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LESSONS IN SCIENCE

A PRELIMINARY COURSE OF

PHYSICS AND CHEMISTRY

BY

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PREFACE

THIS volume is the outcome of a suggestion that the authors' *Lessons in Science*, originally intended to assist pupil teachers and others offering Physics and Chemistry in the Preliminary Examination for the Certificate of the English Board of Education, would, if modified and extended, meet the needs of candidates preparing for the examination in elementary science (Physics and Chemistry) for the Matriculation Examination at the University of Madras.

The present edition will be found to contain a full descriptive and experimental treatment of all the subjects included under Elementary Science in the new Syllabus for Matriculation at the University of Madras, and to cover the work necessary for similar Examinations at other Universities.

Each chapter is provided with a summary of the chief points explained, and is followed by a miscellany of questions set at numerous important public examinations.

The authors gladly take this opportunity to acknowledge the assistance rendered by Mr. G. H. Wyatt, B.Sc., in arranging the present edition, and in writing some of the new sections.

R. A. G.
A. T. S.

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LESSONS IN SCIENCE.

CHAPTER I.

PHYSICAL PROPERTIES AND STATES OF MATTER.

1. PROPERTIES OF MATTER.

i. Porosity.—(*a*) Filter some muddy water through a filter paper placed in a funnel (Fig. 1). The particles of water are small enough to pass through the pores of the paper, but the solid particles of mud are too large, hence they are left behind.

(*b*) Procure a piece of chamois leather; make it into a bag, and pour some mercury into it. Increase the pressure on the mercury by twisting the leather. The mercury is forced through the pores. This is a common way of filtering mercury.

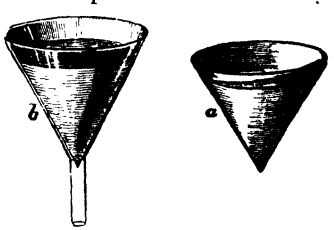


FIG. 1.—Funnel and filter-paper.

(*c*) Half fill a barometer tube with water; then gently add alcohol until the tube is nearly full. Make a mark on the tube at the level with the top of the liquid column, and afterwards shake the tube so as to mix the water and alcohol well together. Observe that the volume of the mixture has diminished, the reason being that some of each liquid has filled up pores between the particles of the other.

ii. Elasticity.—Procure a slab of polished marble or some similar material and smear it with oil. Drop a billiard ball, or a large glass marble, from a considerable height on to the slab. Catch it as it rebounds. Notice that a blot of oil is found where the ball came into contact with the slab. Compare the size of the blot with the spot which is formed when the marble is placed in contact with the slab.

Evidently the ball underwent a compression as the result of collision with the slab, and, by virtue of its elasticity, it regained its original size, causing the rebound.

What is meant by "matter."—Our earliest knowledge of the world teaches us that on every side we have what we familiarly speak of as *things* of all kinds. We become aware of the existence of these things in different ways. Some we feel, some we smell, some we see, some we taste, while others again make their existence known to us by the sounds we hear. On a windy day at the sea-side, standing on the beach, we feel the ground under our feet; we smell, it may be, the tar on a neighbouring boat or the seaweed on the shingle; we see a distant ship at sea or the clouds hurrying across the sky; we taste the salt in the air; and we hear the never-ceasing roar of the waves as they break at our feet. All these things, about which we get to know by our senses, are called material things; they are forms of matter. We must think of **matter**, then, as meaning **all things which exist in or out of our world, which we can become aware of by the help of our senses.**

Different kinds of matter.—Of course the number of different kinds of things is very great, but yet they can all be arranged in three classes, according to certain of the properties they possess, which we shall immediately have to study. The classes are (1) Solid things or solids; (2) Liquid things or liquids; (3) Gaseous things or gases. Sometimes the last two are made into one class and called fluids.

What is meant by "properties."—We shall have occasion to use the word *properties* so often that it will be well to understand clearly what meaning the word conveys, and this can be best accomplished by one or two examples. We say a strawberry is sweet, or a strawberry has the property of sweetness; the paper of this book is white, or the paper possesses the property of whiteness; the sun is bright, or the sun is noted for the property of brightness. Evidently, then, **properties are certain effects caused by the things which are said to possess them.**

Properties possessed by all kinds of matter.—There are certain properties possessed in common by all kinds of matter; these are said to be **general properties.**

1. **Matter must occupy a certain space, or possesses extension;** the larger it is the larger the space occupied by it.

2. **Two material things cannot occupy the same space at the same time.** This property is expressed by saying that matter **is impenetrable.** To be quite accurate, we have to know some-

thing about the way in which matter is built up. We shall see more fully later that every kind of matter is regarded as being built up of small constituent parts, which are incapable of division, and that it is really only these indivisible parts which are impenetrable.

3. **Matter offers resistance.**—We become aware of this, in the case of solids, if we knock ourselves against the wall or the table; if we swim or wade in water we know the same thing is true of water, and we find it to be so of all liquids; if we attempt to run with a screen in front of us we become conscious of the resistance offered by the air to our onward progress, and from this learn that gases, too, offer resistance.

4. **Matter has weight.**—Without knowing the full significance of the expression **weight**, we shall have a sufficiently clear idea of what is meant by this property from its familiar use in every-day conversation. By lifting a solid we become conscious of its possession of this property; if we lift an empty bottle and then when it is full of any liquid, we shall find it is lighter in the first instance, or, as we say, the liquid has weight. By the exercise of sufficient care it can be shown, in just the same manner, that gases have weight.

5. **Matter transfers motion to other things when it strikes against them.**—If we throw a stick at a cocoa-nut on the end of a pole, or send a jet of water at a ball, or blow at a piece of paper, another of the general properties of matter can be demonstrated, namely, the power of giving motion to other things by striking against them. All these general properties can be brought together in a definition—thus: **Matter occupies space, offers resistance, possesses weight, and transfers motion to other things when it strikes against them.**

Other properties of matter.—Matter possesses other characters or properties which it will be useful to study. Though these, too, are general properties, it is possible to form a good elementary notion of matter without taking them into account, and it must be remembered that these properties cannot all be applied to every kind of matter. We shall consider (1) Divisibility, (2) Porosity, (3) Compressibility, (4) Elasticity, (5) Inertia.

Divisibility.—Imagine some material body before you on a table. You know that with suitable means you can divide it

into parts by cutting, that each of the halves can be again divided, and that the bisection can be continued as long as the knife is sufficiently fine and sharp to be able to cut the substance. Evidently, if you could only get sharper and sharper knives, and keener and keener eyes, this process of division could be carried on for a very long time. This property is what is understood by **divisibility.**

Could this division go on for ever? There are reasons for believing that it could not. Eventually a stage would be reached at which each individual part would be indivisible; these parts are called **atoms.** It must be at once understood that atoms have never been seen. We can only imagine what would be the end of our process of division. Our strongest microscopes bring us nowhere near the possibility of seeing an atom.

Porosity.—We are all in the habit of associating this property with certain familiar forms of matter. The sponge we use in the bath has holes through it, or is, as we say, **porous.** A piece of blotting paper is another common example of an obviously porous material; the substances used in filters must also evidently be porous, or else the water would not percolate through them. Porosity refers to the possession of these **interstices** or **pores.** In some cases, though we cannot see these pores with the naked eye, we easily perceive them with a microscope. The pores have often been shown to exist, even where it is difficult to imagine their existence, by forcing water through them. Thus Francis Bacon, in 1604, forced water through a very carefully closed sphere made of lead.

Compressibility.—This property follows as a natural consequence of that just described. If pores exist between the indivisible small particles of which matter is built up, it ought to be possible, by the adoption of suitable means, to make these particles go closer together. This is found to be the case. By pressing upon a body from outside it can be made smaller.

This is well known to be the case in gases, which can actually be made to become successively one-half, one-quarter, one-eighth, and so on up to at least one-hundredth of their original size.

The same thing holds true in the case of solids, though to a much smaller extent. A familiar example of the compression

of solids is seen when a druggist presses a cork between two pieces of iron in order to make it fit a bottle for which it was previously too large. But, generally, in the case of solids the pressure exerted has to be very great to bring about even a small compression.

It was believed for a long time that liquids could not be compressed, but it is now known that they can be slightly reduced in volume, that is, the particles can be forced nearer together.

Thus we learn that **compressibility is not only a consequence of porosity but actually a proof of its existence.**

Elasticity.—Imagine a gas to have been made to assume one-half its size by compressing it. What would happen if the pressure, which is the cause of the diminution, were suddenly removed? The gas would resume its original size or volume, and it would, so far as appearances are concerned, seem to have undergone no change. The gas is said to be perfectly **elastic** and the property which enabled it to go back to its original state is called **elasticity.** Similar results follow with liquids; they also are perfectly elastic.

Some differences arise when solids come to be examined. Though the property can be developed in solids in at least four ways—by **pressure,** by **pulling,** by **bending,** and by **twisting**—we need only in this connection consider the first, as it is the elasticity which is developed by pressure which is most marked in all forms of matter. Ivory, marble, and glass are examples of elastic solids; while putty, clays, fats, and even lead are instances of solids with scarcely any elasticity. In a scientific sense, glass is more perfectly elastic than india-rubber, because it returns to its original shape after it had been forced out of that shape, whereas india-rubber does not return exactly to its original shape.

A solid will only resume its former dimensions when the pressure is removed, provided that the pressure is within a certain limit. If the pressure be more than this minimum amount, or if it exceeds the **limit of elasticity,** as it is called, the sold will not return to the initial size; it will undergo a permanent change. As the student will see later from Expt. 3 i., this limit of elasticity is only exceeded in the case of india-rubber when the pressure applied is very great.

Inertia.—This property will be considered more fully in a later chapter, and here we need only say that it is entirely a negative property; it may be expressed in a general way by the statement that inanimate bodies are incapable by themselves of changing their state of rest or motion.

2. THE THREE STATES OF MATTER.

i. Solid.—Procure a lump of ice and notice that it has a particular shape of its own, which so long as the day is sufficiently cold, remains fixed.

ii. Liquid.—With a sharp brad-awl or the point of a knife break the ice up into pieces, and put a convenient quantity of them into a beaker. Place the beaker in a warm room, or apply heat from a laboratory burner or spirit lamp. The ice disappears, and its place is taken by what we call water. Notice the characters of the water. It has no definite shape, for by tilting the beaker the water can be made to flow about.

iii. Gas.—Replace the beaker over the burner and go on warming the water. Soon the water boils, and is converted into vapour, which spreads itself throughout the air in the room, and seems to disappear. The vapour can only be made visible by blowing cold air at it, when it becomes white and visible, but is really no longer vapour, but has condensed into small drops of water.

iv. Change from solid to gas.—Warm a Florence flask by twirling it between the finger and thumb above the flame of a laboratory burner. When it is too warm to bear the finger upon the bottom, introduce a crystal of iodine, and notice it is at once converted into a beautiful violet vapour.

v. Gradual change of state.—Warm a lump of sealing-wax or bicycle cement in an iron spoon, and notice the gradual conversion into a liquid.

States of matter.—Solids, liquids, gases.—The fact that there are three kinds of material things is well known to every one, and has been stated already. We must now add to this another idea, viz., that **the same matter can exist in three different states.**

In Experiments 2 i.-iii., the same form of matter has been made to assume three states; in other words, ice, water, and steam are the same form of matter in the solid, liquid, and gaseous state respectively.

The change from one state to another may be sudden or gradual.—The circumstances attending the change from the solid to the liquid, or from the liquid to the gaseous state, are not always the same as in the case of water. When solid

iodine is heated, it appears to pass suddenly from the condition of a solid to that of a gas. Camphor is another instance of this sudden transition from solid to vapour. When, on the other hand, sealing-wax is heated, it very gradually passes into the liquid condition, and may be obtained in a kind of transition stage—neither true solid nor true liquid.

There is no hard and fast line between these three conditions of matter. The reader will find as his knowledge extends that intermediate states of matter are known between those described, but for the present it will be best to confine the attention to this simple division, reserving for future work a study of these gradations. We shall now consider the distinctive properties of solids, liquids, and gases.

3. DISTINCTIVE CHARACTERS OF SOLIDS.

i. Stretching of india-rubber.—Fix one end of a piece of india-rubber cord, or tubing, about two feet long, to a support (Fig. 2). Stick two pins through this cord about 18 inches apart. Tie the lower end of the cord into a loop, and then hang a weight by means of a hook from it. Measure the distance between the pins before and after putting on the weight. Repeat the experiment with different weights. You will notice that the cord can be stretched or elongated, because the weights exert a pulling force upon it. If a long metal wire is used instead of the cord, the stretching can be measured in a similar way, but it is much less; for instance, a brass wire $\frac{1}{25}$ inch in diameter and eleven feet long elongates about $\frac{3}{25}$ inch when a weight of 28 lbs. is pulling it.

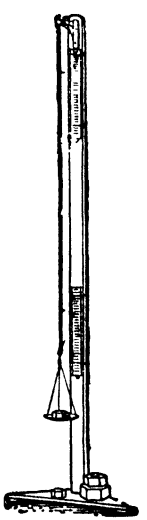


FIG. 2.—Experiment to illustrate the longitudinal stretching of an india-rubber cord.

ii. Bending of a lath.—Procure a flexible wooden lath, and fix it horizontally by clamping one end of it firmly. To the other end attach a pin by means of a little wax. Place a rule vertically near the pin, as in Fig. 3. Hang a weight from the free end of the lath, and observe the amount of bending. Keeping the same weight, clamp the lath so that only half the previous length can be bent, and again notice the amount of bending. Try also with other lengths.

iii. Twisting of a wire.—Suspend a wire with a weight at its lower end, and under it a circle divided into degrees in a manner similar to that shown in Fig. 4. Notice the position at which the weight comes to rest; then twist the weight

through a certain angle, and let it go. The weight untwists back to the starting point and beyond it, and then spins in the way it was twisted, and goes on oscillating in this manner until it comes to

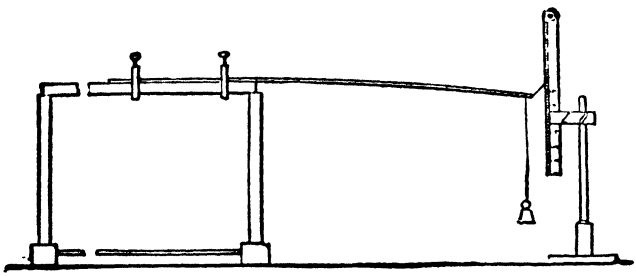


FIG. 3.—Measurement of the bending of a lath.

rest. Observe how long the weight takes to make ten or fifteen complete spins. Repeat the observation with wires of different lengths and diameters, and made of different metals. The time of an oscillation depends upon the tendency to untwist, so that the experiment shows that this depends, among other things, upon the length, diameter, and nature of a wire.

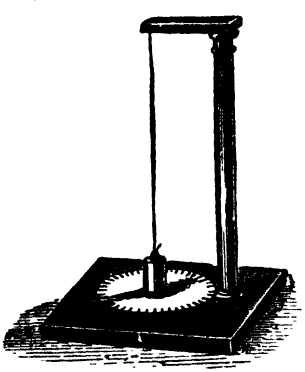


FIG. 4.—Elasticity of torsion.

iv. Tenacity.—Suspend a balance-pan from the lower end of a thin copper wire attached to a beam. Add weights to the pan until the wire breaks. The force required to break the wire is the joint weight of the balance-pan and the weights in it. Repeat the experiment with wires of the same diameter but made of different material.

Ridigity.—A solid body does not readily alter its size or shape. It will keep its own volume and the same form unless subjected to a considerable force.—Another way of expressing the facts contained in the definition is to say that **solids possess rigidity.** Hard solids are more rigid than soft ones. Liquids, on the other hand, possess no rigidity, but the opposite property of **flowing or fluidity.** Liquid particles slide over one another so

easily that the surface of a liquid at rest is always horizontal. Fine sand may be made to flow, but the particles, however, do not move freely over one another, so the surface is left uneven, which fact provides us with a clear means of distinguishing between a solid and a liquid.

Solids possess elasticity.—It has been shown already that elasticity can be called into play in solids by pressure, and also by pulling, bending, or twisting. It would take us too far into the subject of Physics to describe how the elasticity has been measured in these cases, but that there is an alteration of form when solids are treated in these ways is shown clearly by the results of Experiments 3 i.-iii.

Solids possess tenacity, ductility and hardness.—The force required to tear asunder the particles of a body varies with different materials; in other words, some substances are more tenacious than others.

Tenacity is measured by ascertaining what weight is necessary to break solids when in the form of wires.—In making the measurement of tenacity, the area of the cross section of the wire must first be estimated carefully. By the cross section is meant the area of the end of the wire when it is carefully filed to be at right angles to the length. This cross section is estimated by measuring the diameter of the wire and calculating.[1]

It is found that a wire of twice the cross sectional area of another is just twice as tenacious. Evidently, then, if we wish to compare the tenacity of two wires of different materials, it will make the experiment much simpler if wires of the same cross section are selected. Cast steel is the most tenacious of all metals, being about twice as much so as copper and forty times as tenacious as lead. But the tenacity of steel itself is exceeded by that of unspun silk, while single fibres of cotton can support millions of times their own weight without breaking.

Ductility is the property by virtue of which solids can be made into wires.—A ductile material is thus one which can be drawn out. The change of form in this case is produced by pulling. **Malleability** is a similar property to ductility, but the change of form is brought about by the application of pressure; gold,

[1] Area of cross section $= \frac{22}{7} \times \frac{d^2}{4}$.

copper, and lead, for instance, can be beaten out into thin plates, and are therefore malleable substances. Lead is an example of a malleable material which is not ductile—it can be beaten out but cannot be drawn into wires. A solid which can be easily moulded into any shape, which it also retains, is said to be **plastic.** An example is modelling clay.

Platinum is the most ductile and gold the most malleable metal known. Platinum has been drawn out into wire so fine that a mile of it weighs only one and a quarter grains. Gold has been beaten into plates so thin that it would require three hundred thousand of them placed one above the other to make a layer an inch thick.

Hardness is the property by virtue of which solids offer resistance to being scratched or worn by others.—This is a property of great importance in the study of minerals, as it often affords a ready means of distinguishing them. The method of measuring hardness consists in selecting a series of solids, each one of the series being harder than the one above it, and softer than the one below it. At one end of the series, therefore, the hardest solid known is placed; at the other end, the softest which we may wish to measure. A hard solid is frequently also **brittle,** that is, it is easily broken when we attempt to bend it, or when we subject it to a blow. Very hard steel is also very brittle.

4. DISTINCTIVE CHARACTERS OF LIQUIDS.

i. Viscosity.—Procure specimens of treacle and pitch. Soften the latter. Compare the consistency of the treacle and the softened pitch with that of water, and note the gradual increase in the viscosity of the liquids.

ii. Liquid at rest.—Into a shallow glass vessel pour enough mercury to cover the bottom. Attach a ball of lead to the end of a fine string, and so construct a *plumb-line.* Hang it over the surface of the mercury, and notice that the line itself and its reflection are in one and the same line. If this were not the case, that is, if the image slanted away from the plumb-line itself, we should know the surface of the liquid was not horizontal.

iii. Drops.—Sprinkle some powdered resin on a board and then a little water. Notice the water collects in *drops*; the smaller they are, the more nearly spherical they are. Observe the same thing with mercury on a sheet of paper.

iv. Drop of liquid inside another liquid.—Mix methylated spirit and water until a few drops of oil just float when the mixture is

quite cool. Pour fresh oil, by means of a pipette, into the middle of the mixture. Notice spherical globes of oil can thus be formed.

v. **Drops can unite.**—Observe that drops once formed can be made to run together again by coming in contact.

vi. **Cohesion.**—Carefully clean and polish two pieces of plate glass, and then place one on the other. The two surfaces will be found to *cohere*, and considerable force will be required to separate them.

A Liquid easily changes its shape but preserves its volume.—We have now to learn the leading properties which liquids possess, which distinguish them from solids on one hand and from gases on the other. We have learned already that, being forms of matter, they have certain general characters in common with all other material things; but what is there about a liquid which makes us give it a name of its own? **A liquid adapts itself to the shape of the vessel containing it, but the conditions remaining the same, it keeps its own size or volume, however much its shape may vary. When it is not held by the sides of a vessel it at once flows.** This is the commonest everyday experience. You cannot get a pint of water into a glass of half a pint capacity. It does not matter what the shape of the bottle or jug may be—providing it holds a pint, as we say, or provided its capacity is a pint, the quantity of water taken to fill it exactly is always the same. If we turn the jug upside down, the water all runs away because there is no part of the vessel to prevent it from flowing. Moreover, the surface of liquids enclosed by a vessel is always level.

The flow of liquids. Viscosity.—The power of flowing is not perfect in liquids. The small particles making up the liquid always stick to one another a little, and when any part of a mass of liquid moves, it always attempts to drag the neighbouring particle, which is at rest, with it. We can sum this up by saying that liquids would flow perfectly if they possessed no **viscosity.** Those liquids which have little viscosity, or, what is the same thing, are very **mobile** liquids, are instanced by alcohol and water; while treacle and tar have little **mobility,** but are very **viscous.** Evidently, then, there is a gradation in those forms of matter which have as yet come before our notice. At one end we have very mobile liquids, which as the viscosity increases flow less and less easily, until at last there is no power of flowing at all, and we have the solid form of matter.

Liquids find their level.—If several vessels of the most

varied shapes (Fig. 5) are in communication with one another, and water be poured into any one of them, we shall find that as soon as the water has come to rest it will stand at the same level in all the tubes, however different the form of the vessels may be. It is this property of liquids which is utilised in the construction of the water-level. This consists of a horizontal tube connected to a small length at each end held vertically. The upright portions of the tube are of glass, so that the level of the water in them may be easily seen. However the arrangement of tubes may be standing, the line joining the two surfaces

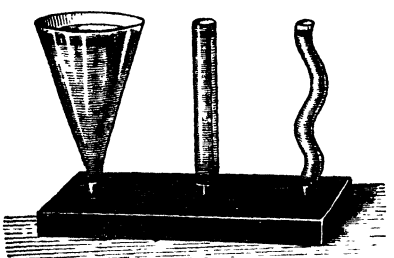


FIG. 5.—Vessels of various shapes in communication. Water standing at same level in all.

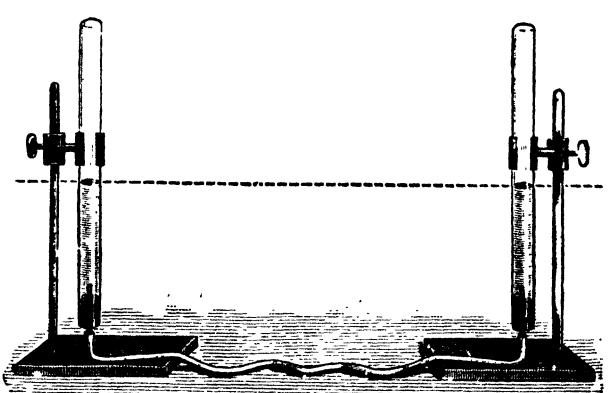


FIG. 6.—Arrangement to show principle of water level.

will always be level. This is of great service to surveyors and such people who must be able to find a level line for the purpose of their observations.

Liquids communicate pressure equally in all directions.—It will be desirable first to satisfy ourselves that liquids

communicate pressure, and then try to understand the second part of the statement, that they communicate it equally in all directions. Imagine we have two cylinders in connection, as shown in Fig. 7. Into each cylinder fit a piston with a plate attached to the top. If we push one piston down, we notice that the other moves up. If we put a weight, say 10 lbs., on to each piston, they will balance one another exactly, and there is no movement in either case. Each piston is pressed upwards and downwards to an equal degree, and in consequence does not move at all. The experiment can be varied by making one cylinder horizontal, while the other remains vertical. In this case, also, if we push in the horizontal piston we notice that the upright piston rises in just the same manner as before, showing us that pressure is communicated *in different directions*. Fig. 8 shows a rubber ball pierced with small holes, and containing water. On squeezing the ball it is easily seen that pressure is communicated by a liquid equally in all directions.

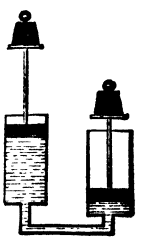


FIG. 7.—To illustrate hydrostatic pressure. The two cylinders are of the same size, so equal weights upon the pistons balance one another.

Hydraulic press.—Reverting to the first example in the last paragraph, and glancing at Fig. 7, suppose that the surface of the piston on the right had been twice as great as the other, and that, as before, weights of 10 lbs. had been placed upon each piston. They would no longer balance; the right-hand weight would be pushed upwards, and to bring about a balance it would be found necessary to put 20 lbs. on the larger piston. Similarly, had the right-hand piston been a hundred times larger, we should have to put 1,000 lbs. upon it to bring about a balance. The upward force, then, is proportional to the extent of the surface of the piston. This principle, which seems so different from what we should naturally expect, is referred to as the *hydrostatic*,

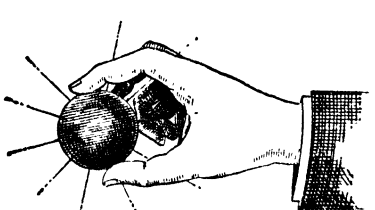


FIG. 8.—Equality of liquid pressure.

paradox, and is utilised in the hydraulic press, called, after its inventor, the **Bramah press**, which is shown in Fig. 9. Here we have exactly the condition of things just described, two cylinders in connection, with pistons fitted into them, one much larger than the other. The application of a comparatively small force to the small one is felt on the larger one, and it is as many times greater in an upward direction as the piston of *B* is larger in area than the piston in *A*. This great upward force is being used in the instance shown in the diagram to compress bales of wool.

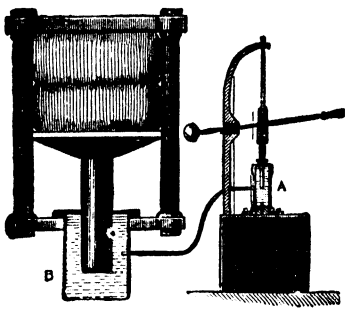


FIG. 9.—Hydraulic press.

Liquids can be separated into drops which will run together again.—Liquids are able to form drops because of a property, known as **cohesion**, which they possess. The larger the drop which can be formed the greater is the cohesion between the particles of the liquid.

Cohesion is the mutual attraction which the particles of a body exert upon one another; it is, indeed, the force which keeps the particles of a substance together. Cohesion is strongest in solids, which in its absence would crumble into powder; it acts between the particles of liquids, but in gases it may be said to be absent.

Cohesion differs from another similar property possessed by some kinds of matter and known as **adhesion.** Adhesion denotes the attraction between unlike particles of matter. A metal plate may be made to adhere to glass; and a postage stamp may be said to exhibit the existence of the force of adhesion when it is stuck upon an envelope.

Experiment 4 iv. makes clear the difference between this adhesion exhibited between the postage stamp and the envelope and the force of cohesion.

Characters of gases.—It has already been explained that the leading difference between a solid and a liquid is the power of

flowing which the latter possesses. Gases also possess fluidity, and to a much more marked degree than liquids. But whereas liquids are almost incompressible, gases are very easily compressed into a much smaller space according to a definite law, viz., just in that ratio in which you *increase* the pressure on a gas do you *decrease* the volume which it occupies, the temperature remaining the same throughout the process. Nor are these the only differences. A liquid always adapts itself to the shape of the containing vessel, and presents a level surface; a gas, on the other hand, will, however small its volume, immediately spread out and do its best to fill a vessel, however large; and it does not present any surface to the surrounding air. We can never say exactly where the gas leaves off and the air begins. Another distinction will be more fully appreciated after we have considered the action of heat upon the volume of bodies. We shall learn that, generally speaking, all bodies become larger as they are heated; this is very much more decidedly the case with gases than with liquids. **Gases, then, are easily compressible and expand indefinitely.**

We shall learn that gases expand equally when heated to the same extent, but this fact and others will be much better understood in their proper places.

5. MATTER CANNOT BE DESTROYED.

i. Weight of water and steam.—Boil water in a flask or a retort, as in Fig. 236, and catch the condensed steam, taking care that none escapes, in another flask kept cool by resting in water. The water thus collected will be found to have the same weight as that boiled away.

ii. Weight of ice and water.—Place a piece of ice in a flask suspended from one arm of a balance. Counterpoise the flask with the ice in it; then melt the ice by warming the flask, and show that the counterpoise is unaltered.

iii. Weight of salt after dissolving.—Put some warm water in a flask and some salt in a piece of paper. Counterpoise the flask of water and the paper of salt together and then dissolve the salt in the water. The total weight remains unaltered.

iv. A burning candle produces moisture.—Over a burning candle hold a white glass bottle which has been carefully dried inside and out. Observe that the inside of the bottle becomes covered with mist and after a short time drops of liquid are formed which run down the sides of the bottle. *The burning of the candle has resulted in the production of a new form of matter.*

v. A burning candle produces a gas.—Allow the candle to burn in a similar bottle placed on the table. After a time the candle ceases to burn. When this has happened take the candle out and cover the bottle over with a glass plate. Notice that no change seems to have taken place in the gas which filled the bottle. Now pour some clear lime-water into another clean bottle and shake it up, the lime-water remains almost if not quite clear. Lift off the glass plate and do the same with the bottle in which the candle has been burnt; the lime-water becomes milky. *The burning of the candle has also resulted in the formation of a new kind of matter, viz., a gas which causes clear lime-water to become cloudy.*

Constancy of weight in different states.—When a solid is converted into a liquid, or a liquid into a vapour, no change of weight is experienced. This has been found to be true in all cases.

No kind of matter can be destroyed.—There is a certain fixed amount of matter in the universe which never gets any less and never any greater. If we confine our attention to the earth, we cannot say that it never receives an addition to the matter of which it is built, for every year it is receiving numbers of small solid bodies which are continually falling upon its surface from outside space. But the proposition means that in those cases in which it is popularly supposed there is a loss of matter, for instance when a fire burns out, no such destruction has taken place, but only a change in the form assumed by the matter. It will make the statement quite clear if we follow out what really takes place when a candle burns, and, as it would seem, gradually disappears.

When a candle burns it ceases to exist as tallow or wax, or whatever the candle is made of, and assumes new forms, still material, one liquid, the other a gas which makes lime-water become milky. If we weigh all the liquid formed and all the gas which renders lime-water milky, we find that the weight of these two things together is actually *more* than that of the part of the candle which has disappeared. The reason why there is an *increase* of weight will be explained later. The arrangement for performing this experiment is shown in Fig. 10. The candle is burned in a wide tube, fitted with a cork at the bottom with holes in it to allow the air, which is necessary to help the candle to burn, to pass in as shown by the bent arrows. The tube is filled with a substance which has the power of arresting the products of the burning. Such a substance is caustic soda,

which is used in the form of lumps. Before the experiment is started, the apparatus is weighed. After the candle has burnt for a few minutes, it will be found that the tube has increased in weight. We are quite sure, therefore, that there has been no loss of matter.

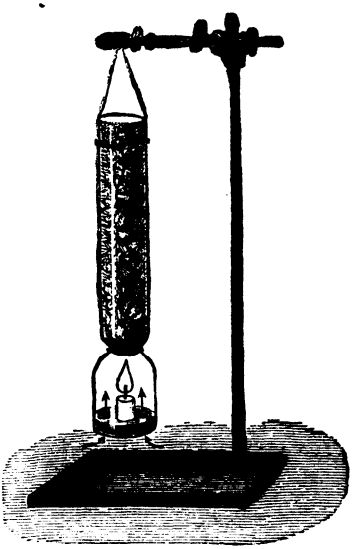


FIG. 10.

Similar experiments to these with candles may be also carried out with ordinary lamp oil or petroleum. The same two substances are produced as the result of the burning, and the same proportion of air is used up. Replacing the lamp glass and candle by a small oil lamp we could show that there is in this case also no loss of weight.

Chemists have satisfied themselves that this fact is universally true, and it must be remembered as a truth of the highest importance, **matter is indestructible.**

CHIEF POINTS OF CHAPTER I.

Matter.—When we speak of matter we mean all things which exist in or out of our world, which we can become aware of by the help of our senses.

Properties are certain effects caused by the things which are said to possess them.

Properties possessed by matter.—Matter occupies space, offers resistance, possesses weight, and transfers motion to other things when it strikes against them.

The same matter can exist in three different states.—The three states of matter are the **solid, liquid** and **gaseous.** By suitable means, as in the experiments described in the chapter, the same portion of matter can be made to assume these states in order. Sometimes the change from one state to another is gradual, some-

times sudden. There is no hard and fast line between the three conditions of matter.

Distinctive characters of solids.—A solid body does not readily alter its size and shape. It will keep its own volume and the same form unless subjected to a considerable force. Or, solids possess rigidity. Solids communicate pressure in one direction only.

Solids possess elasticity, tenacity, ductility and hardness.

Elasticity is the tendency to go back to the original form or volume after being forced out of it.

Tenacity is measured by ascertaining what weight is necessary to break solids when in the form of wires.

Ductility is the property by virtue of which solids can be made into wires.

Malleability is a similar property to ductility which enables certain solids to be beaten out into sheets.

Hardness is the property by virtue of which solids offer resistance to being scratched or worn by others.

Distinctive characters of liquids.—A liquid adapts itself to the shape of the vessel containing it, but the conditions remaining the same, it keeps its own size or volume, however much its shape may vary.

Liquids possess fluidity.—When a liquid is not held by the sides of a vessel, it at once flows. This power of flowing is not perfect in liquids, or liquids are not perfectly mobile. They all possess a certain degree of *viscosity*.

Other properties of liquids.—They find their level. They communicate pressure equally in all directions. They can be separated into drops which will run together again.

Distinctive characters of gases.—They possess fluidity to a much more marked degree than liquids. Unlike liquids they can easily be compressed into a much smaller space. A gas, however small its volume, will spread out and do its best to fill a vessel, however large. Expressed shortly, gases are easily compressible and expand indefinitely.

Constancy of weight in different states.—When a solid is converted into a liquid, or a liquid into a vapour, no change of weight is experienced.

Matter is indestructible.—There is a certain fixed amount of matter in the universe which never becomes any less and never any greater. Whatever changes may occur in the composition of matter there is never any loss of weight.

QUESTIONS ON CHAPTER I.

1. What do you understand by "matter," or, a "material thing"?

2. Give some of the properties which are possessed by all kinds of matter and explain in your own words what is meant by a property.

3. Describe experiments which prove :—
   (*a*) That solids are porous.
   (*b*) That liquids contain pores.

4. What experiment could you perform to show that a solid, say a billiard ball, is elastic ? Explain as well as you can what you mean by elasticity.

5. The same portion of matter can, under suitable conditions, assume different states. Describe fully some experiment which illustrates this statement.

6. What evidence can you give that the different states of matter gradually shade into one another ?

7. What properties are generally associated with matter in the solid form ? Give a definition of a solid which includes the chief of these.

8. In what respects are liquids different from solids ?

9. How do gases and liquids differ ?

10. What property in particular is possessed by liquids and not by solids ? And what character has a gas which neither liquids nor solids possess ?

11. What property is it which liquids possess which enables them to form drops ? Describe another experiment which also shows the possession of this property by liquids.

12. Give in the form of a definition the distinctive properties of a liquid. What name do we give to the property which prevents liquids from being perfectly fluid and what do you know about it ?

13. Describe experiments in support of the assertion that liquids communicate pressure equally in all directions.

14. What is meant by the "hydrostatic paradox" ? Describe some useful apparatus in which this principle is utilised.

15. How would you show by experiment that the weight of the same portion of matter in different states is constant ?

16. On what grounds is the assertion based that "Matter is indestructible" ?

CHAPTER II.

MEASUREMENT OF SPACE.

6. LENGTH.

i. Simple measurements.—Obtain a rule divided into inches and eighths or sixteenths of an inch on one edge, and tenths on the other edge. Measure the length, or width, of any convenient object, such as the top of a table or desk. Write down the length in feet, inches, and fractions of an inch, thus:

Width of desk, 2 feet, $3\frac{5}{8}$ inches.

Length of sheet of foolscap paper, 1 foot, $1\frac{1}{2}$ inches.

ii. Decimal fractions.—Look at the part of a rule having inches divided into tenths. How many tenths are there in half an inch? The tenths of an inch can be written like common fractions, for instance, $\frac{1}{10}$ signifies one-tenth, and $\frac{3}{10}$ means three-tenths. A more convenient way is to separate the inches from the tenths by means of a dot. Thus, a length of $6\frac{3}{10}$ in. is written on this *Decimal System* as 6·3. The two ways of writing tenths in common and decimal fractions may be compared as follows:

Common fractions	$\frac{1}{10}$	$\frac{2}{10}$	$\frac{3}{10}$	$\frac{4}{10}$	$\frac{5}{10}$	$\frac{6}{10}$	$\frac{7}{10}$	$\frac{8}{10}$	$\frac{9}{10}$	$\frac{10}{10}$
Decimal fractions	0·1,	0·2,	0·3,	0·4,	0·5,	0·6,	0·7,	0·8,	0·9,	1

iii. Measurements in tenths.—Measure one or two small objects as before, using the part of your rule divided into inches and tenths. Write down the number of inches and tenths; thus, length of a new lead pencil, 7·1 inches.

iv. Metric measures or length.—(*a*) Look at a rule divided into inches and parts of an inch on one edge, and metric measures on the other, as in Fig. 11. The smallest divisions upon the metric scale are *millimetres* (*mm.*); 10 of these millimetres make 1 *centimetre* (*cm.*), 10 centimetres make a *decimetre* (*dm.*), and 10 decimetres make 1 *metre*. A metre thus contains 10 decimetres, 100 centimetres, and 1000 millimetres.

(*b*) Find how many millimetres there are in the length of this page. Write down the result in (1) millimetres, (2) centimetres and tenths of a centimetre, (3) decimetre and tenths and hundredths of a decimetre.

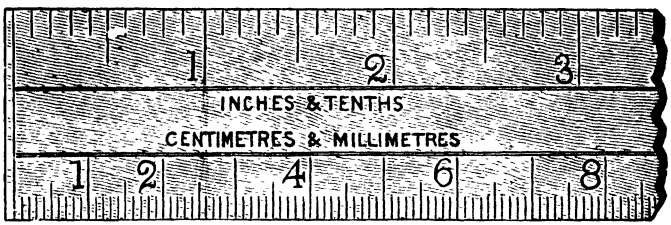


FIG. 11.—A ruler showing Inches and tenths, Centimetres and Millimetres.

v. Relation between Metric and British measures of length.—(*a*) Measure the length of this page both in inches and centimetres; also determine other lengths in the two systems of measurement.

Put down the results in parallel columns, as shown below, and from them calculate the number of centimetres in an inch.

Length in centimetres.	Length in inches.	No. of centimetres. / No. of inches.

(*b*) Upon the back of an ordinary tape measure, mark off a length equal to 100 centimetres, that is, 1 metre, starting from the point where the inches begin on the other side. Prick a hole through the measure at the point where you mark the length of a metre, then turn over the measure and notice where the hole occurs on the inch scale.

Number of inches in one metre = 39·3.
Number of centimetres in one yard = 91·4.

British measures of length.—To secure uniformity, and provide a permanent length with which others may be compared if necessary, a standard of length has to be agreed upon, and this has been described as follows in a publication of the Standards Department of the Board of Trade: "In the **Imperial system of measures** the **yard** is the unit of length, and from this measure are derived all other measures of extension based on that system, whether linear, superficial, or solid. It is represented by the distance between two fine lines

traced on a metal bar which is deposited with the Board of Trade. This bar is made of bronze or gun metal; it is 38 inches long, and one square inch in section, the defining lines being traced on gold plugs or pins inserted at the bottom of a cylindrical hole near to each end of the bar."

Metric measures of length.—When new standards of measurement were being considered in France, in 1795, French geometricians decided that an arbitrary standard such as our yard, in view of the chance of its loss or destruction, was an undesirable one, and they suggested that if a fraction of the circumference of the earth were taken, it would be possible in the event of the loss of the standard, to replace it by an exact copy. They proposed the one ten-millionth part of the distance from the earth's equator to the pole– that is, of the earth's quadrant—as a suitable length, and this they called the **metre.** After bars had been prepared of this length it was unfortunately found that the length of the quadrant had not been exactly determined, and consequently the length of the standard metre at Sèvres is arbitrary. An exact copy of this standard is kept in this country; and for our purposes the metre may be defined as "the length, at the temperature of 0° C., of the iridio-platinum bar, numbered 16, deposited with the Board of Trade." It is equal to 39 37079 inches, or briefly, 3 feet 3 3 inches. The metre is subdivided into ten equal parts, each of which is called a **decimetre**; the tenth part of the decimetre is called a **centimetre,** and the tenth part of the centimetre is known as a **millimetre.** Thus, we get

10 millimetres = 1 centimetre
10 centimetres } = 1 decimetre
100 millimetres }
10 decimetres }
100 centimetres } = 1 metre
1000 millimetres }

The multiples of the metre are named **deka-, hekto-,** and **kilo-metres.** Their value is seen from the following table:

10 metres = 1 dekametre
100 metres = 1 hektometre
1000 metres = 1 kilometre

The kilometre is equal to about five-eighths of a mile, or eight kilometres are equal to five miles.

7. AREA.

i. Square inches.—Mark off lengths an inch apart from one corner along the bottom edge of a page of your exercise book, or of a sheet of paper. Draw lines straight up the page from each of these marks perpendicular to the edge of the paper. From the same corner mark off lengths of inches on the side edge of your paper, and draw lines across the paper parallel to the bottom edge. You will thus divide the paper into square inches. Count how many square inches there are in one row, neglecting the parts of a square inch which may happen to be over at the end, and then count the number of rows. By multiplying these two numbers together, you will evidently obtain the total number of square inches.

ii. Length and width.—Draw an oblong 4 inches in length, and 3 inches in width or height, and divide it into square inches as before. Notice that the number of square inches is equal to the length multiplied by the height.

iii. Rule for areas of oblongs.—The preceding exercises have shown that, to find the area of a square or an oblong, it is only necessary to multiply the length of the bottom or base by the height.

iv. Square inch and square centimetre.—(*a*) Draw two rectangles and determine their areas both in square inches and square centimetres; use your results to find the number of square centimetres in one square inch, thus:

Area of a given rectangle in square inches.	Area of same rectangle in square centimetres.	Square centimetres. / Square inches.

(*b*) Draw a square inch and divide it up into square centimetres as shown in Fig. 12. Notice that it takes about six and a quarter square centimetres to make one square inch.

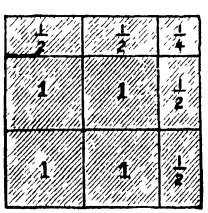


FIG. 12.—Square inch divided into square centimetres.

Measurement of area.—If the floor of a room had to be carpeted, it would not be enough to measure the length of the room only, or its width only, because both of these are measures of length. To know how much carpet is necessary, the amount of surface the floor has, or its **area** must be found. To do this, both the length and width of the floor must be measured, and if the room is square

or oblong, the area can be found by multiplying the two numbers together. When the length and width are measured in feet, then, by multiplying them together, the area of the floor in square feet is obtained; when the measurements of the length and width are taken in inches, the area in square inches is obtained by multiplying the two numbers together.

Whenever areas are measured in this country, square inches, square feet, square miles, or some other unit from square

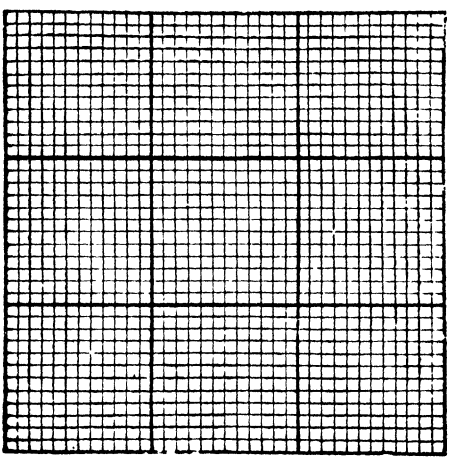


FIG. 13.—This represents a square yard on a small scale. Each small square represents a square inch, and the large squares bounded by thick lines represent square feet.

measure is employed. "Square measure" is obtained from "long measure" by multiplying. Thus, as there are 12 inches in a foot, there are 12×12 square inches in a square foot. This will be understood by examining Fig. 13 Each of the large squares bounded by thick lines represents a square foot, but it is of course smaller than a real square foot.

In the metric system also it is the custom to refer to an area as of so many *square* centimetres, or *square* metres, as the case may be.

Fig. 14 shows, on a scale giving one-quarter the actual area, a square decimetre divided into square centimetres. The area, as drawn, being one-quarter of a decimetre, the length of each side is half a decimetre or five centimetres.

10									
9									
8									
7									
6									
5									
4									
3									
2									
1	2	3	4	5	6	7	8	9	10

FIG 14.—Method of dividing a square decimetre into square centimetres.

8. VOLUME.

1. **Cubic inch.**—Obtain a bar of soap, and cut one end so that it is square with two sides. From one corner, where the end and these two faces meet, mark off the length of an inch along the three edges. Scratch a line along the soap from each of the points thus obtained, and square with the edge. You will then have three square inches

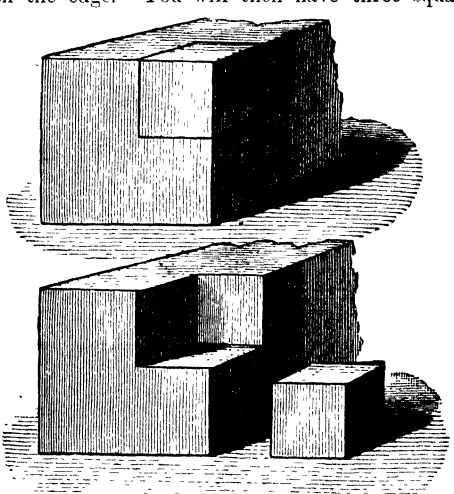


FIG 15.—How to mark and cut out a cubic inch or cubic centimetre from a piece of soap

marked upon the soap. By cutting deep along these lines a cube of soap having each of its edges an inch in length, and each of its faces a square inch in area, is obtained. This is a cubic inch (Fig. 15).

To test whether your cubic inch is true it could be compared with one of the wooden cubes sold for kindergarten teaching.

ii. Cubic centimetre.—Using soap, cheese, clay, or similar material as before, cut one end square with the faces, and then mark off a length of one centimetre along three edges meeting at one corner. Scratch lines from each of these points so as to mark three square centimetres upon the material used, and then cut along the lines to obtain one cubic centimetre, as in the previous exercise. Compare this cube with the cubic inch. If you have time and material, cut a rod of plastic substance sixteen centimetres long and one square centimetre at each end. Divide this into sixteen cubic centimetres, and then show that the cube made by pressing these together is of about the same size as a cubic inch.

iii. Standard capacity boxes.—(*a*) Cut out of cardboard a figure of the shape shown in Fig. 16, and of the size indicated by the numbers.

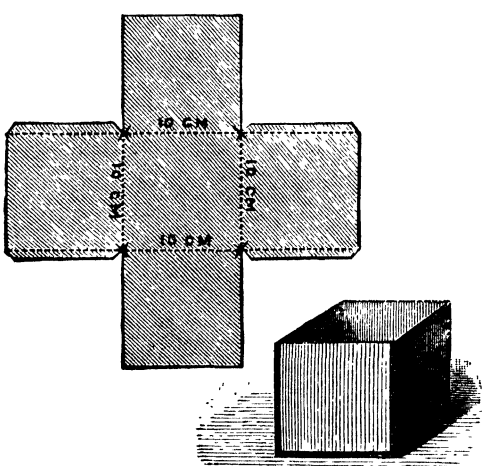


FIG. 16.—Construction of a box to hold 1 cubic decimetre, that is, 1 litre.

Cut the cardboard half-way through at the dotted lines, and then bend it to form a cubical box. Bind the edges together with glued tape, and varnish the box inside and out to make it watertight. This box will hold one cubic decimetre of liquid, and this capacity is called one **litre**.

(*b*) Make a similar box having sides one inch in length, so that it will hold one cubic inch. Also, if time permits, make a box to hold one cubic centimetre.

iv. Cubic inch and cubic foot.—Procure a box or a block one cubic foot in size. Divide the top and the other faces of the cube into square inches (Fig. 17). Notice that the area of each face of the cubic foot is one square foot. Count the number of square inches marked on one face. Notice that 144 cubic inches could be cut out of a slab of the cube one inch thick. How many slabs having a thickness of one inch could be cut from one cubic foot, and how many cubic inches are there altogether in such a cube ?

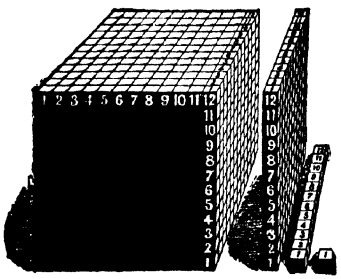


FIG. 17.—To show the relation between 1 cubic inch and 1 cubic foot.

v. Cubic centimetre and cubic decimetre.—Using a block one cubic decimetre in size, divide the top and faces into square centimetres. How many square centimetres are there in each face ? How many cubic centimetres are there in a slab one centimetre thick ? How many cubic centimetres go to make one cubic decimetre ? (Fig. 18).

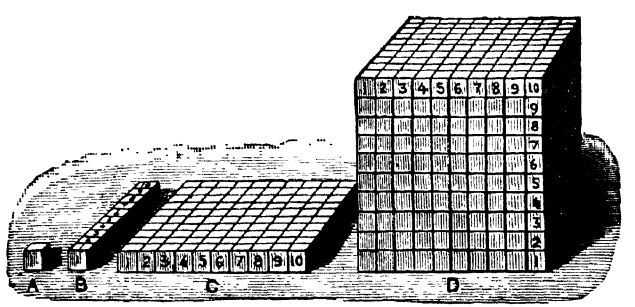


FIG. 18.—To show the relation between 1 cubic centimetre and 1 cubic decimetre

vi. Fluid measure.—Obtain a half-pint graduated glass measure, such as is used by photographers ; obtain also a smaller one (Fig. 21). Examine the divisions upon them : they represent fluid ounces, or parts of a fluid ounce, in Apothecaries' measure.

vii. Graduated measuring jar.—(*a*) Obtain a jar graduated into cubic centimetres, like that in Fig. 19. Test the accuracy of your cubic decimetre box by pouring 1000 cubic centimetres of water, as measured by the graduated jar, into the box, and noticing whether the water just fills the box.

viii. Relation between British and metric measures of capacity.—(*a*) Use graduated jars to determine: (1) The number of cub. cms. in 1 fluid ounce; (2) the number of ounces and drachms in 100 cub. cms.; (3) the number of ounces and drachms in 1000 cub. cms.

(*b*) Fill a measure up to the half-pint mark with water. Pour the water into the metric graduated jar, and so find the number of centimetres in a half-pint. Repeat the operation, so as to determine the number of cubic centimetres in one pint.

(*c*) Pour 1000 c.c., that is, 1 litre, of water into a vessel. Then measure this amount of liquid in pints, so as to find the value of 1000 c.c. in British measure.

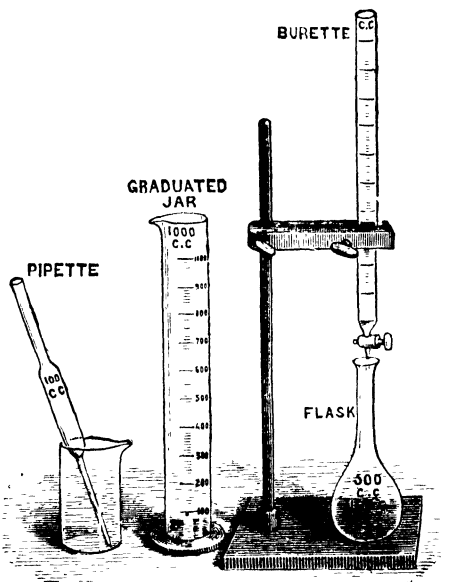


FIG. 19.—Graduated measuring vessels.

ix. Volumes of solids determined by displacement.—(*a*) Fill a metric measuring jar about half way up with water. Notice the level. Hold a cubic inch of wood in the water by means of a long pin. Notice the rise of level. The difference between the two levels evidently represents the number of cubic centimetres equal to one cubic inch.

(*b*) Determine the volume of a stone or any irregular solid by observing the difference of level when it is immersed in water (Fig. 20).

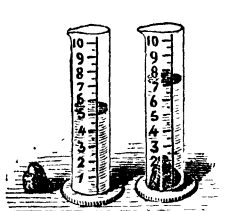


FIG. 20.—The rise of level of the water when the stone is put in shows the volume of the stone.

x. Burette.—Examine a burette (Fig. 20). Notice that the divisions are numbered from the top downwards. Fix the burette upright in a clamp, and about half fill it with water. Notice the level. Find the volume of a stick of slate pencil by gently lowering it into the water and noticing the rise of level produced. (The pencil must, of course, be completely immersed in water.)

Measurement of volume.—The **volume** of a thing is its size, or bulk, expressed in proper units. In dealing with volumes, three dimensions have to be considered. Just as a plane surface or area measuring one foot in each of the directions—length and breadth—is called a square foot from the name of the figure which it forms, so a solid which is obtained by measuring a foot in three directions—length, breadth, and thickness—is called a **cubic** foot, from the name **cube** given to the solid so formed. Similarly using the **Metric system**, we may speak of a cubic metre, or a cubic decimetre.

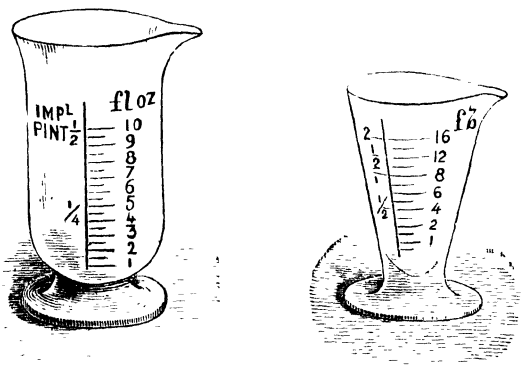


FIG. 21.—Glasses graduated for British fluid measure.

A hollow vessel is capable of holding a certain volume, and this is usually referred to as the **capacity** of the vessel. In the metric system a special name is given to the capacity of a hollow cubic decimetre, that is, a hollow cube having a decimetre edge. It is called a **litre**, and is equal to about one and three-quarters English pints. The sub-multiples and multiples of a litre are named in a similar way to those of the metre. There is no such simple relation between the measures of length and volume in the English system, though the gallon is defined as a measure which contains 10 lbs. of pure water at a certain temperature and pressure. A gallon has a volume of 277¼ cubic inches.

In measuring the volumes of irregular solids, advantage is usually taken of the fact that when immersed in a fluid they displace a volume of fluid equal to their own volume. **This may**

be measured as a difference of level (Fig. 20). The water displaced could also be poured into a cubic inch box or into a cubic centimetre box, and the number of cubic inches or cubic centimetres could be thus found. Or, we could get a glass measure having cubic inches or cubic centimetres marked upon it, and pour the displaced water into it. But the best plan of all is to use a vessel having cubic centimetres marked upon it. Water can be put in such a vessel up to a certain mark, and the number of cubic centimetres of water displaced by the solid can be seen at once by noticing the number of divisions between the levels of the water before and after the solid is put in.

CHIEF POINTS OF CHAPTER II.

Measurement of length.—In estimating lengths, a standard, or unit, of length is necessary.

The *British standard of length* is the *Imperial standard yard.* It may be defined as "the length, at 62° F., marked on a bronze bar deposited with the Board of Trade."

The yard is *divided* into three equal parts, each called a *foot.* A foot is divided into twelve equal parts, each called an *inch.*

The *Metric standard of length* is called a *metre.* In the metric system all measures are connected with one another by *tens* or *tenths.*

The metre is divided into ten equal parts called *decimetres*, each of which is again divided into ten equal parts called *centimetres*, and these in turn are divided into ten equal parts called *millimetres.*

A length which contains exactly 10 metres is called a *dekametre*, one which just contains 10 dekametres, or 100 metres, is called a *hektometre*, and one which is exactly 1000 times as long as a metre is called a *kilometre.*

Measurement of area.—To estimate area, or surface, it is necessary to measure in two directions, *i.e.* to find the length and breadth of the surface.

By multiplying together the length and breadth of a square or an oblong we obtain its area.

The expressions for area are obtained by prefixing the word *square* to the corresponding unit of length thus: square millimetre, square centimetre, etc.

A square centimetre is *not* $\frac{1}{100}$ of a square metre but $\frac{1}{100} \times \frac{1}{100} = \frac{1}{10,000}$ part.

Cubic measure.—The *volume* of a solid is its *size*, or the *amount of room* it takes up, or the *space it occupies.*

The volume of a rectangular solid is obtained by multiplying together the length, breadth, and thickness, the measurements being at right angles to one another.

Metric Measures of Length and Capacity.

Fractions and Multiples.	Length.	Capacity.
$\frac{1}{1000}$	Milli-metre	Milli-litre
$\frac{1}{100}$	Centi-metre	Centi-litre
$\frac{1}{10}$	Deci-metre	Deci-litre
1	**Metre**	**Litre**
10	Deka-metre	Deka-litre
100	Hekto-metre	Hekto-litre
1000	Kilo-metre	Kilo-litre

British and Metric Equivalents.

1 inch = about $2\frac{1}{2}$ centimetres.	1 cm. = about $\frac{2}{5}$ in.
1 foot = ,, 3 decimetres.	1 dm. = ,, 4 in.
1 yard = ,, $\frac{9}{10}$ metre.	1 m. = ,, 3 ft. $3\frac{3}{8}$ in.
1 mile = ,, $1\frac{3}{5}$ kilometre.	1 km. = ,, $\frac{5}{8}$ mile.
1 gallon = ,, $4\frac{1}{2}$ litres.	1 litre = ,, $1\frac{3}{4}$ pints.

EXERCISES ON CHAPTER II.

1. Multiply 10·4 square centimetres by 15·5 decimetres, and state the result both in cubic centimetres and in litres.

If the volume in question were filled with water at 4° C., what would the weight of the water be?

2. What is the cubical content, in litres, of a box of which the inside dimensions are as follows: length 25 centimetres, breadth 12 centimetres, and depth 8 centimetres? What is the weight in kilograms of the water it would hold?

3. One pint is equal to 34·7 cubic inches, and 1 inch is equal to 2·54 centimetres. How many pints are there in 1000 cubic centimetres?

What is the name given to a volume of 1000 cubic centimetres?

4. How would you determine the volume of a pebble in cubic centimetres?

5. Explain how the metric units of volume and length are related. Is there any such simple relation in the case of British units?

6. What is meant by a unit of length, and why is it necessary to have such a unit?

7. Give particulars of the units of length in use in this country and on the Continent, and give your opinion as to whether one of

the units you give is more convenient than the others, and if so, why?

9. Express (*a*) a millimetre as a decimal of a centimetre, (*b*) as a decimal of a decimetre.

10. What fraction is (*a*) 1 millimetre of 1 inch, (*b*) 1 decimetre of 1 foot, (*c*) 1 centimetre of 1 inch?

11. If 25 millimetres are equal to one inch, how many sq. millimetres make 1 sq. inch?

CHAPTER III.

UNITS OF TIME.

9. UNITS OF TIME.

i. A pendulum. – Attach a weight to the end of a cord. Fix the cord in such a way that the pendulum can oscillate freely. Set it oscillating, and notice how long it takes for the pendulum to complete a given number, say twelve, swings. Keeping the cord exactly the same length, attach a heavier weight and repeat the experiment. The time of swing remains unaltered. Keeping any one weight, observe the time taken to complete twelve swings when the length of the cord is varied. It will be seen that the time of swing varies with the length of the cord. Notice also that it does not matter whether the pendulum makes a wide oscillation or a very small one, the time taken being the same in each case.

ii. The sun-dial. – Fasten a small rod at right angles to a flat board. Place the board flat on a table so that the rod is vertical. Move a candle in a semicircle above the table, and note the change in the angle made by the shadow of the rod. Compare the conditions of the experiment with the measurement of the solar day by means of a sun-dial.

The earth's rotation. – The apparent daily motion of the sun and stars across the sky is a direct consequence of the earth's rotation on its axis. The sun appears to go regularly through certain periodic changes of position. It rises, travels higher and higher into the sky, reaches its highest position, sinks lower and lower, and finally sets. When the sun is at its highest altitude on any day it is due south, and is said to **south** or be **southing.** The interval of time between the sun's reaching its highest position on any one day and its corresponding position on the next succeeding day is an **apparent solar day.** These apparent solar days vary in length throughout the year.

Mean solar day.—As the length of days measured by the sun varies throughout the year, no single one of these days will do for a convenient standard of time. But if the lengths of all the days in the year be added together, or the length of a year measured by the sun be divided by the number of days in the year, we obtain an interval of time which is always the same. Such a day, which is of course an imaginary period of time, is called a **mean solar day.** Sometimes the mean solar day is longer than the solar day, sometimes it is shorter, and occasionally both days are of exactly the same length. Solar time is known as **apparent time,** and clock time as **mean time.**

Sidereal day.—Just as in the case of the sun so with all the stars; they rise, south, and set. But whereas with the sun the interval between two successive southings varies throughout the year, it is found that the time which elapses between two succeeding southings of a star at any season of the year is always the same. This interval constitutes **a star** or **sidereal day.**

Period of rotation of the earth.—As the apparent motions of stars across the sky are produced by the rotation of the earth, it will be evident that the exact time of rotation can be determined by finding the interval which elapses between two successive returns of any particular star to the same point of the sky. A star may, indeed, be regarded as a fixed reference mark under which the earth turns; so that by observing it we are able to determine the time taken by the earth to spin round once. The interval between two successive transits of the same star, or, as it is called, a sidereal day, is the time of such rotation.

No matter which star be selected for observation the interval is the same, thus showing that the earth is a rigid body, and that all parts of its surface have the same angular velocity.

Units of time.—The sidereal day, like the mean solar day, is subdivided into hours, minutes, and seconds, but as the latter is four minutes longer than the former, the units are not of the same value. We may take either the **mean solar second** as the unit of time, or the **sidereal second.** In the former case the unit is founded on the average length of the solar day, and in the latter upon the length of the invariable star-day, or the time of rotation of the earth upon its axis. But in either case the

second, that is, the unit of time, is the 86,400th part of the day used.

In physical measurements the unit of time adopted is the mean solar second, that is, it is founded on the average time required by the earth to make one complete rotation on its axis relatively to the sun considered as a fixed point of reference.

Instruments for measuring time.—We need only concern ourselves with the modern contrivances for measuring time, viz., clocks and watches. It will be sufficient to regard these as instruments for measuring intervals of time in terms of the mean solar day to which attention has been directed. In a clock the rate is regulated by means of the pendulum. The chief property of which is the constant time of oscillation at the same place.

If it were possible for the student to perform the experiment, it would be found that the time taken for the pendulum to swing backwards and forwards varies as it is taken from the equator to the poles, on account of the fact that the earth is not exactly spherical in shape. Or, putting the same fact in another way, in order that a pendulum may swing backwards and forwards in the same interval of time, it is necessary to alter the length of the cord in our experiment as we travel from the equator towards either pole. A pendulum of such length that the distance from the point of suspension to the centre of the bob is 39·139 inches, swinging at Greenwich, completes one swing in a second of time. In a clock we have a mechanical contrivance for maintaining the swinging of a pendulum. We must content ourselves with referring the reader to books on astronomy and horology for an account of the construction of a clock. In watches the place of the pendulum is taken by a carefully suspended balance wheel.

CHIEF POINTS OF CHAPTER III.

The solar day is the interval of time between the sun's reaching its highest position on any one day and its corresponding position on the next succeeding day.

The mean solar day is the quotient obtained by dividing the length of the year measured by the sun by the number of days in the year. Or, it is the average of the lengths of all the solar days in the year.

A star or **sidereal day** is of constant length. It is the interval of time between two succeeding southings of a star at any season of the year.

The period of rotation of the earth.—As the apparent motions of the stars across the sky are produced by the rotation of the earth, the exact time of rotation is determined by ascertaining the length of the sidereal day.

Units of time.—In physical measurements the unit of time adopted is the mean solar second, that is, it is founded on the average time required by the earth to make one complete rotation on its axis relatively to the sun considered as a fixed point of reference.

QUESTIONS ON CHAPTER III.

1. What is the difference between a solar day and a mean solar day?

2. Define a sidereal day and explain how it differs from a solar day.

3. What is the common unit of time? How is it related to the period of the earth's rotation?

4. Give a brief description of some mechanism in common use for measuring time.

CHAPTER IV.

MOTION; INERTIA; FORCE; NEWTON'S LAWS.

10. MOTION AND VELOCITY.

i. Motion.—Place a number of marbles in a tray. Disturb them by shaking the tray, or in any other way: they all acquire motion, but move in various directions, and at different rates.

ii. Speed.—Shoot a marble along a table, or let it run round the rim of a tray. The marble travels with a certain speed or rate of motion, and it may be given this speed in any direction.

iii. Velocity.—Bowl a marble along a table in a definite direction, and observe as nearly as possible the number of seconds taken to perform the journey. The distance in feet thus traversed, divided by the number of seconds occupied in traversing it, gives the average rate of motion in feet per second. The velocity of a body is its rate of motion in a definite direction.

iv. Uniform velocity.—Make several lines crosswise on a long table at a distance of one foot apart. Push a cylinder or marble along the table at such a rate that the length from one line to the next is traversed in one second. The velocity of the object is thus one foot per second and is uniform during the movement.

v. Variable velocity.—Shoot a marble or roll a cylinder along the table with the marks upon it, and notice the number of seconds taken in passing from the first to the last mark. In this case the same total distance is traversed as before, but the velocity is variable, that is, the body does not move through equal spaces in equal times throughout its movement as it is gradually coming to rest. If a leng'h of 8 feet is traversed in 4 seconds, what is the average velocity?

vi. Graphic representation of velocity.—(*a*) Taking a line an inch long to represent a velocity of one foot per second, draw lines representing velocities of $3\frac{1}{2}$, $2\frac{3}{4}$, 4, and $1\frac{1}{4}$ feet per second, making the lengths of the lines proportional to the rates of motion.

(*b*) Draw a line to represent the velocity of a river flowing at the rate of 2 miles an hour. Suppose a man who can row 6 miles an

hour in still water to be rowing in this river. Draw lines to represent his velocity with reference to the bank when rowing (1) with the stream, (2) against the stream.

(c) Draw a large circle upon a sheet of paper. Draw two diameters at right angles to each other, as in Fig. 22. Taking ½ inch to represent a velocity of 1 foot per second, and starting from the centre of the circle, represent by graphic construction the path of a body moving with the following velocities: 2 feet per sec. N.E.; 2 feet per sec. N.; 3 feet per sec. W.; 4 feet per sec. S.E.

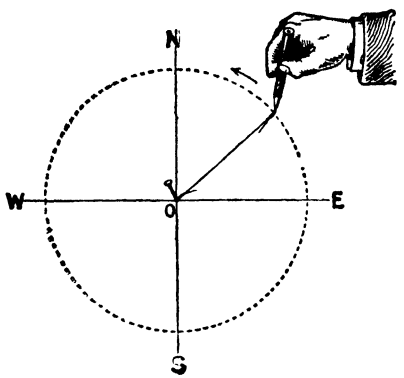


FIG. 22.

vii. Combination of two velocities.—Place a wide glass tube along a table near one edge. Let a marble roll through the tube, and as it does so roll the tube across the table. Suppose a second is taken for the tube to move from its first to its last position, then the positions of the marble at intervals of a tenth of a second are shown in Fig. 23. The marble enters at *A* and emerges at *D*; the distance it has travelled in the direction of the length of the table is therefore represented by *AB*. The distance

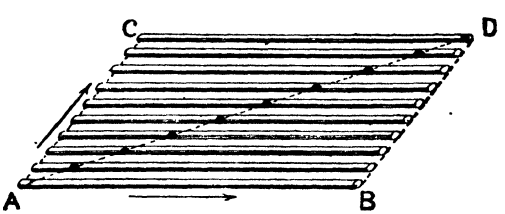


FIG. 23.—To illustrate the combination of two velocities, and the principle of the parallelogram of velocities.

it has travelled *across* the table is in the same way represented by *AC*. Construct a parallelogram *CABD* with lines proportional to *CA* and *AB* as sides. The diagonal *AD* represents the path actually traversed by the marble.

Definition of motion.—The word **motion** is meant to convey the idea of **change of place.** The simplest forms of motion are

changes in the positions of bodies with regard to one another. The change of position may take place slowly or quickly, that is, motion may be fast or slow. The rate at which a body moves is called its **speed.** A ship may have a speed of twenty knots (*i.e.* may change its position at the rate of twenty nautical miles in an hour). A runner may maintain a speed of ten miles per hour. In these statements the direction of the motion has not been considered.

When a boy runs down the street he is in motion ; as regards the houses and lamp-posts he moves. To describe fully the boy's motion it would be necessary to know his **velocity,** that is, the **rate** with which he travels, and the **direction** in which he is moving, or the **line** along which he runs. If during every second through which he moves he travels over a line of five yards in length, he has a **uniform** velocity of five yards a second.

But suppose he does not move regularly over five yards in every second ; he sometimes dawdles, sometimes stops to look at a shop, at other times he puts on a spurt to make up for lost time. How should we describe his motion now? His rate varies from time to time, or his velocity is **variable,** and to describe such a variable velocity it is usual to speak of the velocity **at any instant** as being a certain number of yards per second. Suppose the boy moving with a variable velocity had at a given instant a velocity of eight yards per second. If he continued to move at the same rate he would travel over eight yards in the succeeding second.

Average velocity.—But it is sometimes better to find the **average velocity** of the moving body. Returning to the boy, suppose he travelled 800 yards in 400 seconds ; if we divide the first number by the second we obtain the boy's average rate, namely, two yards in a second ; this, then, is the rate with which he would have had to travel, if he moved uniformly, in order to complete his journey in the same time.

The **unit of velocity** is generally taken as being a velocity of one foot per second. Thus a velocity of six means a velocity of six feet per second.

Measurement of uniform linear velocity.—It is a simple matter to calculate the velocity of a body moving uniformly in a straight line when the distance it has travelled, measured in units of length, and the time it has taken to perform the

journey, measured in units of time, are known. All that has to be done to find its uniform velocity is to divide the number of units of length passed over by the number of units of time taken to complete the distance. Thus :

$$\text{Uniform velocity} = \frac{\text{space traversed}}{\text{time taken}}.$$

Velocities can be completely represented by straight lines.—To determine a velocity completely, all we want to know are its magnitude (or the distance travelled in a given time) and its direction. But, as is well known, a straight line can be drawn of any length in any direction ; and we can arrange that its length shall contain as many inches or feet, whichever is more convenient, as there are feet or yards per second of velocity, depending on the way in which it is decided to measure the velocities. Velocities can therefore be completely represented by straight lines.

Composition of velocities.—Think of the case in Experiment 10 vii., of a marble moving along a tube with a uniform velocity, when the tube itself is all the time being uniformly moved across a table. It is evident that since the marble is in the tube it must have the same velocity *across* the table that the tube has ; and at the same time it moves along the tube, that is, in a direction at right angles to its former velocity. It has two independent velocities. Similarly, we can think of a ship sailing across the ocean with a man on deck walking from one side of the ship to the other. The man has two velocities. He is moving onwards with the ship at a certain velocity, and at the same time he is moving across the ship with another velocity.

But yet a body can only move at any instant in one direction with one definite velocity. How, then, shall we find the actual velocity at any instant in the case of the marble or of the man ? The velocity which we want to find is called the **resultant** of the two independent velocities, which are themselves spoken of as **components.** If the two velocities have the same direction, all we have to do is to add them to obtain the resultant, or if they are in opposite directions along the same straight line we subtract them.

If they have directions which make an angle with one another, it is clear that the resultant must be somewhere

between the components. Referring to the case of the marble, let *OA* (Fig. 24) represent by its length the number of inches the marble moves along the tube in a second, and *OB* the distance moved by the tube, and consequently by the marble in the same time across the table. Draw *BR* parallel to *OA* and *AR* parallel to *OB*, thus completing the parallelogram, then the line *OR* represents the resultant velocity, both in magnitude and direction. In just the same way *OA* could stand for the ship's velocity and *OB* for the man's, then *OR* would represent the direction and magnitude of the man's actual velocity with reference to the sea. This principle is called the **Parallelogram of Velocities.**

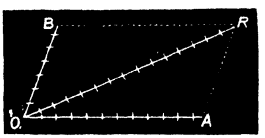


FIG. 24.—Parallelogram of velocities.

11. ACCELERATION.

1. **An instance of accelerated velocity.**—Obtain a smooth board about six feet long, having a slight groove cut in it from one end to the other (Fig. 25). Incline the board slightly at one end. Place a marble or other small sphere near the raised end and let it roll down the board. Notice that as it moves its velocity increases. To show

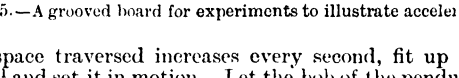


FIG. 25.—A grooved board for experiments to illustrate acceleration.

that the space traversed increases every second, fit up a seconds' pendulum[1] and set it in motion. Let the bob of the pendulum strike against a sheet of paper or some other light object at the end of each swing, so that you can hear when the seconds commence. Now start the marble from a mark upon the board exactly when the pendulum taps the paper on one side. Notice how far the marble has rolled by the time the pendulum taps the paper on the other side. Make a mark at the place reached, and do the same for succeeding seconds until the marble rolls off the board. Measure the length of board traversed by the marble in each second. The distances will be found to increase in proceeding down the board from the starting place.

[1] A seconds' pendulum can be made by suspending, from the ring of a retort stand, a bullet by a string 39·14 inches in length.

Meaning of acceleration.—An express train starting from a terminus begins to move slowly, and, as the journey proceeds, the rate of motion goes on increasing until the train gets its full speed. A stone let fall from a height similarly starts from rest, and as it moves it travels faster and faster until brought to a standstill again on reaching the ground. Or, imagine a cyclist starting for a ride, and regularly increasing his speed until he could not go any faster. In all these examples the velocity of the moving body has periodically increased, and the rate at which the change has taken place is spoken of as **acceleration.**

Acceleration is the rate of change of velocity.—But acceleration may be of an opposite kind to the instances given above. Reverse each of the examples and consider what happens. An express train going at full speed approaches a station and its velocity is regularly diminished until it is brought to rest at the platform. A stone is thrown upwards with a certain velocity, it moves more slowly and more slowly until it comes to rest, and then starts falling. A cyclist travelling at full speed slackens his rate regularly until he comes to a standstill. In all these cases we have examples of an acceleration of an exactly opposite kind to the previous instances, but yet an acceleration. In ordinary language this kind of acceleration is given a name of its own, viz. **retardation.**

Accelerations, like velocities, can be represented graphically by straight lines.

12. INERTIA.

i. A mass at rest tends to remain so.—Place a heavy mass upon a trolley resting upon a level table. Attach a piece of thin cotton or thread to the trolley. Attempt to move the trolley by sharply pulling the thread, which will probably break. The trolley can, however, easily be moved if the pulling force is applied gradually.

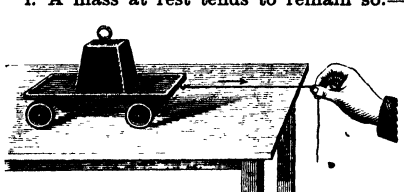


FIG. 26.—Experiment to illustrate the tendency of a body to preserve its state of rest or motion.

ii. A moving mass requires force to stop it.—Fix one end of the thread attached to the loaded trolley and let the thread hang loosely between the trolley and the fixed point. Set the trolley

in motion away from the fixed point. The force of the tendency of the mass to keep in motion will probably be sufficient to break the thread.

The force of gravitation.—Experiments and observations made by Newton led him to the conclusion that it was the rule of nature for every material object to attract every other object, and that this force of attraction is proportional to the masses of the bodies ; a large mass exerts a greater force of attraction than a small mass. But the farther these bodies are apart the less is the attraction between them, though it is not less in the proportion of this distance, but in that of the square of the distance. This diminution of a force according to the inverse proportion of the square of the distance applies to so many cases that it ought to be understood clearly before proceeding. Suppose two bodies of equal mass are one foot away from one another and attract each other with a certain force. If the distance between the masses is doubled, the strength of the attraction between them is only one-quarter of what it was ; for the square of 2 is $2 \times 2 = 4$ and the inverse of 4 is $\frac{1}{4}$. In the same way, if the bodies are placed three feet apart, the force of attraction is $\frac{1}{9}$ of the original force.

Newton's law of gravitation may be expressed thus : **Every body in nature attracts every other body with a force directly proportional to the product of their masses and inversely proportional to the square of the distance between the bodies ; and the direction of the force is in the line joining the centres of the bodies.**

Consider the case of a cricket ball on the top of a house. The earth attracts the ball, and, by Newton's law, the ball attracts the earth. The ball, if free to move, falls to the earth ; to be correct, however, we must think of the ball and the earth moving to meet one another along the line joining their centres. But the ball moves as much farther than the earth as the earth's mass is greater than that of the ball ; and for practical purposes this is the same as saying that only the ball moves and that the earth remains still.

This force of attraction between all material bodies is called the **force of gravitation**, but this is merely giving the force a name. Calling this force "gravitation," and the rule according to which it acts the "law of gravitation," does not teach anything about the nature of the force itself.

The attraction of gravity.—Bearing in mind that weight is really a measure of the attraction between an object and the earth (p. 59), it will be clear from Newton's law of gravitation that since a thing up in a balloon is farther away from the earth (which acts exactly as if its whole mass were collected at its centre) than when at the sea-level, the weight of this thing ought to be less than it is at the sea-level. This is found to be the case, but, actually to demonstrate the difference in weight, the weight must be measured by a spring balance as in Experiment 15 1. (c)

Similarly, because the earth is not a perfect sphere, but is flattened at the poles, points at the surface of the earth in the region of the tropics are at a greater distance from the centre than points similarly situated in the neighbourhood of the poles. Consequently, the weight of a mass situated on the earth in the tropics should be less than the weight it would have if it were moved into the polar regions. This has been found to be the case.

The rotation of the earth is another disturbing influence. While places on the equator are carried round with a velocity of over a thousand miles an hour, those near the poles have but a very small velocity of rotation, while the pole itself is at rest. It is clear that if we consider a mass at the equator its tendency is to obey the first law of motion (p. 45), and to fly off at a tangent, and part of the force of gravitation is expended in preventing this flight—the remainder of the force of gravity is operative as the weight of the mass under consideration. At the pole there is no tendency to move off tangentially, and the whole of the force of gravitation is felt as the weight of the body. For this reason alone the mass would weigh less at the equator.

Inertia.—Common experience tells everyone that things do not move of themselves. An object at rest remains at rest until it is forced to move. Moreover, if it is moving it tends to go on moving in the same direction and with the same velocity until made to change by the application of force. In a word, dead matter is helpless and conservative. It does nothing by itself, and objects to alter its condition whatever that condition may be. **The inabilty shown by a material body to change by itself its condition of rest or of uniform motion in a straight line is called its**

inertia. It is exemplified when a cyclist is stopped suddenly, for the tendency to continue moving is so great that the cyclist, if he is travelling quickly, is shot over the handle-bar of his machine. This law of inertia is often spoken of as Newton's First Law of Motion.

First law of motion.—**Every object remains at rest or moves with uniform velocity in a straight line until compelled by force to act otherwise.**

This law, which Newton first stated as being always obeyed by bodies in nature, means, first, that if a body is at rest, it will remain at rest until there is some reason for its moving—until some outside influence, which is called a **force**, acts upon it. In fact, the law really supplies us with a definition of force. Nobody finds any difficulty in understanding the rule so far. But it is not so easy to see the meaning of the words referring to uniform motion in a straight line. An example will make this clear. Consider a ball moving uniformly along ice. After a time the ball comes to rest, and therefore it does not continue in a state of uniform motion. But it moves for a longer time on ice than it would do on a road. The ice is smoother than the road, and there is a connection between the roughness or smoothness and the length of time during which the ball moves. If the ice could be made smoother and smoother, the ball would move for a longer and longer time, and if both the ball and the ice were perfectly smooth, there is no reason why the ball should ever stop. The roughness or friction represents, then, the force which causes the ball to change its state of uniform motion for one of rest. If a body in a state of uniform motion could be placed outside the influence of what Newton called "impressed forces" it would afford us an example of *perpetual motion*. But because these impressed forces cannot be eliminated perpetual motion is impossible.

Definition of force.—Newton's first law, then, enables force to be defined. **Force is that which produces, or tends to produce, motion in matter; or alters, or tends to alter, the existing motion of matter.** It must, however, be understood clearly that by defining force we do not get to know anything more about it. Nobody can tell what force is; all we can know are the effects produced by it.

13. MOMENTUM.

i. Collision of equal masses.—Suspend two solid balls of equal mass by the side of one another, so that they just touch when at rest. Place a metre rule on edge on the table under the balls (Fig. 27). Notice the positions of the balls, and then pull each ball to an equal distance from the place where the balls rest, by means of threads held in the two hands. Release the two threads at the same instant. The balls will collide, and if they are of equal mass they will rebound by equal amounts.

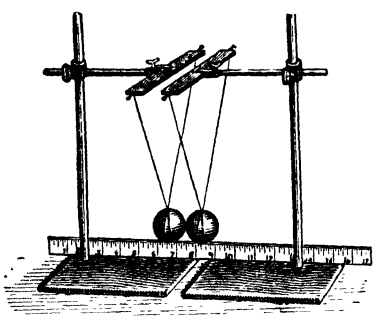


FIG. 27.—Apparatus to illustrate momentum and impact.

ii. Momentum after collision.—Substitute a ball of much smaller mass for one of the balls in the preceding experiment. Pull the two balls away from the position of rest, so that the surfaces which touch one another are at equal distances from this position. Then release the balls simultaneously and observe the rebound. The two balls were at equal distances from their position of rest and are suspended from strings of equal length; therefore their velocities when they meet are equal. The smaller mass has less momentum than the larger before the two meet, but the momentum after collision is the same for each mass; therefore the small mass is given a velocity of rebound as many times greater than the velocity of the large mass as this mass is greater than the smaller one.

Meaning of momentum.—The momentum of a body is equal to the product of the mass and the velocity. Expressed as an equation we have

$$\text{Momentum} = \text{mass} \times \text{velocity}.$$

The unit of momentum is consequently that of a unit of mass moving with a unit of velocity, or, if the unit mass be that of the imperial standard pound, the unit of momentum is the quantity of motion in a mass of one pound moving with a velocity of one foot per second.

Unit of force.—A definition of the unit of force may now be given. It has been explained that force is the name given to the something which is able to set a body at rest in motion or

to change the motion of a moving body. **A unit force acting for the unit of time is able to produce a unit of velocity in a unit of mass ; or a unit of force produces a unit of acceleration in a unit of mass.** But, since the product of a mass and its velocity is spoken of as the momentum of the body, we can measure force by the momentum it generates. The unit force gives rise to the unit of momentum.

Change of momentum is proportional to the acting force. The force which will produce a velocity of one foot per second in a mass of twenty pounds will produce a velocity of two feet per second in a mass of ten pounds, and a velocity of twenty feet per second in a mass of one pound, acting in each case for the same time. Or to express these facts concisely in an equation :

$$\text{force which acts} = \frac{\text{mass of the body}}{\text{acted on}} \times \frac{\text{change of velocity produced}}{\text{in a unit of time}}$$

$$= \frac{\text{mass of the body}}{\text{acted on}} \times \text{acceleration.}$$

Newton's second law of motion.—Change of motion is proportional to the impressed force, and takes place in the direction in which that force acts.

This law speaks of "change of motion," and by motion is understood momentum or quantity of motion. It states that the momentum generated by a force of two units will be twice as great as that produced by one unit ; and further it implies that a force of one unit acting for two seconds will produce twice the momentum which it would do if it only acted for one second. This is why it is necessary in defining the unit of force to introduce the words "acting for the unit of time." The change of motion spoken of in the law must therefore be understood to mean that produced in a unit of time, or the *rate* of change of motion or momentum. But we know that the momentum of a body is the product of the mass of the body into the velocity, and since the mass remains constant, the rate of change of momentum must be the product of the mass and the rate of change of velocity, which is, as we have seen, the acceleration. Hence we come to the very important fact that **the number of units of force in any force is equal to the product of the number of units of mass in any body on which it may act, and the number of units of acceleration produced in that mass by the force in question.**

Newton's third law of motion.—Action and reaction are equal and opposite ; or the mutual actions of two bodies on one another are always equal and in opposite directions.

This law can be exemplified in a variety of ways. If you press upon the table with your hand with a certain force, the law asserts that your hand is pressed by the table with an equal force in the opposite direction. We say in common language simply that the table offers resistance. Similarly, if you hold a mass of a pound on the palm of your outstretched hand it presses your hand downwards with a force which we have learnt to call its weight ; your hand presses the weight upwards with a force of exactly the same amount. If the scrimmage at Rugby football does not move, it is clear that the opposing "forwards" are pressing one another with equal forces in opposite directions, and it is easy to understand in this case that the action and reaction are equal and opposite. Every force, then, is one of a pair of forces, the second being the reaction to which it gives rise. Such a pair of forces is called a **stress.**

14. PARALLELOGRAM OF FORCES.

i. Relation between tension and extension.—Clamp a thin india-rubber cord with a tray suspended from the lower end, and arrange a wire by its side as in Fig. 28. Stick a pin through the cord opposite the end of the wire. Place 10 grams in the tray and measure the elongation produced by it—that is, measure the distance between the pin and the bottom of the wire. Place another 10 grams in the tray and again notice the extension produced. By subtracting this reading from the former one, you can find the extra extension produced by the additional 10 grams. In the same way find the extension produced by 30 grams, 40 grams, and so on, up to 100 grams, also determining in each case the extension for the additional 10 grams.

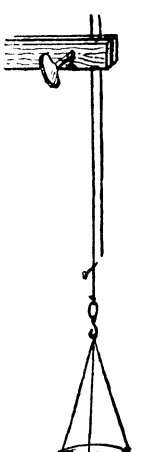


FIG. 28.—Experiment to show the extensions produced by different forces.

ii. Representation of forces by lines.—Draw vertical lines to represent the extensions produced by the various loads, and also lines to represent the extension for every additional 10 grams. As the loads represent forces, the lines will represent the extension produced by various forces. You will find that the extensions are proportional to the forces—that is, double the force double the extension, and so on. *The strength or magnitude of a force can thus be shown by the length of a line.*

iii. Parallelogram of forces.—The following experiments illustrate how the *directions* as well as the *magnitudes* of forces can be represented by lines.

(*a*) To the middle of a piece of elastic cord about 6 inches long fasten a piece of the same cord about three inches in length. From the point where the three cords meet measure off an equal distance, say two inches, and put a pin through each cord at this distance.

Fasten the pin at each end of the six-inch piece of cord upon a sheet of paper lying flat upon a drawing board or table. The cord should lie in a straight line between the pins, but it must not be stretched. Now pull the third cord in any direction so as to stretch the others as well as itself, and fasten its pin in the board while the cords are thus stretched (Fig. 29). Make a pin mark upon the paper at the point where the cords meet, and then pull the pins out of the paper and put them and the cords aside. Pin-holes will be marked at *ABCD* in Fig. 29. Mark off from each of the lines, starting from *A*, *B*, and *D* respectively, the original length, that is, 2 inches, measured on each cord. The additional lengths *EC*, *FC*, and *GC* thus represent the amounts by which the cords have stretched.

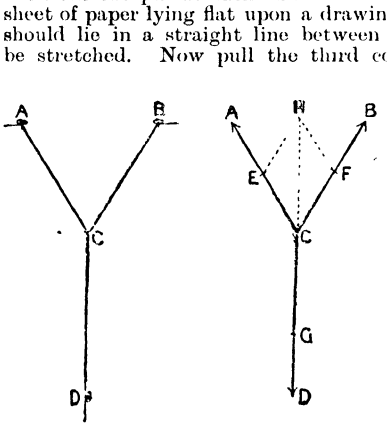


FIG 29. —A three-way elastic cord, used to illustrate the principle of the parallelogram of forces.

(*b*) Construct the parallelogram *ECFH* by drawing *EH* and *FH* parallel to *CF* and *EC* respectively. Draw the diagonal *CH*; it will be found to have the same length and be in the same line as *CG*. Assuming that the extensions are proportional to the forces (and they practically are so in this case), the two sides *EC*, *CF* of the parallelogram represent two of the forces both in magnitude and direction, and the third force, which is equivalent to the other two taken together, is represented on the same scale by the diagonal of the parallelogram.

(*c*) Repeat the experiment with the cords stretched by different amounts and in different directions. Construct a parallelogram similar to *ECFH* for each case.

When the cords are arranged in any position they represent three forces acting upon the particle *C*. Since *C* is at rest, the force represented by *CG* produces the same effect as the two forces represented by *CE*, *CF*, taken together. If, therefore, a force equal in magnitude to *CG*, and acting in the direction *CH*, is substituted for

CE, *CF*, the particle still remains at rest. The diagonal *CH* represents such a force both in magnitude and direction.

A single force which can be substituted for two separate forces in this way is termed a *resultant* force.

Graphic representation of forces.—Every force has a certain strength or magnitude, and acts in a certain direction. It is, therefore, possible to represent a force completely by a line, the length of which is made proportional to the magnitude of the force, and the direction of which represents the direction in which the force is exerted. If the length of an inch is taken to represent a unit force, then a force of 5 units would be represented by a line 5 inches long, and two forces of 5 and 3 units acting together in the same direction would be represented by a line 8 inches long. If, however, a body were acted upon by a force of 5 units in one direction, and 3 units in the opposite direction, then the effect would be that of a force of 2 units acting in the direction of the foree of 5 units; for 3 of the units of this force would be rendered inoperative by the three units acting in the opposite direction.

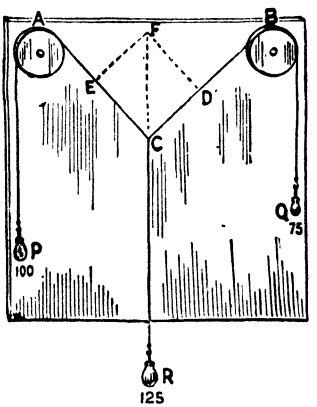


FIG. 30. The parallelogram of forces can be illustrated by connecting two weights *P* and *Q* with a cord passing over two pulleys, and keeping them in equilibrium with a weight *R*.

Parallelogram of forces.—A body can only move in one direction at any given instant, though it may be acted upon by any number of forces. Each force acts with a certain strength and in a certain direction, and, in consequence of their joint action, the body moves with a certain velocity, if it is free to do so. The same velocity could be given to the body by a single force instead of the separate forces, and the single force which would produce the same effect as the separate forces is called the **resultant** of the forces. When two forces act upon a body at the same time, their resultant can usually be found by means of the parallelogram of forces,

which may be expressed thus : **If two forces acting at a point be represented in magnitude and direction by the adjacent sides of a parallelogram, the resultant of these two forces will be represented in magnitude and direction by that diagonal of the parallelogram which passes through this point.**

Let *O* represent a material body acted upon by two forces, represented both in magnitude and direction by the lines *OB*, *OA*. To find the *resultant* of these two forces, both as regards its amount and direction, complete the parallelogram *OBRA* and join *OR*, which will be the resultant required (Fig. 31).

Calculation of resultant.—When the two forces of which the resultant is required act at right angles to one another, the calculation is a simple application of a proposition in the first book of Euclid (I. 47). In these circumstances the triangle *ORA* is right-angled, and Euclid proves that $(OA)^2 + (AR)^2 = (OR)^2$, and consequently $(OA)^2 + (OB)^2 = (OR)^2$, from which when *OA* and *OB* are known we can calculate *OR*.

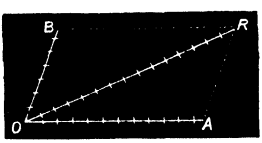


FIG. 31.—Graphic representation of the parallelogram of forces.

When the directions of the two forces *OB* and *OA* are inclined to each other at an angle which is not a right angle, the calculation involves an elementary knowledge of trigonometry. This can be obviated, however, by the simple expedient of what is called the graphical method. This consists in drawing two lines inclined at the angle at which the directions of the forces are inclined, and making them of such lengths that they contain as many units of length as the forces do units of force (Fig. 31). The parallelogram is then completed by drawing *AR* and *BR* parallel respectively to *OB* and *OA* and joining the diagonal *OR*, the direction of which will be that of the resultant, and its length will be as many units as there are units of force in the resultant force. It is immaterial what lengths are used to represent the units of force so long as the components and the resultant are measured in the same units.

Resolution of forces.—A single force can be replaced by other forces which will together produce the same effect. Such a substitution is called **resolving** the force, or a **resolution** of the force. The parts into which it is resolved are spoken of as

components. When this has been done it is clear that we have made the original force become the resultant of certain other forces which have replaced it. Referring back to what has been said about the parallelogram of forces, it will be seen that any single force can have any two components in any directions we like; for by trying, the student will be able to make any straight line become the diagonal of any number of different parallelograms. The most convenient components into which a force can be resolved are those the directions of which are at right angles to each other. In this method of resolution, neither component has any part in the other.

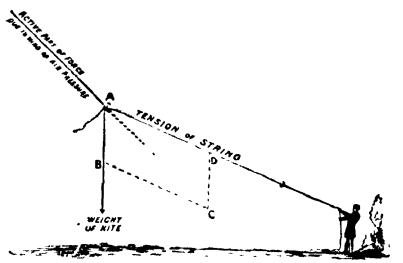


FIG 32.—Forces acting on a kite at rest in the air.

A kite at rest in the air affords an example of the principle of the parallelogram of forces (Fig. 32). There are two downward forces—one represented by AB, due to the weight of the kite, and the other represented by AD, due to the pull of the string. The pressure of the air on the face of the kite can be resolved into two forces, one acting along the face and the other at right angles to it. The latter force is an upward one, and if the kite is at rest it is equal to the resultant AC of the two downward forces. If it is greater than the resultant AC, the kite rises; if it is less, the kite falls.

CHIEF POINTS OF CHAPTER IV.

Motion is change of place. *Speed* is rate of motion in any direction.

Velocity is the rate of motion in a given direction. When a body's velocity is *uniform* it moves over equal distances in every second, when it is *variable* unequal distances are moved over in equal times.

$$\text{Uniform velocity} = \frac{\text{No. of units of length travelled}}{\text{No. of units of time taken}}.$$

The unit of velocity is generally taken as being a velocity of one foot per second. Velocities can be completely represented by straight lines. The resultant of two velocities can be determined by the *parallelogram of velocities.*

Acceleration is the rate of change of velocity. In measuring uniform accelerations it must be known what addition to, or subtraction from, the velocity of the moving body there has been during each second of its journey.

The unit of acceleration is an increase of the unit velocity in a unit of time. It is generally taken as equal to an increase of velocity of one foot per second for each second.

Force is that which changes, or tends to change, a body's state of rest or motion. Forces may be completely represented, both in magnitude and direction, by straight lines.

The **momentum** of a body is the quantity of motion it has, and is equal to the product of its mass and its velocity. The total momentum of several moving masses is unaltered by impact.

A unit of force acting for a unit of time is able to produce a unit of velocity in a unit of mass.

The force of gravitation.—Every particle in nature attracts every other particle with a force which is directly proportional to the product of the masses of the particles and inversely proportional to the square of the distance between them.

$$\text{Gravitational attraction} = \frac{\text{mass} \times \text{mass}}{\text{square of distance between masses}}.$$

Newton's laws.—(1) A body once set in motion and acted upon by no force will move forward in a straight line for ever.

(2) If a moving body be acted upon by any force, its deviation from the motion defined in the first law will be in the direction of the acting force and proportional to it.

(3) Action and reaction are equal and in opposite directions.

Parallelogram of forces.—If two forces acting upon a point are represented in magnitude and direction by the adjacent sides of a parallelogram, their resultant is represented in magnitude and direction by the diagonal which passes through that point.

Resolution of forces.—A single force can be replaced by other forces which will together produce the same effect. Such a substitution is called *resolving* the force, and the parts into which it is resolved are called *components.*

EXERCISES ON CHAPTER IV.

1. What is the difference between *uniform* and *variable* velocities? How are uniform linear velocities measured?

2. Explain how the average velocity of a body moving with a variable velocity is estimated.

3. A line is drawn upon the floor of a railway carriage from door to door. When the carriage is at rest a ball is dropped from the roof and falls upon this line. What difference would be observed:

(*a*) If the train is moving when the ball is dropped?

(*b*) If the train starts when the ball is half way down?

(*c*) If the ball is dropped when the train is in motion, but the train stops suddenly when it is half way down?

4. Define acceleration and give an example of a body moving with an accelerated velocity.

5. How is uniform acceleration measured? Explain the difference between a positive and negative acceleration, giving an example of each.

6. Describe an experiment for demonstrating the principle of the parallelogram of forces to a class.

A nail is driven into a wall and two strings are tied to its head. When the two strings are pulled horizontally and at right angles to one another with forces equal to 6 and 8 lbs. respectively, the nail is dislodged. What force would be needed if the strings were brought together and the nail pulled straight out? Illustrate your answer with a diagram.

7. Explain what is meant by the inertia of a material body. Give as many of the results of the possession of this property by a material body as you can.

8. The horizontal and vertical components of a certain force are equal to the weights of 60 lbs. and 144 lbs. respectively. What is the magnitude of the force?

9. Two spring balances are attached by strings to a ring which is placed round a nail driven into the table, and the balances are stretched in directions at right angles to one another till one indicates a pull of 16 lbs. and the other of 9 lbs. Draw a figure showing the direction and magnitude of a *single* force which would produce the same pressure on the nail as do the *two* forces due to the pulls of the balances.

How would you prove experimentally the correctness of your result?

10. State the "parallelogram of forces." You are provided with three small spring balances (sometimes called "dynamometers"), a blackboard, chalk, string, etc.; how can you verify the proposition?

11. A weight which is hung by means of a piece of elastic from a nail in the ceiling is pulled some way to one side by a thread which is always kept horizontal. Explain why this operation will increase the stretching of the elastic. Illustrate your answer by a diagram.

12. As a train draws up at a station on a rainy day, water flows from the roofs of the carriages by their front ends. As the train leaves the station the water flows off from the rear ends of the carriages. How do you account for this ?

13. A train moving out of a station covers in successive seconds 4, 8, 12, 16, 23, 30, 34, 38, 42, and 44 feet. Name the various kinds of motion with which the train is moving during this period. What is the mean velocity in miles per hour during the last second ?

CHAPTER V.

MEASUREMENT OF MASS, WEIGHT, DENSITY.

15. MASS AND WEIGHT.

i. Meaning of mass and weight.—(*a*) Take two pieces of iron or brass, called in ordinary language, "pound" and "half-pound" weights; or a "pound" and a "two-pound" will do. Lift the two pieces of metal. One feels heavier than the other, that is, the masses are different.

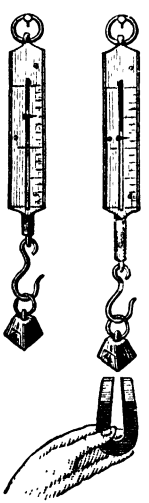


FIG. 33.—The iron weight may be made to appear heavier by a magnet, but the mass does not change.

(*b*) Place a certain amount of lead in one pan of a balance, and counterpoise it with cotton-wool in the other pan. The masses are equal but the volumes are different.

(*c*) Examine the parts of a spring balance (Fig. 33). Attach one ounce to the balance and show that the marker is pulled down to the division 1. The pull of the spring upwards and of the ounce downwards are equal.

(*d*) If possible, using a delicate spring balance, such as is used for weighing letters, show that the downward pull of a mass of iron can be increased by holding a strong magnet under it (Fig 33).

ii. The balance. - (*a*) Uncover the balance and identify the different parts by reference to Fig. 34. Raise the beam, *AB*, of the balance, off the supports by turning the handle *C*. Notice whether the pointer *F* swings equally on both sides of the middle of the scale *G*: if it does the balance is ready for use; but if not, let down the beam and turn the small screw at *B*, then try again. Repeat this adjustment until the swings to right and left are equal.

(*b*) When using the balance *always* place the substance to be weighed in the *left-hand pan*, *E*, of the balance. Examine your box

of weights, and, selecting a weight which you estimate to be about the same as an object placed in the pan, *E*, take hold of it *with*

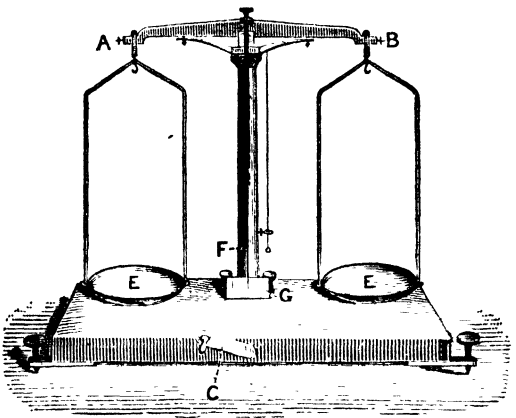


FIG. 34. —The Student's Balance.

the forceps and place it in the *right-hand pan*. Now raise the beam slightly to see whether the estimated "weight" is nearly equal to the mass of the substance in *E*. If the weight appears a little below what is wanted, pick up, with the forceps, the next heaviest weight in the box, and try it in the pan with the other. If the two together are too heavy, take out the smaller weight and put in the one below it, and so on, adding one weight after the other, *without missing any*, until you find the correct weight. When you have completed the weighing, write down and add up the weights that are missing from their places, and check the figures as you put each weight back in its place.

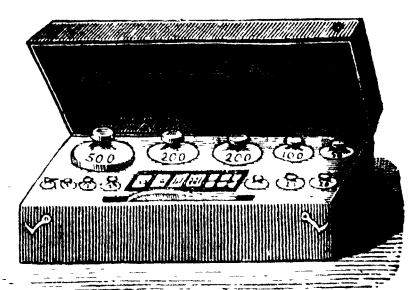


FIG 35.—Box of metric weights.

(c) Examine examples of British masses, *e.g.* an ounce, a pound, a half-hundredweight. Also examine a box of metric masses, generally spoken of as a box of "weights" (Fig. 35).

desired to measure, with standard masses. This is done with the help of a balance, or pair of scales.

In the simple balance the beam, AB, is supported at the centre C, and if the balance is true, the scale pans have no

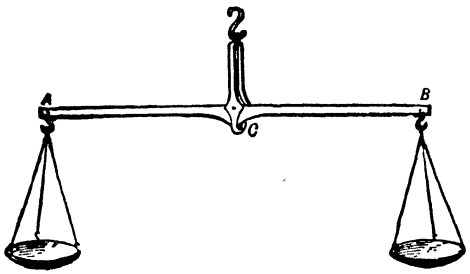


FIG 37.—A simple balance, or pair of scales.

tendency to dip either on one side or the other. When, as in this case, the pans are suspended from points at equal distances from the pivot, any mass in one pan is balanced by an equal mass in the other.

The earth's attraction for equal masses, at any one place, is the same, or shortly stated **equal masses have equal weights.** Thus any weight can be compared with the weight of standard masses. The weight of a body which, as has been stated, is the pull of the earth upon it, may be measured directly by a suitably graduated spring balance.

In ordinary language this distinction between mass and weight is not made. The word weight is used to mean both the amount of matter in a body and also the pull of the earth upon it. Having directed attention to the distinction, we shall in future use the word weight in its ordinary sense.

16. DENSITY.

1. The density of different bodies varies.—(*a*) Determine, by means of a balance, the weight of a cubic centimetre of wood, lead, cork, and marble, and record the results thus:

Weight of the cubic centimetre of

wood (oak)	= 0·82	grams.
lead	= 11·35	,,
cork	= 0·24	,,
marble	= 2·84	,,

(*b*) Place, in one pan of a balance, the cubic centimetre of lead used in the previous experiment and cut a regular piece of soap from a bar so that it just counterpoises the lead. Find the number of cubic centimetres in the soap.

(*c*) Counterpoise two small bottles of the same size. Fill one with water and the other with methylated spirit. Notice that the bottle of water is heavier than the bottle of spirit, though the volume of each liquid is the same.

(*d*) Counterpoise a pint measure or bottle with some sheet lead. Fill the bottle with water, and place metal weights in the opposite pan to balance it. Notice that the size of the metal is much less than the size of the pint of water.

The meaning of density.—1. Pieces of different substances of the same size or volume may have unequal weights.

2. Pieces of different substances which have equal weights may have very different sizes or volumes.

It is usual to speak of these facts by saying that things have different **densities**. A pound of feathers, or cotton-wool, has exactly the same weight as a pound of lead, but the feathers (or the cotton-wool) take up much more room, or have a larger volume, than the piece of lead. The matter in the lead may be packed more closely than in the feathers, which accounts for its taking up less room. The shortest way of saying all this is to say that lead is **denser** than either cotton-wool or feathers.

If a small thing is comparatively heavy, then it is called a *dense* thing, or it is said to have a *high density*. If, on the other hand, a large thing has a small weight, it is said to have a *low density*. Moreover, the densities of things having equal volumes are in the same proportion as their weights.

Standard of density.—But to compare densities it is better to have a standard, just as we have a standard of length, the yard, with which to compare other lengths; or a standard of area with which to compare other areas. The density of water at a certain fixed temperature is taken as the standard. Why it is necessary to specify the temperature will be understood later.

The weight of one cubic centimetre of water at a temperature of 4° C. (see p. 58) is one gram, and its density is taken as the standard of density, and is called 1. Similarly, a substance, the weight of a cubic centimetre of which is two grams, would be said to have a density of 2, for it must contain twice as much matter as water does, packed into one cubic centimetre. The

weight of a cubic centimetre of quicksilver is 13·6 grams, *i.e.* it contains 13·6 times as much matter in one cubic centimetre as there is in one cubic centimetre of water. Its density is therefore 13·6.

Density is the weight of unit volume of a substance.—Suppose a cube of soap and a cube of lead be cut so as to have equal weights; the soap will evidently be larger than the lead, and it will be just as many times larger as its density is less than the density of the lead. With equal weights, the greater the density the smaller is the volume. It follows from this that if the volume of a body is multiplied by its density, we obtain its weight. Or, expressed as an equation,

$$\text{volume} \times \text{density} = \text{weight},$$

from this it follows that

$$\text{density} = \frac{\text{weight}}{\text{volume}}.$$

In using this relation between the volume and weight care must be taken that the values of weight and volume are reduced to the proper units. In all scientific work it is customary to adopt the cubic centimetre and gram as the units of volume and weight respectively.

The ratio of the weight of *any volume* of a substance to the weight of the same volume of water is equal to the **relative density** of the substance, or, as it is frequently called, the *specific gravity*.



FIG. 38.—A bottle for use in determining relative densities.

17. SOME METHODS OF DETERMINING RELATIVE DENSITY.

1. By means of a relative density or specific gravity bottle.—(*a*) Counterpoise an empty specific gravity bottle, or a flask having a mark on its neck (Fig. 38). Fill the flask up to the mark with methylated spirit and weigh it; then empty the flask, dry it, and fill with water up to the same mark. Weigh again, and from the two values thus determined find the relative density of the spirit, remembering that

$$\text{Relative density} = \frac{\text{weight of substance}}{\text{weight of equal vol. of water}}.$$

(*b*) Following the method of the previous experiment, determine the relative densities of two or three liquids, such as turpentine, milk, vinegar, beer, wine, sea-water or a solution of salt, and ink.

(*c*) Weigh out about 100 grams of shot. Fill the specific gravity flask with water, and counterpoise it together with the shot. Next put the shot into the bottle, and remove the water displaced. Add weights until the index of the balance swings evenly. The weights added must equal the weight of the water displaced, that is, the weight of a volume of water equal in volume to the shot. Therefore

$$\text{Relative density of the shot} = \frac{\text{weight of shot}}{\text{weight of water displaced}}.$$

(*d*) Find by the method used in the preceding experiment the relative densities of such common things as tin tacks (which are really made of iron), bits of slate pencil, brass wire, and brass nails.

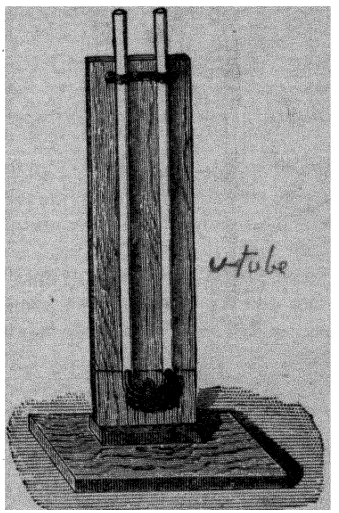


FIG. 39.—An arrangement for balancing columns of liquid. Mercury is in the bend of the tube, up to the line on the upright board.

ii. The relative density of liquids determined by balancing columns.—(*a*) Make a U-tube by bending a piece of glass tubing, or by connecting two pieces of tubing of equal bore with a piece of india-rubber. Mount the tubes upon a strip of board. Pour quicksilver into one of the branches of the U-tube until it reaches a horizontal line drawn on the board (Fig. 39). Now introduce water into one of the tubes, and notice that the mercury on which the water rests is pushed down; afterwards introduce enough water into the other tube to bring the mercury back to its original level. By measuring you find the length of each column is the same. Repeat the experiment with different quantities of water.

(*b*) Remove the water and dry the tubes, and see that the mercury is up to the mark. Nearly fill one of the tubes with some liquid, such as methylated spirit, and balance it with water introduced into the other tube. Measure the lengths of the columns of liquid. Determine the relative density of the liquid, using the equation that

$$\text{Relative density of spirit} = \frac{\text{length of water column}}{\text{length of spirit column}}.$$

iii. Hare's apparatus.—Connect two glass tubes with a three-way tube as shown in Fig. 40, Let the lower ends of the glass dip into different liquids; then apply suction to the short tube at the top, and when the liquids have risen to a convenient height close the clip upon the india-rubber. Measure the height of the liquid in each tube above the level of the liquid below. The heights are inversely proportional to the densities.

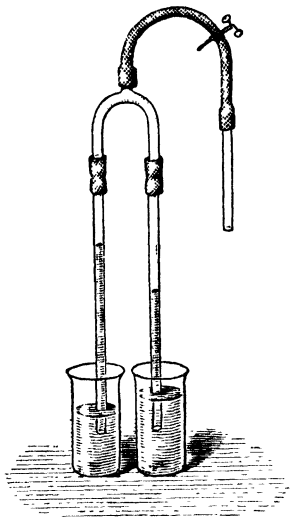


FIG. 40.—Hare's apparatus for comparing the densities of liquids.

Determination of relative densities by relative density bottle.—A simple method of determining the relative densities of substances is to use a specific gravity or relative density bottle. Such a bottle often consists of a small glass flask, holding about fifty grams of water. It is provided with a nicely-fitting ground stopper, with a very small bore through it or a vertical groove cut upon it (Fig. 38). To use the bottle in determining the relative density of liquids and powders, the weight of the empty bottle and stopper must first be known. The bottle is then filled with pure water, the stopper inserted, and the water which is forced through the hole in the stopper wiped off, and the bottle and its contents weighed. In this way the weight of water which just fills the bottle is found. If now we empty the bottle and carefully dry it inside and out, and fill it with the liquid of which the density is required, say spirits of wine, and weigh again, we have the weight of the liquid which just fills the bottle, or the weight of equal volumes of the liquid and water. The ratio—weight of liquid divided by weight of water—gives the relative density of the liquid.

Instead of a bottle of this kind, a flask having a narrow neck around which a horizontal mark has been made may be used. The weight of water which fills the flask up to the mark may thus be compared with the weight of liquid which fills it to the same mark.

Suppose the weight of water in a relative density bottle was found to be 50 grams, and the weight of the same volume of methylated spirit was found to be 40 grams. Then these numbers show the relative densities of the two liquids, and as we take the density of water as the standard or unit, the density of the spirit is equal to 40 divided by 50. It is thus seen that the relative density of spirit is represented by the fraction $\frac{40}{50}$ or $\frac{4}{5}$, which expressed as a decimal fraction is 0·8.

Determination of relative densities by balancing columns of liquid.—A convenient way to balance liquids against one another, and so compare their densities, is by means of a glass tube bent in the form of a **U**.

When a **U**-tube is arranged, as in Fig. 39, the mercury in the bend acts just like a pair of scales, and a column of water in one of the upright arms can be balanced against a column of the same length in the other. The columns of water are the same size, or have the same volume, and they balance one another, consequently their weights must be the same ; and since their weights are equal and their volume the same, they must have the same density.

But suppose water is put in one arm of the **U**-tube, and enough methylated spirit into the other to make the mercury stand at the same height in the two arms. Here there is a different state of affairs. The column of spirit which balances the column of water will be the longer, hence its size or volume is greater, since the tubes are the same width. But because they balance, their weights must be the same ; and the lengths of the spirit and water columns are inversely proportional to the densities of the two liquids. Thus,

$$\text{Relative density of spirit} = \frac{\text{length of water column}}{\text{length of spirit column}}.$$

Hare's apparatus for determination of relative densities.—A convenient arrangement to use in determining the relative densities of liquids which mix is represented in Fig. 40, and is known as Hare's apparatus. Two straight glass tubes are connected at the top by a three-way junction, upon the unconnected end of which a piece of india-rubber tubing is placed. The lower ends of the tubes dip into beakers containing the liquids the relative densities of which have to be determined. By

applying suction to the free end of the rubber tube, the two liquids are drawn up the glass tubes, and the heights of the liquid columns above the level of the liquids in the beakers will be inversely proportional to the relative densities of the liquid employed. The principle is thus precisely the same as that of the U-tube, but by using the form of apparatus here described, the relative densities of liquids which mix can be more conveniently found than by the ordinary U-tube into which the liquids have to be poured.

CHIEF POINTS OF CHAPTER V.

Mass is the amount of matter in a thing. The British standard of mass is the Imperial Pound avoirdupois, and the metric standard in common use is the Kilogram. The metric standard of mass used in scientific work is the *gram*. Masses may be compared by a pair of scales.

The **weight** of an object is the force with which the object is pulled toward the centre of the earth. Weight can be measured by a spring balance.

Density.—Equal volumes of different substances may have different weights. Equal weights of different substances may have different volumes. Or, different substances may have different densities.

Specific gravity, or **relative density**, is the ratio of the weight of any volume of a substance to the weight of an equal volume of water.

$$\text{Relative density} = \frac{\text{mass (or weight) of a substance}}{\text{mass (or weight) of an equal volume of water}}.$$

Balancing columns of liquids.—The densities of two liquids balanced in a U-tube are in the inverse proportion of the lengths of the columns. Or,

$$\text{Relative density of a liquid} = \frac{\text{length of water column}}{\text{length of liquid column}}.$$

Metric Measures of Mass.

1 = **Gram.**

$\frac{1}{1000}$ = Milli-gram. | 10 = Dekagram.
$\frac{1}{100}$ = Centi-gram. | 100 = Hekto-gram.
$\frac{1}{10}$ = Deci-gram. | 1000 = Kilogram.

British and Metric Equivalents.

1 gm. = about 15½ grains. | 1 oz. Avoir. = about 28¾ grams.
1 kgm. = „ 2¼ lbs. | 1 lb. Avoir. = „ $\frac{9}{20}$ kilogram.

The following summary, dealing with standards of weight and measure, is exhibited in the Victoria and Albert Museum, South Kensington:

Standards of Weight and Measure.

The "yard" of 36 inches, is the unit or standard measure of extension, from which all other imperial measures of extension, whether linear, superficial, or solid, are ascertained.

The "pound" is the unit of the measure of weight from which all other imperial weights, or measures having reference to weight, are ascertained. The one seven thousandth part of the pound is a "grain." One sixteenth part of the imperial standard pound (437½ grains) is an ounce (avoirdupois).

The "ounce troy" consists of 480 grains and is now the only legal standard troy weight.

All articles sold by weight must be sold by avoirdupois weight; except that—

(1) Gold and silver, and articles made thereof, including gold and silver thread, lace or fringe, also platinum, diamonds, and other precious metals or stones, may be sold by the ounce troy or by any decimal parts of such ounce—and

(2) Drugs when sold by retail, may be sold by apothecaries' weight.

[The "apothecaries'" ounce is of the same weight as the troy ounce of 480 imperial grains.]

The unit or standard measure of capacity from which all other measures of capacity, as well for liquids as for dry goods, are derived, is the "gallon" containing ten imperial standard pounds weight of distilled water (weighed in air against brass weights) at the temperature of 62° Fahrenheit and with the barometer at 30 inches.

Metric System of Weights and Measures.

The third Schedule to the Weights and Measures Act, 1878, sets forth the equivalents of imperial weights and measures and of the weights and measures therein expressed in terms of the metric system, and Sect. 21 of the Act provides that a contract shall not be invalid if weights or measures of the metric system, or decimal subdivisions of imperial weights and measures, are used in it.

The unit of Length is the "Metre," = 39·3708 inches.
" " Surface " "Are," = 100 sq. metres = 119·6033 sq. yds.
" " Capacity " "Litre," = $\frac{1}{1000}$ cub. metre = 1·76077 pints.
" " Weight " "Gram," = 15·4323487 grains.

(A gram is the weight of a cubic centimetre of distilled water at 4° centigrade.)

EXERCISES ON CHAPTER V.

1. Define the *mass* and *weight* of a material body, carefully distinguishing between the terms.

2. Give the British and metric measures of mass.

3. What is a spring balance and what can be measured with it?

4. What particulars about a given mass could you determine (*a*) by using a spring balance, (*b*) by means of a pair of scales?

5. A body of which the volume is 1·5 litres, has a mass of 25 kilograms. What is the mass in grams of 1 cubic centimetre of the body.

What is the volume at 4° C. of 1,712 grams of water?

6. How is the relative density of a liquid determined with a specific gravity bottle?

7. Explain a simple method for ascertaining the relative density of small shot or tin-tacks.

8. Being provided with two pieces of glass tube and a piece of india-rubber tubing, explain how you would proceed to (i) compare the relative densities of olive oil and spirits of wine, (ii) ascertain the relative density of a specimen of milk.

9. You are given a small rectangular block of brass, and you have at your disposal a measuring rod divided into centimetres and millimetres, a balance and weights, some fine wire, and a vessel of water. How will you determine, by two perfectly independent methods, so that the results may form a check on one another, the volume of the block? Describe exactly what calculations you will have to make, and say upon what scientific principle, if any, each of your methods depends.

CHAPTER VI.

PARALLEL FORCES. CENTRE OF GRAVITY. THE LEVER AND OTHER MACHINES.

18. PARALLEL FORCES.

i. **Example of parallel forces.**—Place the ends of a stiff lath or rod of uniform thickness upon two letter balances, or support the rod by hanging each end from a spring balance. Notice the load borne by each balance; then weigh the rod, and so determine the fraction of the load supported at each end.

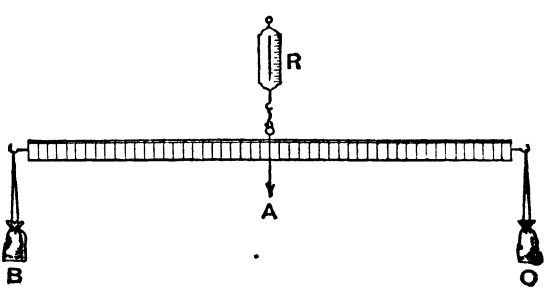


FIG. 41.—The reading of the spring balance is equal to the total weight of the lath and the loads.

ii. **Resultant of parallel forces.**—(*a*) Suspend a light lath from a spring balance by a ring above its centre (Fig. 41). Notice the reading of the balance. Hang two *equal weights* in bags from the ends of the lath, and again observe the reading of the spring balance. Repeat the experiment with *unequal weights* placed in bags arranged

on the lath so as to counterpoise one another. Record your results thus:

Weight of lath *A*.	Weight *B*.	Weight *C*.	Total *A*+*B*+*C*.	Reading of Spring Balance, *R*.

By comparing columns 4 and 5 it will be seen that the three downward forces *A*, *B*, and *C*, acting upon the lath, are kept in equilibrium by one upward force *R*.

(*b*) Hang the two weights *B* and *C* together from the middle of the lath suspended from the spring balance. Notice that the reading of the balance is the same as when the two weights are hung from the ends of the lath.

Parallel forces.—It has been seen that the earth exerts a downward pull upon all objects on its surface, and that in consequence of this all things fall to the ground if unsupported. It follows, therefore, that every mass which is supported above the earth's surface is constantly being pulled downwards, even though it does not fall. If a beam, for instance, is supported horizontally by resting the ends upon two posts, each particle of it may be regarded as being pulled earthwards by an attractive force. The direction of the pull is everywhere towards the centre of the earth, so for any one spot on the earth's surface we may consider the attractive forces due to gravity to be parallel to one another.

When a stiff lath or rod of uniform thickness rests upon two letter balances, or is supported by hanging each end from a spring balance, the experiment represents on a small scale the case of a beam referred to before; and by using spring balances it can be proved that the weight supported at its ends is equally divided between the two supports. In other words, the two upward forces exerted by the balances are together equal to the downward force represented by the weight of the beam.

If a load be placed anywhere upon the lath, the balances still show that when the lath is in equilibrium the sum of the upward forces is equal to the sum of the downward forces.

Principle of parallel forces.—The principle of parallel forces demonstrated by the experiments referred to may now be definitely expressed as follows: **The resultant of a number of parallel forces is numerically equal to the sum of those which act in one direction, less the sum of those which act in the opposite direction.** In other words, the resultant is equal to the *algebraic* sum of the forces.

If two *equal parallel forces* act in the same direction upon a body, the total force will be obtained (as might be expected) by adding the two individual forces together. In like manner, if two *unequal parallel forces* act in opposite directions the net effect can be found by subtracting the smaller of the two forces from the greater; the direction of the resultant is that of the greater force.

19. DETERMINATION OF CENTRES OF GRAVITY.

i. Experimental methods of determining centres of gravity.—(*a*) Procure a disc of sheet cardboard and find by trial the point on which it may be balanced, that is, the centre of gravity of the disc. Make a hole in the card near the edge, and take a plumb-line consisting of a thread with a piece of lead tied at one end and a hook of thin wire at the other. Hang the disc from the hook, and then suspend both as shown in Fig. 42, so that the disc and lead are both suspended and the thread passes over the point of suspension. The thread also passes through the centre of gravity. Do this for various holes in the edge of the disc, and see that in all cases the *vertical line through the point of suspension passes through the centre of gravity.*

(*b*) Another way to find the centre of gravity of a plane figure, such as a triangular plate of metal or card, is to tie strings at each corner of the plate. Then hang the plate by one of the strings to a support such as one of the rings of a retort stand. Allow it to come to rest, and, using a straight edge, draw a chalk mark across the plate in the same straight line with the string as shown by the dotted line in the figure. Now attach the same plate by one of the other threads exactly as before, and again make a mark in continuation of the string. The two chalk marks intersect at a point marked *G*, Fig. 43. Untie and do the same with another string, the third line passes through the intersection of the first two. Obtain a similar point for each of the other plates. Also determine in the same way a similar point for irregular plates of wood, zinc, or cardboard.

ii. Centres of gravity of skeleton solids.—(*a*) Procure a skeleton cube or tetrahedron, and suspend it as in the preceding experiments. Mark the verticals through the point of suspension by light wires attached by wax, and thus find the position of the centre of gravity.

The centre of gravity will not be on any of the bars of the skeleton solid used.

(*b*) Find the centre of gravity of an open wickerwork basket, such as a waste-paper basket. To do this, suspend the basket, and hang

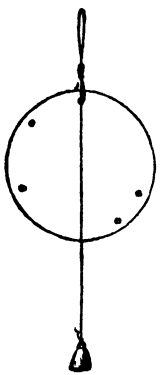


FIG 42.—Determination of centre of gravity of a disc.

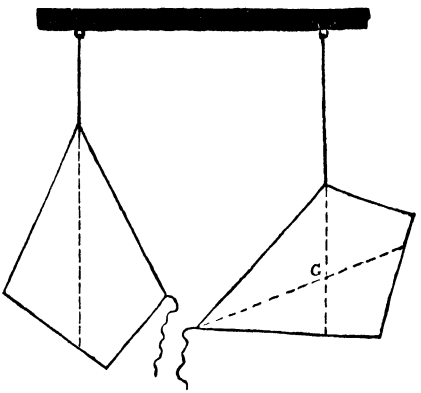


FIG. 43.—A method of determining centre of gravity.

a plumb-line from the point of suspension. Tie a piece of thread across the basket in the direction of the plumb-line; then suspend the basket from another point, and notice where the plumb-line crosses the thread. The point of intersection, which need not be actually on the framework itself, is the centre of gravity.

Centre of gravity.—Consider a large number of weights, some heavier than others, suspended from a horizontal rod. A certain position can be found at which a spring balance would have to be attached in order to keep the rod in equilibrium. When the rod is hung from this point the tendency to turn in one direction is counteracted by the tendency to turn in the other, so the rod remains horizontal. The weights may be regarded as parallel forces, and the pull of the spring balance as equal to their resultant. Now consider a stone, or any other object, suspended by a string. Every particle of the stone is being pulled downwards by the force of gravity, as indicated in Fig. 44. The resultant of these parallel

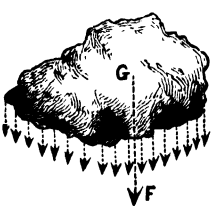


FIG. 44.—Parallel forces due to gravity.

forces is represented by the line GF, and the centre of the forces is the point G. The point G, through which the resultant (GF) of the parallel forces due to the weights of the individual particles of the stone always passes, is known as the *centre of gravity*. For the stone to be in equilibrium, the string must be attached to a point in the line GF, produced upwards.

Every material object has a centre of gravity, and the position of this point for a particular object is the same so long as the object retains the same form.

Experimental methods of determining centres of gravity.—The centre of gravity of such geometrical figures as circles, squares, and parallelograms is really the centre of the figures, and can therefore be determined geometrically. In the case of unsymmetrical figures, however, the centre of gravity cannot be so easily found by geometry, and is best determined by experiment.

The experimental method adopted for determining the centre of gravity of any material body depends upon the considerations set forth in the preceding paragraphs. The body, the centre of gravity of which is required, is allowed to hang quite freely, either by means of a cord or on a smooth peg, and when it has come to rest a vertical line through the point of support is marked upon it. If a string is employed, this vertical line will be a continuation of the string, and is at once drawn by the help of a ruler. If the body which is being experimented with is hung from a smooth nail, by means of a hole bored in the body, the vertical line must be drawn with the help of a plumb-line. The point of support is then shifted and the operation repeated. Since the centre of gravity of the plate is in both straight lines it must be located at their intersection.

Plates of all shapes balance about their centres of gravity.—After the centre of gravity of a sheet of metal, or other stiff material, has been determined by hanging it from a support in the manner described in Experiment 19 i. (a) and (b), it will be found that if this sheet be so arranged that a pointed upright is immediately under the centre of gravity, the plate will be supported in a horizontal position. This affords a convenient means of checking the correctness of the experiment performed.

Geometrical determination of centres of gravity.—It has been sufficiently explained that the centres of gravity of straight

lines, circles, squares, and other regular figures is at their geometrical centres. Hence, the geometrical constructions for determining these central points also locate the positions of their centres of gravity.

The centre of gravity of a **parallelogram** is at the intersection of its diagonals.

The centre of gravity of a **triangle** is determined by bisecting any two sides and joining the middle points so obtained to the opposite angles. The intersection of the lines so drawn gives the centre of gravity. The centre of gravity is found, by measuring, to be one-third the whole length of the line drawn from the middle point of the side to the opposite angle, away from the side bisected.

We may, in fact, consider a triangular plate as made up of a number of narrow strips of material which decrease in length

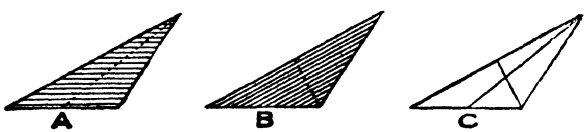


FIG. 45.—Geometrical illustration of centre of gravity of a triangular plate.

from the base to the apex. The centre of gravity of each strip is the middle of the strip; hence the line drawn from the apex to the middle of the base passes through each centre of gravity (Fig. 45, *A*). By taking another side as base, a similar line can be drawn from the middle to the opposite angle (Fig. 45, *B*). These lines intersect at one-third the distance up the line so drawn, measured from the base, and the point of intersection is the centre of gravity of the triangular plate (Fig. 45, *C*).

To find the centre of gravity of a quadrilateral by construction, the plan is to divide it into two triangles by drawing a diagonal. By the method just described the centre of gravity of each triangle is found, and the points so obtained are joined. The centre of gravity of the quadrilateral lies on this line. Repeat the process drawing the other diagonal. Join the centres of gravity of the second pair of triangles, the centre of gravity of the quadrilateral lies on this line. Hence, it is situated at the point of intersection of this line and the first one obtained in the same way.

Centres of gravity of other bodies.—The method of drawing lines across the surface of a thin plate is not practicable in the case of bodies such as blocks or skeleton solids having three dimensions. The experiments described in 19 ii. explain two methods which can be employed usefully in cases of this kind. The ingenuity of the student will provide other equally suitable plans for particular cases.

20. EQUILIBRIUM.

i. Conditions of equilibrium.—Place upon a square-edged table or board one of the cardboard figures of which you have found the centre of gravity. Gradually slide the figure near the edge until it would just topple over; keeping it in this position, draw a line

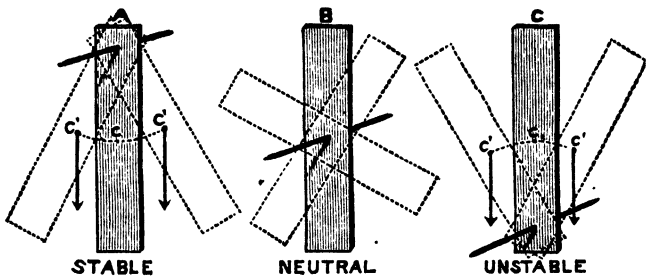


FIG. 46.—Relative positions of centre of gravity and point of support for stable, neutral, and unstable equilibrium.

along the under side of the cardboard where the edge of the table touches it. Then place the cardboard in another position and again mark where the edge of the table touches it when it would just topple over. The intersection of these lines is the centre of gravity, and it will be noticed that the cardboard would just topple over when the centre of gravity falls outside the edge of the table.

ii. Stable, neutral, and unstable equilibrium.—Procure an oblong strip of wood or cardboard (Fig. 46). Support the strip as at *A* by a long pin pushed through it; it is then in stable equilibrium, for the slightest turn either to right or left raises the centre of gravity. When supported as at *B*, the strip is in neutral equilibrium; and when supported as at *C*, it is in unstable equilibrium, for the slightest movement lowers the centre of gravity.

Relation of centre of gravity to base of support.—A circular disc, in which it will be remembered the centre of

gravity coincides with the geometrical centre, will not rest upon a table if the centre is beyond the edge of the table, but will topple over. In a similar way, if any plane figure lies flat upon a table, the centre of gravity of the figure must be within the edge of the table. The same conditions apply to any object resting upon a support. For an object resting upon a base to be in equilibrium, a vertical line drawn from the centre of gravity downward must fall within the base. When this vertical line falls outside the base the body topples over.

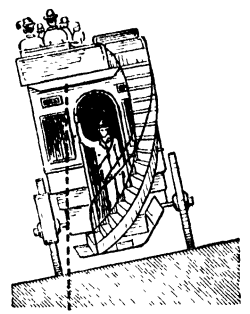


FIG 47 —If the vertical line from the centre of gravity should fall outside the base of support, the omnibus would topple over.

Consider the case of an omnibus on level ground. The centre of gravity is somewhere inside the omnibus, and a vertical line drawn from it downwards would fall within a line traced around the omnibus upon the ground. But, if the outside of the omnibus is filled with people and the vehicle happens to be running across a sloping road, it might topple over, for a jerk might cause so great a change of position of the centre of gravity as to make the vertical line from the centre fall outside the base of support, and in such a case an accident must happen (Fig. 47).

Equilibrium.—When a body is at rest, all the forces acting upon it balance one another (or, what is the same thing, any force is equal and opposite to the resultant of the remaining forces), and it is said to be in equilibrium. It is in **stable equilibrium** when any turning motion to which it is subjected raises the centre of gravity; in **unstable equilibrium** when a similar movement lowers the centre of gravity, and in **neutral equilibrium** when the height of the centre of gravity is unaffected by such movement. Consequently, if a body in stable equilibrium is disturbed, it returns to its original position; if in unstable equilibrium, it will, if disturbed, fall away from its original position: while if the condition of equilibrium is neutral it will, in similar circumstances, stay where it is placed.

Conditions of stability of suspended and resting objects.—The centre of gravity must in every case be below the point of support for a suspended object to be in equilibrium. The greater the distance between the point of support and the centre of gravity the greater is the tendency to return to the position of equilibrium.

When the centre of gravity and the point of support of a suspended object are close together the equilibrium of the object is easily disturbed. A good balance partly owes its sensitiveness to this condition, the centre of gravity and point of support being designedly brought close together.

It has been shown that in the case of a freely suspended object the centre of gravity is at its lowest point when the object is in equilibrium. Let us see how this applies to a body supported upon a surface below the centre of gravity.

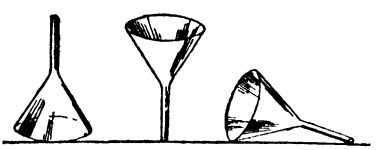


FIG. 48.—A funnel in (*A*) stable equilibrium, (*B*) unstable equilibrium, (*C*) neutral equilibrium.

A body is least liable to be upset when the centre of gravity is at a considerable distance from all parts of the edge of the base; for, when this is the case, the body has to be tilted through a large arc before the centre of gravity falls outside the base.

A funnel standing upon its mouth is an example of a body which cannot be easily overturned on account of the low centre of gravity and its distance from the edge of the base (Fig. 48, *A*). It is then in stable equilibrium. If the funnel is stood upon the end of the neck it can be easily overturned, because very little movement is required to bring the centre of gravity outside the base. It is thus in unstable equilibrium. When the funnel lies upon the table it is in neutral equilibrium, for its centre of gravity cannot then get outside the points of support.

21. THE LEVER.

1. Balancing equal weights on a lever.—(*a*) Make or obtain a lever consisting of a strip of light wood graduated in centimetres, and having a thin ring screwed into one edge, above the central point,

and a hook screwed into each end (Fig. 49). Hang the lever from a nail by the middle ring. If it does not exactly balance, plane off a little wood from the end which sinks, or slightly unscrew the hook at the end which rises above the horizontal, until the lever does set itself horizontally.

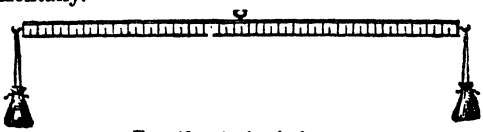


FIG. 49.—A simple lever.

Hang two pill-boxes or small linen bags by their strings from the lever, one on each side of the fulcrum or pivot, at equal distances from it. Into one of the bags place say 50 grams, and find how many grams you must place into the other bag in order to restore equilibrium. Repeat the experiment with the bags at a different distance.

It will be found always that equilibrium is obtained when equal weights are at equal distances from the turning point or fulcrum.

ii. Principle of moments.—(*a*) Place some pieces of lead in each of four linen bags, and adjust by means of shot or small bits of lead until they weigh respectively 50 grams, 100 grams, 200 grams, and 300 grams. Place the 100 gram bag about 12 cm. from the fulcrum of the lever, and balance it with a 50 gram bag on the other side. Record the distance from the fulcrum in each case. Repeat the experiment by balancing 50 grams against 100 grams, 50 grams against 200 grams, 100 grams against 300 grams and other combinations.

Record your observations in columns as below:

LEFT SIDE OF LEVER.			RIGHT SIDE OF LEVER.		
Load.	Distance from Fulcrum.	Load × Distance.	Load.	Distance from Fulcrum.	Load × Distance.

Compare the numbers in columns 3 and 6, and state in words the law or rule indicated by the results.

The experiments show that there is a definite proportion between the weights on the two sides of a lever, and their distances from the fulcrum. The proportion is:

$$\text{left load} : \text{right load} :: \text{right distance} : \text{left distance}.$$

Or, expressed in another way, the loads are inversely proportional to their distances from the fulcrum.

The turning effect of any force acting upon a lever, as each load did in the experiment, is termed the *moment* of the force. The comparisons of columns 3 and 6 prove that the *moment* or turning effect is measured by the product of load into distance from the fulcrum.

(*b*) Hang two bags on the same side of the lever at different distances and one bag on the other side. Move the single bag until equilibrium is obtained. Do this several times with the bags in different positions, and compare the sum of the moments of the forces acting on one side with the moment of the force on the other side.

(*c*) Hang a small book or a bag with shot or nails in it on one side of the fulcrum, and the 100 gram bag from the other. Move this bag along the lever until equilibrium is obtained. Then, remembering that

$$\text{load}_1 \times \text{distance}_1 \text{ from fulcrum} = \text{load}_2 \times \text{distance}_2 \text{ from fulcrum},$$

calculate the weight of the book or bag of nails. Repeat the experiment, using the 200 gram bag.

In the preceding experiments the fulcrum has been between the forces due to the loads hanging from the lever. The forces may, however, both act on one side of the fulcrum. It is convenient to call one of the forces the effort and the other the resistance.

iii. **Load between effort and fulcrum.**—Suspend the lever from its middle hook as before. Attach a spring balance near one end, and suspend a load from some point between the balance and the fulcrum. The principle of moments applies to this case as to the others, and you will find that

$$\text{reading of balance} \times \text{distance of balance from fulcrum} = \text{load} \times \text{distance from fulcrum}.$$

The forces acting in this experiment are the same as those called into play when a man lifts the handles of a wheelbarrow having a load in it.

iv. **Effort between load and fulcrum**—Suspend a load from one end of the lever supported upon the central ring. Attach a spring balance to the lever between the load and the fulcrum, and hold the lever horizontal by means of it. Keeping the balance between the load and the fulcrum, show that whatever the relative distances of the two from the fulcrum, the principle of moments holds good when the lath is kept horizontal. This class of lever is similar, as regards the distribution of forces, to sugar-tongs and ordinary fire-tongs.

Machines.—A machine is a contrivance by means of which a given force is made to resist or overcome another force acting in a contrary direction.

What are termed "the mechanical powers" are really simple machines which can be used to overcome resistance We have now to consider the mechanical principles underlying the action of such a simple machine as the lever.

The lever.—A lever is a rigid bar which can be turned freely about a fixed point. The **fulcrum** of a lever is the fixed point about which the lever can be turned. The force exerted when using a lever is often described as the *power* and the body lifted or resistance overcome as the *weight*. These words are convenient, but they are not correctly used in connection with levers, as their true meanings are confused by so doing. It is better to substitute the word *effort* for power, and *resistance* or *load* for weight. It should be borne in mind, that, so far as mechanical principles are concerned, there is no difference between the power and the weight; both represent forces, and as such they must be considered in the action of levers.

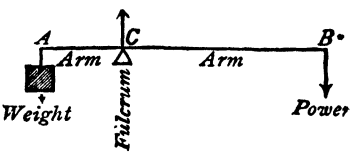


FIG. 50.—Terms used in connection with levers.

The perpendicular distances from the fulcrum to the lines of actions of forces acting upon a lever, are known as the **arms of the lever.** In Fig. 50 the distance AC is the arm at one end of which the "weight" acts, and BC is the arm at one end of which the "power" acts.

Classes of levers.—For convenience, levers are divided into three orders or classes, according to the relative positions of the fulcrum and the forces in action. The classes are as follows:

Class I. Fulcrum between resistance and effort. Examples: see-saw, a pump-handle, a balance, a spade used in digging.

Class II. Resistance between effort and fulcrum. Examples: nutcrackers and a wheelbarrow.

Class III. Effort between fulcrum and resistance. Examples: sugar-tongs, ordinary fire-tongs, and the pedal of a grindstone.

These classes are of no real consequence, for the principle underlying the action of all levers is the same.

It will, however, be a valuable and interesting exercise to follow out these distinctions in the Figs. on p. 81.

Principle of the lever.—It is easy to show by experiment, that when a lever is in equilibrium the following equation holds good:

$$\text{force on one side} \times \text{distance from fulcrum} = \text{force on other side} \times \text{distance from fulcrum.}$$

FIG. 51.—Levers of the first order. FIG. 52.—Levers of the second order. FIG. 53.—Levers of the third order.

This principle of moments applies to all levers, so that all that need be remembered when considering the action of a lever of any kind, are the forces working in one direction and their distances from the fulcrum, compared with the forces or resistances which oppose them and the distances of these from the fulcrum.

Moments.—Refer to the diagram (Fig. 54), where F represents the point of support, or fulcrum, of a lath or other straight lever, and M_1 is a weight at a distance AF in equilibrium with a smaller weight M_2, at a greater distance FB.

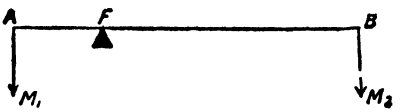


FIG. 54.—To illustrate moments of forces.

The force acting at A is the weight of M_1, acting vertically downwards; and the force at B is the weight of M_2, acting in the same direction. Each force tends to turn the lever in a particular direction, and this turning effect is called the **moment of the force.** The moment of the force acting vertically downwards at A is the product of the force equal to the weight of M_1 into the distance AF, which, as the diagram shows, is measured at right angles to the direction in which the force acts. Similarly, the moment of the force equal to the weight of M_2, about the point F, is equal to the product of this force and the perpendicular distance BF.

This is a rule of universal application for taking moments which will have to be used several times later on, and should be well borne in mind.

The moment of a force about any point is the product obtained by multiplying the force by the perpendicular distance between the point and the line of action of the force.

The principle of work.—Though the subjects of work and energy cannot be considered fully in this book, it may be stated here that work is measured by the product of a force into the distance through which a body is moved in the direction of the force. Now, in every case where such a simple machine as the lever is concerned, the work done by one set of forces is equal to that done by the other. What is gained in force is lost in distance, so that in each case the product of force and distance is the same, whether the lengths of the lever arms, or

the distances through which the ends of the lever move, are taken as the distances.

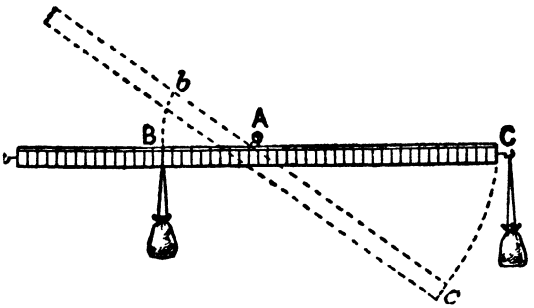


FIG. 55.—Principle of work illustrated by a lever.

From the principle of work it follows that if a man, by exerting a force of 10 lbs. on one end of a crowbar, moves 100 lbs. at the other end, he has to exert his effort through 10 inches in order to move the mass one inch. Thus, what is gained in effort has to be made up by distance moved.

22. THE PULLEY, INCLINED PLANE, SCREW, WHEEL AND AXLE.

i. The pulley.—Attach a spring balance to one end of a piece of string. Make a loop in the other end and hang from it a weight and a pulley. The spring balance, if correct, indicates the weight of the two together. Pass the string under the pulley as in Fig. 56, *Bb*. The total weight is now supported by two parallel portions

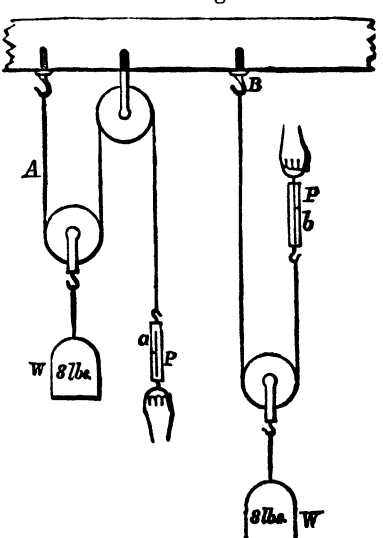


FIG. 56.—Actions of pulleys.

of the string, and the tension of the string, as indicated by the spring balance, is only one-half what it was before. Pass the string over a fixed pulley as shown by *Aa* in the left-hand part of Fig. 56. The tension will be found practically the same as before.

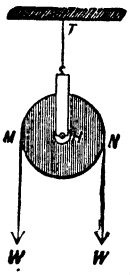


FIG. 57.—A single fixed pulley.

ii. Single fixed pulley.—Fix a pulley as in Fig. 57. Pass a flexible cord over the groove, and hang a weight on each end of it. The weights have to be equal in order to balance one another.

iii. Single movable pulley.—Arrange one fixed and one movable pulley as in Fig. 58. Hang a weight *W* from the pulley, and find the weight *P* required to balance it. Repeat the experiment with several different weights. In each case the value of *P* is found to be about one half the sum of *W* and the weight of the pulley to which it is attached.

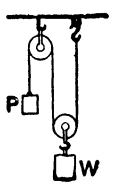


FIG. 58.—A fixed and a movable pulley.

iv. Inclined plane.—(*a*) Arrange a hinged board with a weight attached by elastic to the free end. Show that the tension is less when the weight rests on the board than when it is suspended freely.

(*b*) Procure a toy carriage such as is shown in Fig. 59. Place some shot in it and find the total weight *W* of the carriage and shot. Attach one end of a string to the carriage after passing it over a pulley or smooth rod *B*, and fasten the other end to a small bucket or box *P*. Put sufficient shot in the box to balance the carriage. Find by several experiments the relation between *W* and *P*.

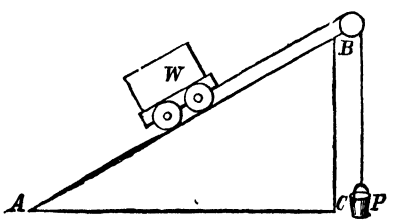


FIG. 59.—Experiment to illustrate the advantage of the inclined plane.

v. The Screw.—Cut out of paper a right-angled triangle such as *ABC* (Fig. 60), and wind it around a lead pencil. The slant side of the triangle forms a spiral upon the pencil, similar in appearance to the thread of a screw. Mark off from the point *C*

a part of the base CD which will just go around the pencil, and draw the perpendicular DE. The small triangle CDE is similar to the large one and represents one turn of the screw-thread.

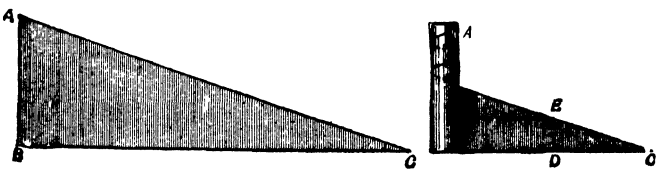


FIG. 60.—To illustrate the principle of the screw.

The pulley.—A pulley is a wheel having a grooved rim, and capable of rotating about an axis through its centre. The frame which holds the pulley is called the block.

Experiment 22 iii., shows that although a movable pulley, that is, one which can move up and down, reduces the effort which has to be exerted to support a given weight, a fixed pulley is of no advantage in that respect. The ratio of the resistance, or weight overcome to the effort exerted is known as the **mechanical advantage** of a machine.

Experiment 22 ii., shows that, with a single **fixed pulley,** no mechanical advantage is obtained. All that this pulley does is **to change the direction of the pull** ; if one of the weights, for instance, is pulled down, the other rises. The pulley thus acts in the same way as a lever balanced at its centre ; the distance from the centre to the circumference, in other words, the radius of the pulley, being regarded as one arm of the lever. A pulley having a radius of three inches has therefore an equivalent lever-arm three times as great as one with a radius of one inch.

Single movable pulley.—Advantage is gained by the use of **movable** pulleys. The fixed pulley is of no advantage in reducing the force required to raise a weight ; the advantage gained is derived from the use of the movable pulley. One half of the total weight is supported by the part of the string hooked to the beam, and the other half is supported by the part of the string which goes around the fixed pulley to the weight marked P. There are several different combinations of pulleys, but the principle exemplified by the foregoing experiments, namely, that every movable pulley reduces by one-half the effort required to raise a given weight, is utilised in them all.

The principle of work applied to pulleys.—With pulleys, as with levers, there is neither loss nor gain of work. If, in any combination of pulleys, a force of 10 lbs. balances a force of 120 lbs.—the mechanical advantage (resistance÷effort) thus being 12—the effort will have to be exerted through twelve feet in order to move the resistance through one foot. For it is an invariable rule that

$$\text{Effort} \times \begin{cases} \text{the distance it} \\ \text{acts through} \end{cases} = \text{Resistance} \times \text{distance moved.}$$

The mechanical advantage of any system of pulleys can therefore be determined by finding the relation between the weight moved and the effort exerted. By comparing the

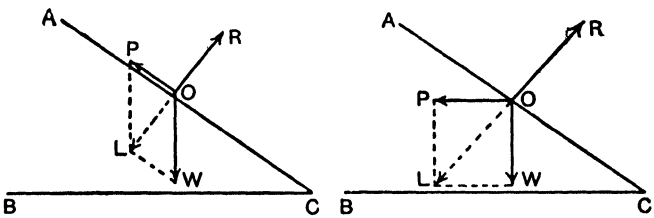


FIG. 61.—Parallelogram of forces applied to the inclined plane.

distance through which the effort is exerted with that through which the weight moves, we obtain the **velocity ratio** of a machine.

If there were no waste of energy in a machine, the mechanical advantage would be equal to the velocity ratio; but this condition is impossible to attain in actual practice.

The inclined plane.—A plane in mechanics is a rigid flat surface, and an inclined plane is one that makes an angle with the horizon.

An object is prevented from sliding down an inclined plane by a force less than the weight of the body. The force required may be measured by a spring balance.

The reason for the decrease of tension in an elastic cord attached to a mass resting on an inclined board, compared with the tension when the mass hangs freely, will be understood from the parallelogram of forces. Suppose an object *O* (Fig. 61) is kept in position upon a smooth inclined plane by a force acting up the plane. The object is acted upon by three forces, namely, *W*, due to its weight, acting vertically downwards, *P*, the force exerted up the plane, *R* the reaction of the plane.

The weight acting downwards is kept in equilibrium by the forces R and P acting upwards. The last-named force represents the effort exerted, and being one of two it is less than the weight whenever the object rests upon the plane. Let OP and OW be drawn of lengths respectively proportional to the force exerted up the plane, and the weight of the object O. Complete the parallelogram $PLWO$, and draw the diagonal OL. This parallelogram represents graphically in magnitude and direction the forces in equilibrium when a body of definite weight is kept in position upon a smooth inclined plane, by a force acting up the plane. If the sustaining force acts horizontally, the parallelogram shown on the right in Fig. 61 represents the forces concerned.

If we measure the force required to draw a weight up an inclined plane, as in Experiment 22 iv., the results obtained show that the weight of the bucket and shot is less than that of the carriage and shot, and that the proportion which one bears to the other differs with different inclinations of the plane. There is a definite relation between this proportion and the slope of the plane on which the carriage travels. When the effort is exerted parallel to the plane, this proportion is as follows:

$$\text{Weight moved } (W) : \text{Effort exerted } (P) :: \text{Length of plane } (AB) : \text{Height of plane } (BC).$$

This rule can also be deduced from the principle of work. If the carriage starts from A and moves to B, it is lifted through the vertical height BC. For this to take place, the effort will have to be exerted through a distance equal to the length of the plane AB. Therefore

$$\frac{\text{Weight}}{\text{Effort}} = \frac{\text{Distance through which the effort is exerted}}{\text{Vertical distance through which weight is lifted}} = \frac{\text{Length of plane}}{\text{Height of plane}}.$$

When the effort acts horizontally, the ratio which it bears to the weight is in the proportion which the base of the plane AC bears to the height BC.

A wedge can be considered to be two inclined planes base to base, this double plane being forced forward by an effort exerted parallel to the direction of the base.

The screw.—A screw may be regarded as an inclined plane wound around a cylinder. This is easily seen in Fig. 60.

Comparing the screw-construction now with an inclined plane, it will be seen that the

height of inclined plane represents distance between threads,
base of „ „ „ circumference of screw.

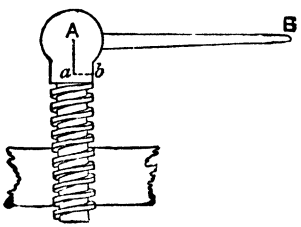


FIG 62 - A screw turned by a lever

The angle of inclination of the inclined plane is represented by the angle *ECD*, and this determines the **pitch** of the screw.

In considering the use of a screw, the resistance to be overcome can be regarded as a weight upon an inclined plane. With a screw such as is shown in Fig. 62 the effort acts in a direction parallel to the base of the plane and under this condition

Weight : Effort :: Base of Plane : Height of plane.

Or expressing the proportion in the terms which apply to screws :

$$\text{Resistance} : \text{Effort} :: \begin{matrix}\text{Circumference} \\ \text{of screw}\end{matrix} : \begin{matrix}\text{Distance between} \\ \text{successive threads.}\end{matrix}$$

When the force is applied at *B*, leverage is gained in the proportion of *AB* to *ab*, and so further mechanical advantage is obtained on this account. But, in order to advance the screw by a distance equal to that between two successive threads, the end of the handle *B* has to be turned through a complete circumference. This fact can be used to deduce the mechanica advantage of a screw from the principle of work. We get, in fact,

$$\frac{\text{Resistance}}{\text{Effort}} = \frac{\text{Circumference of circle described by power arm}}{\text{Distance between successive threads}}.$$

The wheel and axle is a well-known contrivance used in raising water from a well. The resistance overcome is the weight of the bucket of water and the effort exerted is the force applied at the handle. The weight of the bucket acts

through the rope which coils around the axle and the effort acts at the circumference of the circle described by the handle. We may think of the machine therefore as consisting of two cylinders, as indicated in Fig. 63.

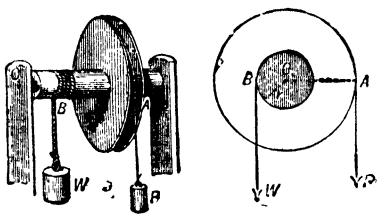


FIG. 63.—The wheel and axle

The lever-arm of the effort P is AC and that of the resistance W is BC. The proportion between these lengths is therefore that between P and W, or,

$$\frac{W}{P} = \frac{CA}{CB}$$

CHIEF POINTS OF CHAPTER VI.

Parallel forces.—The resultant of a number of parallel forces is numerically equal to the sum of those which act in one direction, less the sum of those which act in the other direction.

Moreover, not only is the magnitude of the resultant of two parallel forces equal to the algebraical sum of the components, but when there is equilibrium, one force multiplied by its distance from the point of action of the resultant, is equal to the other force multiplied by its distance from the resultant.

Centre of gravity.—The point through which the resultant of the parallel forces, due to the weight of the individual particles of any mass, always passes, is known as the centre of gravity of the mass.

The centre of gravity of a body may be determined experimentally by allowing the body to hang quite freely, and when it has come to rest tracing a vertical line through the point of support by means of a straight edge. By repeating the process for a second point of support two such lines are obtained, the intersection of which is the centre of gravity.

Plates of all kinds balance about their centres of gravity.

Positions of centres of gravity.—(*a*) Those of straight lines, circles, squares, and other regular figures, are at their geometrical centres.

(*b*) That of a parallelogram is at the intersection of its diagonals.

(*c*) That of a triangle is on the line drawn from one of its angles to the middle point of its opposite side, and at a distance of one-third of this line's length from that side of the triangle.

Equilibrium.—In a body in equilibrium all forces acting upon it balance one another.

It is in *stable equilibrium* when any turning motion to which it is subjected raises its centre of gravity.

It is in *unstable equilibrium* when any turning motion lowers its centre of gravity.

It is in *neutral equilibrium* when the height of the centre of gravity is unaffected by such movement.

A machine is a contrivance by means of which a given force is made to resist or overcome another force acting in a contrary direction.

A lever is a rigid bar which can be turned freely about a fixed point (the *fulcrum*). The force exerted when using a lever is called the *effort*, and the body lifted, or force overcome, the *resistance*.

Classes of levers.—1. Fulcrum is between resistance and effort. Examples: see-saw, pump-handle, balance, crowbar.

2. Resistance is between effort and fulcrum. Examples: nut-crackers, wheelbarrow, and boat oar.

3. Effort between resistance and fulcrum. Example: pair of tongs.

Principle of lever.—

$$\text{Force on one side} \times \text{distance from fulcrum} = \text{force on other side} \times \text{distance from fulcrum.}$$

Moments.—The turning effect of a force is called the *moment of the force*. The moment of a force about any point is the product obtained by multiplying the force by the perpendicular distance between the point and the line of action of the force.

The mechanical advantage of a machine is the ratio of the resistance or weight overcome to the effort exerted.

The pulley.—With one fixed pulley no advantage is gained; but every movable pulley used reduces by one-half the effort required to support or raise a given weight.

The inclined plane.—An object is prevented from sliding down a plane by a force less than the weight of the object. If the force supporting the body acts parallel to the plane, then the weight of the body is to the effort exerted as the length of the plane is to the height. When the force acts horizontally the ratio between the weight and the effort is that of the base to the height.

The screw.—The principle of the screw is obtained from that of the inclined plane. The resistance is to the effort as the circumference of the screw is to the distance between successive threads.

The wheel and axle.—This is the ordinary arrangement for drawing

water from wells. The mechanical advantage of the wheel and axle is obtained from the ratio between the diameters of the wheel and of the axle.

EXERCISES ON CHAPTER VI.

1. Describe the principle of the action of a simple lever.

A stiff wooden rod, six feet long, and so light that its weight may be neglected, lies upon a table with one end projecting four feet over the edge. Upon the end of the rod lying on the table a weight of 8 lbs. is placed. What weight must be placed upon the other end so as just to tip the rod ?

2. What is a lever ? What is the " fulcrum " of a lever ?

Name four or five levers in common use, and say where the fulcrum of each may be.

3. What is meant by the resultant of two forces ?

Describe an experiment to prove that the resultant of two parallel forces is equal to the algebraic sum of the forces.

4. How would you determine the centre of gravity of an iron hoop made by joining together two semicircles, one thicker than the other ? Explain how the observations could be used to find out which was the thicker half of the hoop.

5. How would you determine experimentally the centre of gravity of a sheet of cardboard of irregular shape ?

6. When is a body said to be in equilibrium ? Distinguish between stable, unstable, and neutral equilibrium. What is the condition which determines the nature of the equilibrium ?

7. A solid hemisphere made of uniform material is placed with any part of its curved surface upon a horizontal plane. Show that, however thus placed, it will always tend to a position of stable equilibrium with its flat surface horizontal and uppermost. What other positions of equilibrium are there ? Which of them are stable and which unstable ?

8. Being given an object like a sugar-bowl, how would you propose to find experimentally the position of its centre of gravity ?

9. A square sheet of cardboard weighing 8 ozs. is suspended by a thread fastened to one corner, and a weight of 4 ozs. is fastened to one of the corners adjacent to the corner of suspension. Draw a diagram to show the position in which the sheet will hang, and say what is the total weight that the thread supports.

10. What is meant by the mechanical advantage of a machine ?

11. Describe an advantageous method of using two pulleys, one being fixed.

12. Find the force necessary to support a weight of 10 lbs. on an inclined plane, angle 30°, (*a*) when the force is horizontal, (*b*) when the force is vertical.

13. Indicate clearly how to calculate the mechanical advantage of the screw.

14. Illustrate the 'principle of work' in the arrangement known as the Wheel and Axle.

CHAPTER VII.

PRINCIPLE OF ARCHIMEDES.

23. FLUID DISPLACEMENT AND FLOATING BODIES.

i. Some things sink others float in water.—Fill a fish-globe or finger-bowl with water, and carefully place lumps of different things, *e.g.* pieces of lead, iron, oak, pine, and cork, one after another, into the water. Observe that (1) some sink and others float, (2) of those which float some sink further into the water than others. Take the objects which sink in water and place them in mercury. Notice that they float. (Fig. 64).

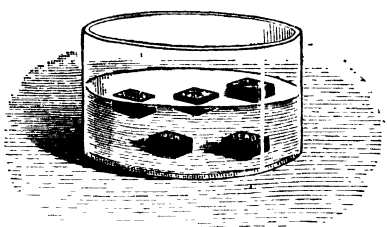


FIG. 64.—When substances are denser than water they sink in water, and if less dense, they float.

ii. Volume of water displaced by solids which float.—(*a*) Obtain a rectangular rod of wood, 1 square cm. in section and about 15 cm. long, with lines around it 1 cm. apart. Gouge a small piece of the wood out of one end, and put lead into the hole; flatten the end by filling in with wax.

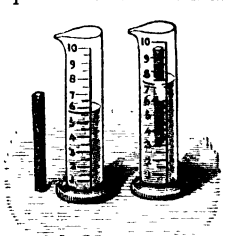


FIG. 65.—The number of cubic centimetres in the part of the rod under water is equal to the number of cubic centimetres of the water displaced.

Put some water in the graduated jar and notice its level. Find the weight of the rectangular rod, and then place it in the jar with the leaded end downwards. Notice how many cubic centimetres of the rod are immersed, and also how many cubic centimetres of water are displaced (Fig. 65). Since the weight of 1 cub. cm. of water is 1 gram, the number of cubic centimetres of water displaced is also the weight in grams of the

water displaced. This weight will be found equal to the weight of the whole rod.

(*b*) Fill the divided glass cylinder with water up to a certain mark. Notice the level of the water.

Draw lines at equal distances apart across a narrow strip of paper and fix the paper inside a test-tube, as in Fig. 66. Float the test-tube in water in the graduated jar and put mercury or shot into it until a certain mark upon the strip of paper inside it is on a level with the surface of the water. Notice the number of cubic centimetres of water displaced when the test-tube is thus immersed.

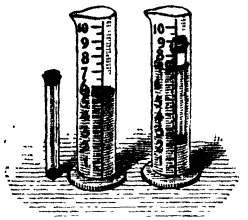


FIG. 66.—The weight of the test-tube and contents is equal to the weight of water displaced.

Then take out the test-tube, dry it, and weigh it together with the mercury it contains. The total weight of the test-tube and contents will be found equal to the weight shown by the number of cubic centimetres of water displaced. Repeat the experiment with the test-tube immersed to a different mark.

Float the test-tube and mercury in spirits of wine and milk in succession. Notice that in the former case it sinks deeper than the mark, while in the other not so deep.

(*c*) Place the loaded test-tube or a hydrometer (1) in milk, (2) in water, (3) in a mixture of milk and water. Observe the depth to which it sinks in each case.

Water displaced by solids.—If a solid one cubic centimetre in size sinks in water it pushes aside one cubic centimetre of water to make room for itself. If its size is two cubic centimetres, it makes two cubic centimetres of water rise above the level the water had at first. Whatever the size of the solid it must have room, and this room is obtained by displacing an amount of water of exactly the same size.

Floating bodies.—A solid which sinks in water or any liquid, displaces a volume of liquid equal to its own volume. When a solid floats, the case is slightly different. Part of the solid is in water and part out of the water, and, of course, only the part immersed is pushing the water aside in order to make room for itself. **In the case of a floating object, therefore, the volume of liquid displaced is equal to the volume of the part of the solid below the surface.**

When any object is floating in water, a certain volume of it is under water, and a certain volume is above the surface. The depth at which it floats depends upon its density. A rod

of heavy wood sinks deeper in water than a rod of light wood of the same size. The water displaced by the heavy wood has therefore a greater volume, and consequently a greater weight, than that displaced by the light wood. But there is one important fact which applies to both cases, and should be kept well in mind. It is that **the weight of the water displaced by the immersed part of a floating object is equal to the whole weight of the object.** If, therefore, the question is asked, how far does an object which floats sink into water, the answer is—it goes on sinking until it has displaced an amount of water having a weight equal to the whole weight of the floating object.

Since the depth at which an object floats in water is decided by this rule, we have a ready way of deciding whether an object will sink further in another liquid or not so far. If the liquid into which it is put is less dense than water, like spirits of wine, it is clear that to make up a given weight we shall want more of the liquid. Consequently, to make up a weight equal to the weight of the floating body, the object will have to sink further into the spirit than into the water. If, on the other hand, the object is placed in a liquid such as mercury, which is denser than water, it will not sink so far, because it will not take so much of this denser liquid to have a weight equal to that of the floating body.

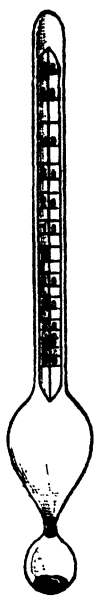


FIG. 67.—A hydrometer for determining the quality of milk. It is called a lactometer.

The hydrometer.—The construction of a simple instrument called the **hydrometer** is based upon these conclusions. The instrument is made in various shapes and graduated differently according to the special use for which it is intended. All hydrometers are, however, used to measure the densities of fluids by observing the depth to which they sink when immersed in them. The **lactometer** is a form of hydrometer employed for measuring the density of milk. When placed in pure milk a lactometer should float with the mark *P* (Fig. 67) on a level with the surface of the liquid. In a mixture of milk and water the lactometer floats with

some other division level with the surface of the liquid. Thus, in milk 10 per cent. below the average density, the 10 above the *P* mark is level with the surface.

An experienced observer is, therefore, able from the readings of a lactometer to tell whether a sample of milk has a correct density, or whether it is heavier or lighter than it should be. At the same time it must be clearly understood that it is not possible to decide at once from the reading of a lactometer whether a sample of milk has been adulterated or not. There are other considerations to be taken into account.

× 24. THE PRINCIPLE OF ARCHIMEDES.

1. Principle of Archimedes.—(*a*) Suspend a metal cube, or any other fairly heavy object, from a spring balance, and notice the reading of the balance. This indicates the weight of the object in air. Immerse the cube in water, as in Fig. 68, and again notice the reading of the balance. It is less than before, and the loss of weight shows the buoyant power or upward pressure of the water.

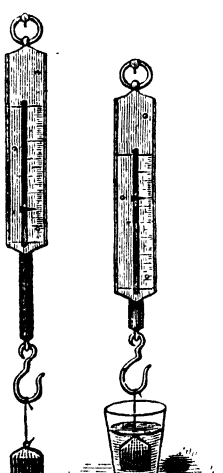


FIG. 68.—The block weighs less when immersed in water than when suspended in air.

(*b*) Find the volume of the cube, or other object used in the last experiment, by noticing the volume of the water it displaces in a graduated jar.

Hang the object from one pan of the balance, as shown in Fig. 69, and determine its weight in grams. Now bring a vessel of water under the pan so that the object is immersed in it, as in Fig. 70. The pan rises. Put gram weights in the pan until the balance sets horizontally as before. You thus find the loss of weight due to the buoyancy or upward pressure of the water. Notice that the number giving this loss in grams is the same as that giving the volume of water in cubic centimetres displaced by the object.

(*c*) This principle, *viz.*, that when an object is immersed in water it experiences a loss of weight equal to the weight of water displaced, can be convincingly demonstrated as follows:—Suspend an object from the left-hand pan of a balance. Place in the short scale pan a small measuring-glass graduated in cubic centimetres. Counterpoise the suspended object and the measuring-glass

together. Now pour water into a graduated jar until the jar is about two thirds full. Notice the level. Bring the jar under the

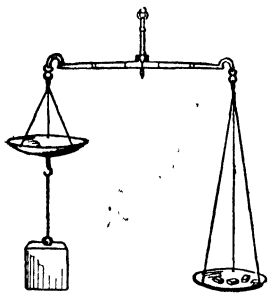


FIG. 69.—Weighing an object in air.

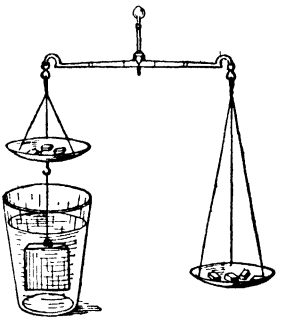


FIG. 70.—The same object weighed in water. Notice that weights are in the short pan to make up for the buoyancy of the water.

short scale pan so that the object is immersed as in Fig. 71. Notice the amount of water displaced. Put water gradually into the measuring-glass by means of a pipette. Equilibrium will be restored when the amount of water added is equal to the amount displaced.

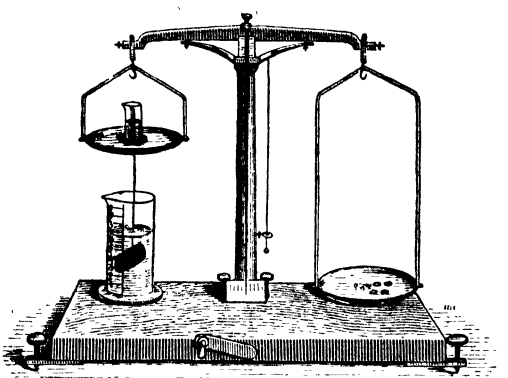


FIG. 71.—When an object is immersed in a liquid it experiences a loss of weight equal to the weight of the liquid displaced by it.

Buoyancy.—Most people have noticed when in a bath that if there is water enough, and they take hold of no support, the

water buoys them up, or they experience a tendency to rise up through the water. It is as if the water resists being displaced and presses the displacing object upwards. In the case of things which float, such as a wooden rod or a lead pencil, the results of this buoyancy which the water exerts can easily be seen, by pushing either the rod or pencil down into the water and then letting go, when the solid rises up through the water. Even in the case of bodies which sink, there is the same buoyancy on the part of the water, but it is not enough to float them. The effect which the water has upon such bodies can, however, be seen in the loss of weight which they experience if they are weighed by a spring balance when hanging in water (Fig. 68).

Loss of weight of things immersed in water.—It is easy to prove by experiment that an object weighs less in water than out of it. If a cubic centimetre of lead or any other heavy material, is hung from a spring balance and then suspended in water, it will be found to weigh the weight of one gram less in water than out. If two cubic centimetres are suspended from the balance, the loss of weight is the weight of two grams. In every case the loss of weight measured in this way is equal to the number of cubic centimetres of the solid immersed in the water. The loss is thus equal to the **weight of the water displaced.** This fact brings us to a highly important conclusion, known after its discoverer as the **Principle of Archimedes.**

The Principle of Archimedes.—When a body is immersed in water it loses weight equal to the weight of the water displaced by it.—If the body be immersed in any other liquid, then the loss of weight is equal to the weight of an equal volume of that liquid. It does not matter what substance the thing is made of; the amount of loss of weight depends upon the *volume* of the part immersed, and not upon the material.

This principle explains many interesting facts. For instance, a ship made of iron, and containing all kinds of heavy things, is able to float in water although the material of which it is made is denser than water. This is because the ship and all its contents only weigh the same as the volume of the water displaced by the immersed part of the hull. Or, the ship as a whole weighs less than a quantity of water the same size as the ship would weigh.

Now, too, it can be seen why some solids float and some sink. When an object weighs more than an equal volume of water it sinks. When an object weighs less than an equal volume of water it floats. When an object weighs the same as an equal volume of water it remains suspended in the water.

A balloon rises in the air because the gas in the balloon, together with the bag and tackle, weighs less than an equal volume of air. If the balloon were free to ascend it would rise to a height where its weight would be equal to the weight of an equal volume of the air around it.

25. RELATIVE DENSITY OF SOLIDS.

1. **Determination of density of solids.**—(*a*) Suspend the solid, the density of which is to be determined, to one side of the balance, letting it hang in an empty beaker standing upon a platform, *H*, *H*,

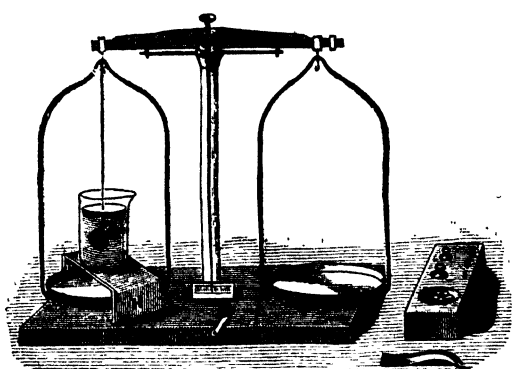


FIG. 72.—How to find the weight of an object suspended in water.

as shown in Fig. 72. Find the weight of the object. Then pour water into the beaker, and find the weight of the object when immersed in water. By subtracting the number thus found from the weight of the object in air, determine the loss in weight of the solid when suspended in water.

(*b*) Another plan of determining the weight of an object in water was explained in the last section (Fig. 70).

This loss of weight equals the weight of a volume of water equal to the volume of the solid. We can therefore write:

$$\text{Relative density of solid} = \frac{\text{weight of the solid in air}}{\text{its loss of weight in water}}.$$

How the relative density of a solid is determined.—The fact that when a body is immersed in water it loses weight equal to the weight of the water displaced by it, provides a means of determining the density of a solid compared with water.

All it is necessary to know is :

1. The weight of the object ; this can be determined by weighing it in air.
2. The weight of an equal volume of water.

The Principle of Archimedes enables this to be done in the following manner :

The object is suspended by means of a fine thread, from one side of the beam of a balance in such a way that it is completely immersed in water. Then by weighing it is found that the weight is less than when hanging in air, on account of the loss of weight in the water. The buoyancy of the water acting upwards overcomes part of the pull of the earth downwards. The difference in the weight of the object in air and its weight when immersed in water gives the weight of a volume of water equal to the volume of the object. From these numbers the density of the solid compared with water as a standard can be at once calculated.

$$\left.\begin{matrix}\text{Relative density of the} \\ \text{solid}\end{matrix}\right\} = \frac{\text{weight of the object in air}}{\text{weight of an equal volume of water}}$$

$$= \frac{\text{weight in air}}{\text{loss of weight in water}}.$$

To find the weight of an equal volume of water, the object could be placed in a graduated jar and the amount of water displaced could be taken out and its weight determined. Or, if the number of cubic centimetres of water displaced is observed, the same number shows the weight of the displaced water in grams.

Example.—A piece of lead was found to have a weight of 100 grams in air, and when suspended in water its weight appeared to be 90 grams. What is its density compared with water ?

What must be done with these numbers to find out the density of the lead compared with water ? Two facts must be known —first, the weight of the object in air, and secondly, the loss of

weight in water, as this gives the weight of an equal volume of water. The loss of weight is obtained by subtraction, thus :

$$\text{Loss of weight} = \begin{matrix}\text{the weight of}\\ \text{the lead in air}\end{matrix} - \begin{matrix}\text{the weight in water}\end{matrix}$$

$$= \text{weight of 100 grams} - \text{weight of 90 grams}$$

$$= \text{weight of 10 grams.}$$

$$\therefore \text{Density of the lead} = \frac{100}{10} = 10.$$

As the density of this piece of lead is 10, one cubic centimetre of it will have a weight equal to that of 10 grams. Also, since the weight of the piece of lead used in the example is 100 grams, the volume must be ten cubic centimetres.

CHIEF POINTS OF CHAPTER VII.

Floating bodies.—When an object floats in a liquid, the *volume of liquid displaced* is equal to the volume of the immersed portion of the object.

The weight of the liquid displaced by a floating object is equal to the weight of the object.

The hydrometer.—The hydrometer is a practical application of the principle of floating bodies. It consists of a suitably loaded tube which is graduated in such a manner that the level at which it floats indicates the density of the liquid in which it floats.

The **lactometer** is a special form of hydrometer used in testing milk. In using it, however, certain other facts, as well as the density, have to be taken into account before pronouncing upon the purity or otherwise of the milk.

Principle of Archimedes.—When a body is immersed in a liquid it loses weight equal to the weight of the liquid displaced by it.

Expressed differently, the up-thrust experienced by an object in water is equal to the weight of the water displaced.

It is in consequence of this that bodies which are bulk for bulk heavier than water, sink, and those bulk for bulk lighter, float.

EXERCISES ON CHAPTER VII.

1. Define mass, volume, and density, and state the relation that exists between them.

Suppose you were given two irregular pieces of metal, one of which was gold and the other gilded brass. How would you find out, by a physical method, which piece was gold?

2. Explain why a ship made of iron will float in water, though iron itself is heavier, bulk for bulk, than water?

3. A number of nails are driven into a rough piece of wood, one cubic centimetre of which weighs 0·5 gram. It is required to find the weight of the nails without pulling them out. How could this be done by experiment ?

4. A bottle weighs 2 ounces. When holding $3\frac{1}{2}$ ounces of shot it will just float in water, when holding 3 ounces it will just float in oil, and when holding $3\frac{3}{4}$ ounces it will just float in brine. Find the specific gravity of the oil and the brine.

5. A piece of metal weighs $19\frac{1}{2}$ grams in air, and $17\frac{1}{4}$ grams in water. What is its specific gravity ?

With what apparatus and in what manner would you find the weight in water ?

6. A stone, weighing in air 1 kilogram, is suspended by a piece of cotton so that it is entirely immersed in water. On attempting to lift the stone out of the water the cotton breaks when the stone is partly out of water. Why is this ?

If, when the stone is completely immersed, the cotton would bear an additional pull equal to the weight of 150 grams, what volume of the stone will be out of the water when the cotton breaks ?

7. A beaker of water stands on the pan of a table spring balance. A block of iron hangs from the hook of an ordinary spring balance, and is gradually lowered into the water till completely submerged. Will there be any changes in the readings of the balances ? If so, what changes, and why ?

8. Two blocks of glass, each having a volume of 10 c.c., are hung from the scale-pans of a balance by means of hooks under the pans, and balance one another. Under one is brought a beaker of water, under the other a beaker of alcohol, so that the blocks are immersed in the liquids. The balance is now disturbed, and it is found that 1·82 grams have to be added to one pan to restore equilibrium. To which pan has this weight to be added, what is the explanation of the fact, and how can you determine from the figures now at your disposal the density of the alcohol ?

9. Explain the statement that " the specific gravity of iron is 7·8." Describe one method of finding this specific gravity.

10. What do we mean when we say that the " specific gravity " of mercury is $13\frac{1}{2}$? What things are necessary for the experimental determination of this specific gravity ? Say how you would proceed to make the determination.

11. Explain how you would ascertain (1) the volume and (2) the specific gravity of an apple or potato.

12. A porcelain weight of one pound breaks into two unequal fragments. How could you without using a spring balance or ordinary balance determine the weight of each fragment ?

CHAPTER VIII.

FLUID PRESSURE.

26. PRESSURE OF LIQUIDS.

1. Relation between pressure of liquid and depth.—(*a*) Bend a piece of glass tubing, as shown in Fig. 73, the long arm having a length of about 40 cm. Mount the tube firmly upon a strip of wood having half-centimetre divisions marked upon it, or upon a metric scale. Pour enough mercury into the tube just to fill the bend. Notice that the level of the mercury is the same in the two arms of the tube.

Lower the frame into a tall jar of water so that the open end of the short arm of the tubes is 10 cm. below the surface of the water. Notice the difference of level of the mercury in the two arms, and record it.

Lower the tubes an additional 10 cm. and observe the effect. Then lower the frame as far as it will go, and again note the difference between the heights of the mercury in two arms.

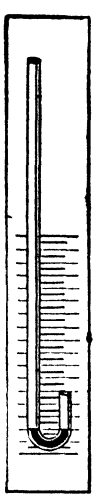


FIG. 73.—Arrangement for determining the pressure of water at different depths.

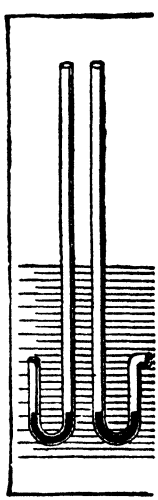


FIG. 74.—Arrangement for showing that pressure in any particular liquid depends only on depth.

(*b*) Repeat the preceding exercise, using turpentine, or a strong solution of salt, or any other convenient transparent liquid instead of water, and record your results as before.

(*c*) To show that the pressure in any particular liquid depends only upon the depth, mount two tubes upon a frame, as shown in Fig. 74—

one with the lower opening turned upwards, and another with the opening pointing sideways. Pour the same quantity of mercury into the tubes. Arrange the tubes so that the lower openings are at the same level. Lower the frame, a few centimetres at a time, into water, and observe the difference of level of mercury produced in each tube.

(*d*) Immerse the frame of tubes to any level which causes the mercury to rise by a clearly-marked amount. Turn the frame to face in various directions, keeping it at the same depth, and notice that the pressure, as indicated by the rise of mercury, is the same in all directions at any particular depth. Test this for several depths.

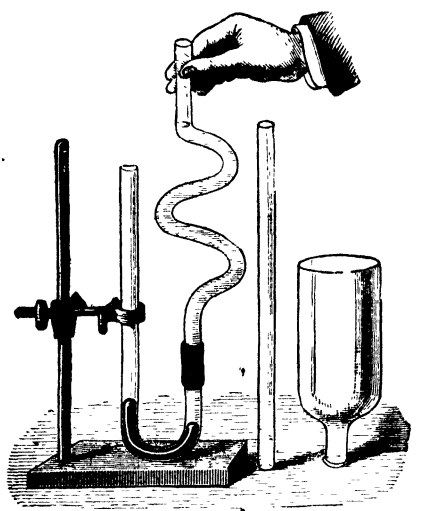


FIG. 75.—Experiment to show that liquid pressure is independent of the form of containing vessel.

ii. Pressure of liquid independent of the form and volume of containing vessel.—(*a*) Bend a short length of fairly wide glass tubing into a ∪-form, as shown in Fig. 75. Fit a short piece of india-rubber tubing on the short arm. Pour sufficient mercury into the tube to cover the bend. Connect a straight piece of glass tubing with the india-rubber, and introduce water into it. Measure the difference of height of the mercury in the two arms produced by the pressure of the column of water. Substitute a funnel, or a curved tube for the straight one, and fill it to the same height with water. Measure again the difference of height of the mercury in the two arms of the ∪-tube. The pressure upon the mercury depends merely upon the height or the **head of liquid.**

iii. Upward pressure of liquid.—(*a*) Procure a piece of wide glass tubing, or a straight lamp glass, having one end flat. If stiff leather is available, cut out a disc of slightly larger diameter than that of the glass, and pass a knotted thread through its centre. If a leathern disc cannot be obtained, make a disc of wood or stiff cardboard. Hold the disc tightly upon the flat end of the glass by means of the thread, and while doing so, lower the glass into a jar of water (Fig. 76). When the end of the glass with the disc upon it is a few inches below the surface of the water, the thread can be released and the disc will be found to remain in its position.

Pour water very gently into the inside of the glass, and notice the height inside and out at that moment when sufficient water has been introduced to make the disc drop.

The pressure a liquid exerts depends upon the depth below the surface.—Since water and other liquids are material substances, they are pulled downwards by the attractive force of the earth. The amount of this pull measures their weight. The greater the distance below the surface the longer is the column of liquid above any area under consideration ; and, as a natural consequence, the greater is the weight of the column of liquid above the given area.

The weight of the column of liquid above unit area, say a square inch, at any depth, measures the **pressure of the liquid** at that depth, and this pressure increases from the surface of the liquid downwards, and is, indeed, directly proportional to the depth. Moreover, as the pressure is measured by the weight of the column of liquid above unit area, it is clear that the denser the liquid the greater is the weight of a column of it, and the greater the pressure it exerts.

The pressure per unit area exerted by a liquid depends *only* upon the depth. The direction is immaterial, a fact which is clearly brought out by Experiment 26 i. (*d*), where it is seen that whatever the direction in which the open ends of the bent tubes containing mercury point, the difference in height of the level of the mercury in the two limbs is the same when the depth of the open ends of the tubes beneath the water is the same.

Relation between the pressure of a liquid and the area upon which it acts.—The pressure upon a given area at any depth below the surface of a liquid is measured by the weight of the column of liquid above it. So long as the depth remains the same the pressure upon unit area, a square inch for instance, remains the same. But the larger the number of square inches upon which the pressure of liquid is felt, the greater the **total pressure** upon the area. These facts may be expressed by saying that the total pressure varies in proportion to the areas of the surfaces at any given depth ; or, though the pressure (which means per unit area) at any given depth is constant, the total pressure is dependent upon area.

These considerations, too, provide an explanation of Experiment 26 ii., where vessels of different shapes and volumes are

attached, one by one, to a bent tube with mercury in its bend (Fig. 75). In the circumstances of the experiment, the difference in level of the mercury in the two limbs of the tube containing it remains the same, and consequently the pressure exerted by the water is the same whichever of the variously shaped vessels is attached, provided that the vertical height of the liquid in the tubes is the same. It is necessary to remember that the height and cross-area of the column of liquid, and the density of the liquid, are the only things which alter the pressure. None of these factors vary, however much the shape of the vessel alters. Whatever the shape of the vessel, it acts just as if it were of the simple cylindrical form shown in the middle of Fig. 75. When the wide vessel is attached to the bent tube provided with mercury, the effective pressure at the base is due to the water contained by a cylinder of cross-sectional area equal to that of the upper surface of the mercury, and a vertical height equal to that of the water in the vessel above the mercury.

Upward pressure in a liquid.—It has been seen how the pressure at any point in a liquid is estimated, and how the amount of such pressure is determined. It has also been proved by experiment that the pressure is the same in different directions. It would, consequently, be surmised that the pressure acting vertically upwards at a point in a liquid is equal to the downward pressure of the liquid above the point. That this is actually true is easily demonstrated by the simple apparatus shown in Fig. 76. A wide glass cylinder, one end of which is covered with a disc of leather to which a string is attached, as shown in the illustration, is lowered into a liquid. It is found that water may be poured into the cylinder, without disturbing the disc, until the level of the water inside the cylinder becomes the same as that outside; but, if more water be added, the downward pressure on the upper surface of the disc now exceeds the upward pressure of the outside liquid on

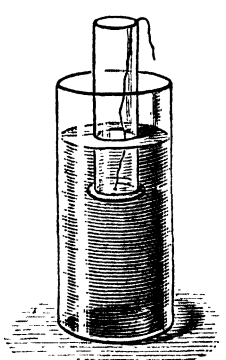


FIG. 76.—Upward pressure of a liquid.

the lower surface of the disc, which is therefore forced downward. So long as the level of the water inside the cylinder is below that of the water outside, the upward pressure on the disc from outside exceeds the downward pressure from inside, and the disc is held firmly against the end of the cylinder. When the water level is the same inside and outside the cylinder, the disc is acted upon by exactly equal upward and downward pressures.

27. THE PRESSURE OF THE ATMOSPHERE.

1. **Atmospheric pressure.**—(*a*) Tie a piece of thin sheet india-rubber, such as that used in toy air-balls, over the top of a funnel. While the rubber remains flat the pressure is evidently the same on both sides of it. Blow into the funnel. Explain why the india-rubber is forced out. Suck the air out of the funnel. What is pressing upon the outside and forcing the india-rubber into the funnel? Suck the air out of the funnel and place your thumb over the open end so as to prevent air from entering. Turn the funnel in various directions and see whether you can detect any difference in the amount of bulging of the india-rubber. If not, you may conclude that the pressure of the air upon the outside is the same in all directions.

(*b*) Fill a gas collecting jar, or a tumbler having a flat rim, with water; cover it with stout paper and invert. Why does not the water drop out?

(*c*) Place in water one tube of the Hare's apparatus for determining density (p. 64), and the other tube in mercury. Suck out the air. The liquids rise. Why? Notice the difference in the level of the mercury and the water. Explain the cause of this difference. Let one of the tubes of the Hare's apparatus be much wider than the other. Place the ends of the tubes in mercury, and suck out the air. Is there any difference in the height of the mercury in the large and small tubes?

The pressure exerted by the atmosphere.—The gaseous envelope which surrounds the earth is a fluid, and the pressure it exerts at different distances from the earth varies considerably. The condition of things is different from what has been found in the case of liquids. Gases are easily compressible. The consequence is, in the case of the air, that the pressure is greatest near the earth or at the bottom of the gaseous ocean round the globe—for so we may regard the atmosphere. But the pressure does not increase in the same simple way as in the case of water. At a point midway between the surface and the

bottom of the water in a tank the pressure is just half of that at the bottom of the tank. But in the case of the air the pressure at a height of 3½ miles is just half that at sea-level, though the atmosphere extends for as much as 150 miles from the earth's surface. Owing to the compressibility of air the lower layers are much denser than those at a greater height; they are consequently heavier, bulk for bulk, and are able, per unit volume, to exert a greater pressure.



FIG. 77.—A barometer. The pressure of the atmosphere acting upon the open end of the short tube supports a column of mercury about 30 inches long in the closed tube.

Just as in the case of water, however, the deeper we go into the gaseous envelope, that is, the nearer the earth we get, the greater is the pressure the air exerts. Or, conversely, the higher we ascend into the atmosphere, or the farther from the earth's surface we travel, the smaller is the pressure of the air, simply because, as we ascend we diminish the column of air above us.

28. MEASUREMENT OF THE PRESSURE OF THE AIR.

1. The principle of the mercurial barometer.—Procure a barometer tube and fit a short piece of india-rubber tubing upon its open end. Tie the free end of the tubing to a glass tube about six inches long, open at both ends. Rest the barometer tube with its closed end downwards and pour mercury into it (being careful to remove all air bubbles) until the liquid reaches the short tube. Then fix the arrangement upright as in Fig. 77. The mercury in the long tube will be seen to fall so as to leave a space of a few inches between it and the closed end. The distance between the top of the mercury column in the closed tube and the surface of that in the open tube will be found to be about 30 inches.

ii. The cistern barometer.—Procure a thick glass tube about 36 inches long and closed at one end. Fill the tube with mercury; place your thumb over the open end; invert the tube; place the open end in a cup of mercury and take away your thumb (Fig. 78). Measure the distance from the surface of the mercury in the basin to the top of the mercury column.

Slant the barometer and measure the vertical height of the mercury by means of a plumb-line. Compare it with the previous reading.

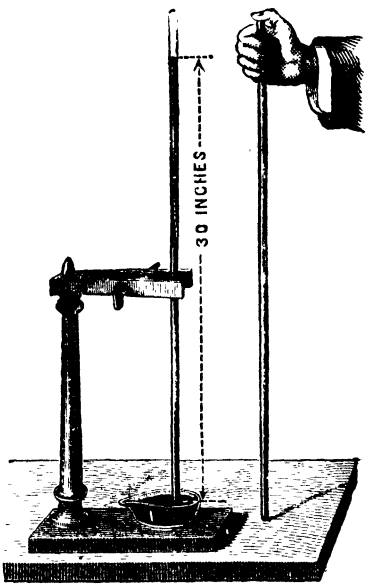


FIG. 78.—Construction of a barometer.

The mercurial barometer.—The apparatus represented in Fig. 77 calls to mind the U-tube method of determining densities by balancing columns of liquid (p. 63). Here, however, the column balancing the mercury in the tube *A* is not apparent. The experimental fact that the column *A* does not rush up through *O* to adjust the liquid levels to equality demands explanation. Remember that the top of the column *A* is closed and protected while *O* is open. Something acts at *O* and is capable of sustaining the column *A*. This something is the weight of the air. In fact, a column of air many miles high balances a column of mercury about 30 inches high.

If, however, a hole were made in the closed end of the tube the balance would be disturbed, and the mercury would fall down the long tube and flow out of the short tube *O*.

The column of mercury is kept in its position by the weight of the atmosphere pressing upon the surface of the mercury in the short open tube. The weight of the column of mercury and the weight of a column of the atmosphere with the same

sectional area is exactly the same; both being measured from the level of the mercury in the short stem of the apparatus shown in Fig. 77, the mercury column to the top of the column in the long tube, the air to its upper limit, which, as has been seen, is a great distance from the surface of the earth. If for any reason the weight of the atmosphere becomes greater, the mercury will be pushed higher to preserve the balance; if it should become less, then similarly the amount of mercury which can be supported will be less, and so the height of the column of mercury is diminished.

The height must in every case be measured above the level of the mercury in the tube or cistern open to the atmosphere. In the usual arrangement, illustrated by Fig. 77, a line is drawn at a fixed point *O*, and the short tube is shifted up or down until the top of the mercury in it is on a level with this line.

The student will now understand why it is so necessary to remove all the air bubbles in Experiment 28 i. If this is not done, when the tube is inverted the enclosed air would rise through the mercury and take up a position in the top of the longer tube, above the mercury. The reading would not then be thirty inches, for instead of measuring the whole pressure of the atmosphere, what we should really be measuring would be the difference between the pressure of the whole atmosphere and that of the air enclosed in the tube. In a properly constructed barometer, therefore, there is nothing above the mercury in the tube except a little mercury vapour.

An arrangement like that described constitutes a barometer, which may be defined as an instrument for measuring the pressure exerted by the atmosphere.

The cistern barometer.—Other forms of barometer are often employed for the determination of the pressure of air. A very common arrangement is that of Experiment 28 ii., which is a repetition of one by an Italian physicist, Torricelli. The principle of its action is precisely that of the barometer just described, except that the U-tube principle is not immediately apparent. There is, however, the same balance maintained between the column of mercury in the tube and a column of air outside it, pressing down upon the mercury in the basin.

A column of mercury will be supported in the tube by the

pressure of the atmosphere. The distance between the top of the column and the surface of the mercury in the cup will be about 30 inches, or 76 cm., when the tube is vertical (Fig. 79, *b*). If the tube is inclined so that the closed end of it is less than this height above the mercury in the cup (Fig. 79, *c*) the mercury

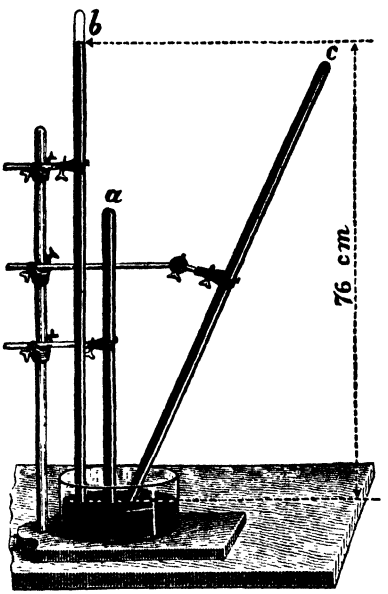


FIG. 79.—The atmosphere at sea level will support a column of mercury up to 30 inches, or 76 cm. in length.

fills it completely ; and if the tube is less than 30 inches long, it is always filled by the mercury whether it is inclined or not (Fig. 79, *a*). On an average the atmosphere at sea-level will balance a column of mercury 30 inches in height. No matter if the closed tube is 30 feet long, the top of the mercury column will only be about 30 inches above the level of the mercury in the basin.

The empty space above the column of mercury in the tube is often referred to as the **Torricellian vacuum.**

29. BOYLE'S LAW.

1. Relation between pressure and volume of gases.—(*a*) Select a glass tube about 20 cm. long, and neatly closed at one end (Fig. 80, *A*).

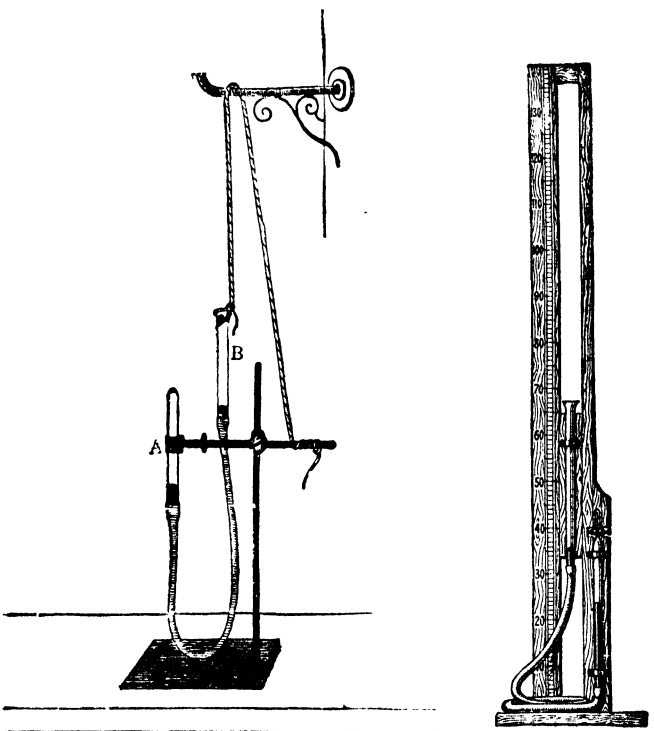


FIG. 80.—Two forms of apparatus to show the relation between the pressure and volume of a gas.

Tie a piece of stout india-rubber tubing about a metre long upon the open end of the tube, and fix the other end of the tubing upon a glass tube *B*, about 20 cm. long, open at both ends. Now carefully fill the tubes with mercury until the level of the liquid is about 10 cm. from the open end. Afterwards fix the closed

tube upright in a retort stand with the sealed end upwards, place your finger upon the open end of the other tube, and lower the open end so as to make the air pass into the closed tube. This apparatus will enable you to measure the expansion and compression of air.

Support the tube *B* with its open end upwards, and at such a height that the mercury stands at the same level in the closed and open tubes. The imprisoned air is then at the same pressure as the air outside.

If the closed tube is uniform in bore, and the inside of the sealed end is nearly flat, the volume of the imprisoned air is proportional to the length of tube occupied by the air, so that if the air is made to occupy one-half the original length of tube its volume is one-half the original volume. The pressure upon the enclosed air is equal to the pressure due to the column of mercury between the level of mercury in the closed tube and that in the open tube, *plus* the pressure of the atmosphere. Observe the height of the barometer, and make the difference of level of the mercury in the two tubes equal to it by lifting the open tube. Then record as follows:

Height of the barometer, - - - - - - ... cm.

Length of air column when the mercury is at the same level in both tubes, that is, when the imprisoned air is at atmospheric pressure, - - ..... cm.

Height of barometer.... cm. + equal height in tube, - - - - - - - - - - .... cm.

Length of air column under pressure of two atmospheres, - - - - - - - - - ... cm.

The pressure to which the enclosed air is subjected in the second case is double that in the first case; find the amount by which the volume of the air, represented by the length of the air column, has been diminished.

(*b*) Lower the open tube until the air in the closed tube almost reaches the india-rubber junction. Measure the length of the air column, and the difference of level of the mercury in the two tubes. Repeat the experiment by reading the volume of air and the head of mercury at every few centimetres up to the highest point you can raise the open tube. Record your results as indicated below:

Height of Barometer in cm.	Difference of Level of Mercury in cm.	Total Pressure on the Air, *P*.	Volume of Air, *V*.	Volume × Total Pressure. ($P \times V$.)

ii. **A simple form of Boyle's Law apparatus.**—Procure a length of thermometer tubing, AB (Fig. 81), about 75 cm. long and 1 mm. bore. Seal it at B and expand the end A somewhat. Clamp AB in a vertical position by the side of a metre scale, and connect a small funnel to A by means of a short piece of rubber tubing. Pour a little pure, clean mercury into the funnel and induce it to run down the bore of the tube by inserting a thin, clean, steel wire. In this way any desired volume of air can be enclosed.

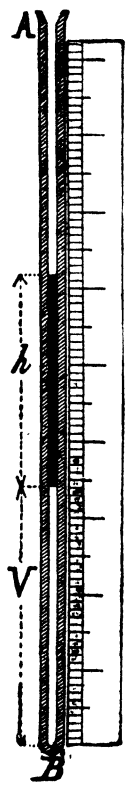


FIG. 81.—Simple form of apparatus for verifying Boyle's Law.

The length of the column of enclosed air may be taken to represent its *volume* (V). If H = the height of the barometer, and h = the length of the mercury thread (both expressed in the same units), then the total pressure on the enclosed air = $(H+h)$.

Introduce more mercury in the same manner, and in this way alter the values of V and h. The volume of the air under the pressure of the atmosphere alone can be observed by laying the glass tube flat on the table.

Perform several experiments and record your results in the following way:

Volume (V).	Pressure ($H+h$).	Volume × Pressure.

Boyle's Law.—Before a student can clearly understand how and why the density of the atmosphere varies (p. 107), it is necessary to become acquainted with the rule expressing the relation between the volume and pressure of a gas. This can be satisfactorily done by one of the forms of apparatus employed in Experiments 29 i. and 29 ii., which provide a means of subjecting an enclosed quantity of air to varying pressures, by the addition of smaller or larger quantities of mercury. When in the apparatus shown in Fig. 80 the mercury in both tubes stands at the same level, the enclosed air is at the same pressure as the air of the room, but as the tube B is raised, the mercury in it stands higher than that in A, and the air enclosed

in *A* is under a total pressure equal to the sum of that due to the atmosphere and that due to a column of mercury equal in length to the difference of levels of the mercury in *A* and *B*. In these circumstances the volume of the air in *A* decreases, and it decreases more and more as the total pressure is increased. When the results of experiments with any satisfactory form of apparatus are tabulated, certain very important relations between the volume of a gas and the pressure to which it is subjected become evident. It is found that the volume regularly diminishes as the pressure is increased, and in the same ratio. The converse is also found to be true, viz., that when the volume of a gas increases the pressure upon it has diminished, and exactly at the same rate.

But, in both these cases, it is understood that the temperature of the gas remains the same, that is, the temperature of the gas under the different pressures must not alter.

The tabulated results of the experiments reveal another important relation, which is, however, another way of expressing those already noticed. It is found that, when there is no alteration of temperature, the product obtained by multiplying the volume of a given quantity of gas by the pressure to which it is subjected is always the same, or remains constant.

These facts were discovered by Boyle, and are included in what is known as Boyle's Law. It can be expressed by saying that when **the temperature remains the same, the volume of a given mass of gas varies inversely as its pressure.** Or, what is the same thing, **the temperature remaining the same, the product of the pressure into the volume of a given mass of gas is constant.**

But it has been seen that if the volume occupied by a given weight of a substance is increased, its density is decreased, and if the volume is decreased, its density is increased. Therefore, by decreasing the volume of the enclosed air in the above experiment, its density is increased. The increase of density and the increase of pressure are proportional to one another. It is not difficult to apply these facts to the case of the atmosphere. It has been learnt that the pressure of the atmosphere decreases as we ascend, and we are now able to add that its density decreases also and at the same rate. Therefore the densest atmosphere is that at the surface of the earth, leaving out, of course, the air

of mines and other cavities below the surface, where the air is denser still. The air gets less dense, or rarer, as we leave the earth's surface, until eventually it becomes so rare that its existence is practically not appreciable.

30. SOME INSTRUMENTS DEPENDING UPON FLUID PRESSURE.

i. The air-pump.—(*a*) Select a cork or india-rubber stopper which will fit tightly into a brass or glass tube about a centimetre in diameter. Bore a hole through the cork lengthways, and then tie a narrow strip of gold-beater's skin, or oiled silk, over one end of the cork so as to cover the hole. You will find that the silk does not prevent you from blowing through the cork from the open end, but that you cannot suck air through. An arrangement of this kind,

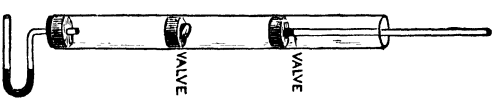


FIG. 82.—Model of an air-pump.

which allows a passage in one direction but not in the other, is called a *valve*. Wax the cork and fit it tightly in the brass tube (which should be about a foot long) so that the valve is inside. Push a similar valve half way through another brass tube of about the same length, but wider, and wrap darning cotton around the stopped end of the narrow tube until this tube fits nicely into the wider one, and can be moved to and fro like the piston of a popgun. Fit a short U-tube into the cork at one end of the wide tube, as in Fig. 82, and pour a little mercury into it. The instrument you have constructed is exactly the same in principle as a simple air-pump (Fig 83).

Moisten the cotton on the piston, either with oil or water. Describe what happens to the enclosed air in the wide tube, and how the valves and the mercury in the U-tube are affected, in each of the following cases: (1) When the piston is pushed inwards; (2) when the piston is pulled outwards. Explain how such an instrument as this enables the enclosed air to be rarefied.

ii. The common pump.—Examine a glass working model of a common pump (Fig. 84). Use the pump to lift water, and explain clearly how it acts. Notice that both the valves open outwards or upwards as in the air-pump.

iii. The siphon.—(*a*) Make a siphon by bending a glass tube in a fish-tail gas flame: make one limb about 6 inches long and the other at least a foot. Fill the siphon with water, either by placing it in a bucket of water and covering each open end with one of your fingers before lifting it out; or, by sucking water through it as when using

your pipette. Allow the siphon to empty itself; from which end does the water flow?

Fill the siphon again, dip the short limb in a beaker of water, and notice what occurs when you take your finger from the long arm. Fill again and let the water flow into a tall narrow vessel. Keep the beaker full of water and notice when the flow of water stops.

(*b*) Connect two short pieces of glass tubing with india-rubber tubing. Fill the tubes with water, and insert the ends below the surface of water in beakers or flasks about half full. Lift one of the flasks, and notice the flow of water which takes place. Show that no flow occurs when the level of liquid is the same in both vessels. Test whether a flow takes place when the bend of the india-rubber is below the lower of the two vessels.

Observe whether a siphon will act if there is a hole in it.

The air-pump.—Several forms of air-pumps are in use, but in this place it will be sufficient to describe one of the simplest, that designed by Hawksbee, the essential parts of which are shown in Fig. 83. V is the receiver, from which it is required to remove air. The plate on which V rests is connected with a cylinder *C* by means of a tube, shown in the illustration, bent twice at right angles. At the end of this tube, remote from the receiver and just at the bottom of the cylinder *Cv*, is a valve *v* opening upwards. In the cylinder works, in an air-tight manner, a piston provided with a valve *v'* opening upwards; and a handle for pulling the piston up and pushing it down is provided. The action is very simple. Imagine the piston to be at the bottom of the cylinder to begin with, and that it is then gradually pulled up. As this takes place, the air in the receiver and below the valve *v* is subjected to a diminished pressure, and consequently expands, filling the space which is formed as the piston moves upwards. This continues until the piston arrives at the end of its stroke. The piston is now pushed down. This compresses the air between *v'* and *v* and increases its pressure,

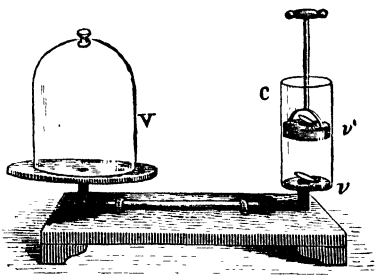


FIG. 83.—Simple form of air-pump.

causing the valve v to shut. But as the piston descends the pressure on the under surface of the valve v' becomes greater than that of the atmosphere upon its upper surface, with the result that the valve v' opens upwards and the air in the space vv' rushes through the open valve into the outside air. The final result, when the piston reaches the bottom of the cylinder, is that there is less air in the receiver and tube connecting therewith than there was originally. As the piston is worked up and down the same opening and shutting of valves is repeated, with the result, that by and by, all the air is removed from the receiver.

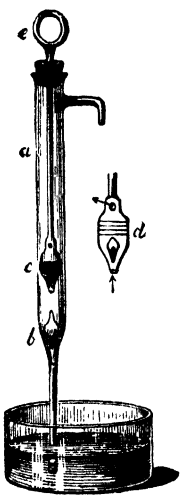


FIG. 84.—Model of a common pump.

The common pump.—After examining a glass model like that shown in Fig. 84, there is no difficulty in understanding the action of a common pump. To begin with, suppose that the pump is full of air and that the end of the tube below the valve b is dipped into a basin of water. The piston ec is, to start with, at the bottom of its stroke near the valve b. As the piston is raised the air in the cylinder above b expands, and its pressure consequently decreases, the pressure on the lower surface of the valve b is, therefore, soon greater than that on its upper surface and it is pushed upwards by the air below it, the air flowing into the cylinder acb. The result of this is that the air in the cylinder below the piston is of a lower pressure than that of the outside air, and as a consequence water is pushed up the tube b. This action continues until the piston reaches the end of its stroke towards the top of the pump.

As the piston descends, the air in the cylinder below the piston c is compressed and its pressure becomes gradually greater. This closes the valve b and opens that in c, through which latter, of course, the air in the cylinder escapes. On raising the piston again the same effects are repeated until all the air in the pump is removed and the outside air pushes the water up until it reaches the spout and escapes.

It has been learnt that the air is able to support a column of mercury 30 inches in height, and as mercury is about $13\frac{1}{2}$ times heavier than water it is clear that the air could support a column of water of a height equal to

$$30 \text{ inches} \times 13\tfrac{1}{2} = 2\tfrac{1}{2} \times 13\tfrac{1}{2} \text{ ft.} = 33\tfrac{3}{4} \text{ ft.},$$

and it will at once be understood, since the efficacy of the common pump depends wholly upon the pressure of the air, that theoretically the spout of the pump must never be more than $33\frac{3}{4}$ feet from the level of the water. Practically the spout of an ordinary pump cannot be much more than 30 feet above the level of the water.

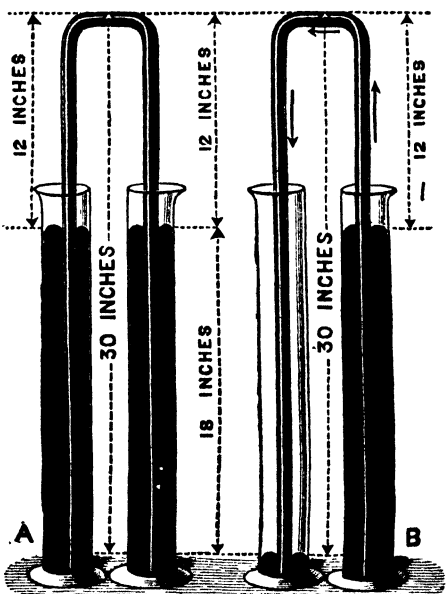


FIG. 85.—To explain how the action of a siphon depends upon atmospheric pressure.

The siphon.—The siphon is a simple instrument which can be easily understood after what has been said about atmospheric pressure. It consists usually of a bent tube, one leg of which is

longer than the other. It is filled with the liquid to be transferred from one vessel to another, and while both ends of the tube are kept closed, the shorter limb is placed into the vessel of liquid. The result is that the liquid flows until the level of the liquids is the same in both vessels, or the higher liquid has been siphoned to the lower level.

Suppose a siphon having limbs of equal length to be placed with the ends in two jars containing mercury at the same level, as in Fig 85, *A*. For simplicity, suppose each tube to have a length of 30 inches, and let the surface of the mercury in each vessel be 12 inches from the bends of the tube. Under normal conditions the atmosphere is able to support 30 inches of mercury, but in the case illustrated the height of the mercury columns is only 12 inches. There is thus a surplus atmospheric pressure equal to 18 inches of mercury acting on the surface of the mercury in the jars; but as it is the same in each, the mercury in the bent tube does not move.

But now consider the conditions represented in Fig. 85, *B*. The 30 inches of mercury in the left-hand tube just balances the atmospheric pressure on the surface of the mercury into which it dips. On the right-hand side, however, the mercury column is only 12 inches high, so there is a surplus atmospheric pressure equal to 18 inches of mercury. The mercury is therefore forced through the tube, and the flow goes on until the level is the same in each vessel. It is thus seen that the **force tending to push the liquid up an arm of a siphon is equal to the atmospheric pressure minus the pressure due to the liquid in that arm.**

Since the instrument depends upon the pressure of the atmosphere for its efficiency, it is clear that if the bend of the siphon is at a greater height above the level of the liquid than that which can be supported by the pressure of the atmosphere, then the siphon will not act.

When the liquid is water, the height of the bend above the higher liquid surface must not exceed 33 feet (p. 119), and when mercury is being transferred by a siphon, this height must be under 30 inches.

CHIEF POINTS OF CHAPTER VIII.

Pressure on liquids.—Liquids communicate pressure equally in all directions. The pressure *per unit area* which a liquid exerts depends only upon the depth below the surface—the direction is immaterial. The *total* pressure on any area within a liquid varies with the area considered. The pressure acting vertically upwards at a point in a liquid is equal to the downward pressure of the liquid above the point.

The pressure of the atmosphere is due to its weight. The pressure diminishes as an ascent from the earth is made. At a height of 3½ miles the pressure is only one-half that at sea-level.

The barometer is an instrument for measuring the pressure exerted by the atmosphere. The pressure of the atmosphere balances a column of mercury about 30 inches in height.

Boyle's Law states the relation between the volume and pressure of a gas. It can be expressed by saying that, the temperature remaining the same, the volume of a given mass of gas varies inversely as its pressure. Or, what is the same thing, the product of the pressure into the volume of a given mass of gas is constant.

The air-pump is an instrument for removing the air from a closed vessel or space.

The **common pump** and the **siphon** depend for their action upon the pressure exerted by the atmosphere.

EXERCISES ON CHAPTER VIII.

1. Some deep-sea animals, when brought to the surface, become very much larger. Why is this?

2. Describe a simple experiment to show that the pressure in a liquid increases with the depth below the surface. What pressure, expressed in grams' weight, would be exerted upon a square of glass of a decimetre side immersed to a depth of 10 metres in the water of a lake?

3. How would you prove to a class that the upward pressure per unit area at a depth of 10 feet below the surface of the sea is equal to a column of water 10 feet high and unit area in cross-section?

4. Given a glass tube, thirty-two inches long, closed at one end, a bottle of mercury (quicksilver), and a small cup. State how you proceed (*a*) to construct a barometer, and (*b*) to show the readings of this barometer.

5. State the principle on which the action of a mercurial barometer depends. Why is a water barometer longer than a mercurial barometer? What occupies the space above the mercurial column in the latter instrument? If a hole were to be bored through the glass above the column of mercury, what would happen?

6. How can the weight of the air be determined? In what way is the pressure exercised by the atmosphere on the earth's surface, in consequence of its weight, stated? How is it that we are able to move about under the weight of the atmosphere?

7. Enunciate the law associated with the name of Boyle. Describe a simple form of apparatus to prove that, if a given mass of air is acted upon by a pressure twice as great as that of the atmosphere, its volume becomes one half what it was originally.

8. Draw a common pump, and explain how it acts. At what depth will such a pump become ineffective?

9. What is a siphon? What experiments would you perform to explain its action to a class?

10. Make a sketch of some form of air-pump and describe how the instrument enables the air to be removed from a closed vessel.

11. About what height does the mercury column of the barometer generally stand? Explain why the tube must be kept in a vertical position.

12. What do you know about the contents of that part of the barometer tube which is not filled by the mercury? What simple experimental proof of the truth of your answer can be given?

13. Three similar bottles of water are turned upside down, the first under water, the second free to the air, the third free to the air but fitted with a cork, through which a narrow hole has been drilled. Compare and explain what happens in each case.

CHAPTER IX.

EFFECTS OF HEAT. THERMOMETERS.

31. EXAMPLES OF EXPANSION BY HEAT.

1. **Expansion of solids.**—(*a*) Take a metal ball suspended by a chain as shown in Fig. 86, and suspend it by the side of a metal ring, through which it just passes easily. Heat the ball in a laboratory burner for a few minutes, and then try to drop it through the ring. It is too large and rests on the ring. Now allow it to cool

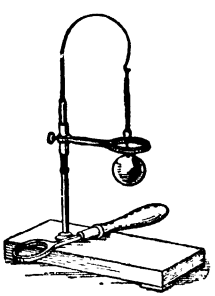


FIG. 86.—Apparatus for showing that a metal ball is larger when hot than it is when cold.

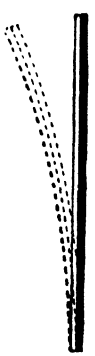


FIG. 87.—Compound strip of ebonite and wood to show the greater expansion of ebonite when heated.

slowly and notice that after a short time it gets smaller and will slip through.

(*b*) Solder a strip or wire of brass, about two feet long, to one of iron of the same length. Straighten the compound strip by hammering; then heat it. Notice that the strip bends, because the brass expands more than the iron. The same effect can be shown by means of a strip of ebonite glued to a strip of wood, on account of the ebonite expanding more than the wood (Fig. 87).

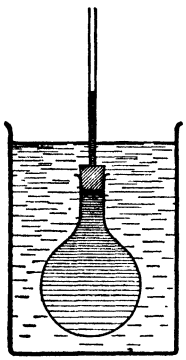


FIG. 88.—The expansion of a liquid.

ii. Expansion of liquids.—(*a*) Procure a 4-oz. flask and fit it with a cork. Bore a hole through the cork and pass through it a long glass tube which fits tightly. *Fill* the flask with water coloured with red ink. Push the cork into the neck of the flask and so cause the coloured water to rise up the tube. See that there is no air between the cork and the water. Now dip the flask in warm water, and notice that the liquid soon gets larger and rises up the tube (Fig. 88). Take the flask out of the warm water, and see that the coloured water gets smaller as it cools, and that it sinks in the tube.

(*b*) Arrange two other flasks as in the last experiment, but filled respectively with alcohol and turpentine. Push in the corks till the liquid stands in each tube at the same height. Put the flasks to the same depth into a vessel of warm water. Notice that the expansion of the glass causes a momentary sinking of the liquids; and that ultimately the expansions of the liquids are very different.

iii. Expansion of gases.—(*a*) Procure a well-made paper bag and tightly tie a piece of tape round the open end. Hold the bag in front of the fire, and notice that the air inside gets larger and inflates the bag.

(*b*) Or, obtain a flask with a cork and tube as in Fig. 89, i. Remove the cork and tube, and, by suction draw a little red ink into the end of the tube near the cork. Re-insert the cork, and gently warm the flask by clasping it in your hands. Notice that the air in the flask gets larger and pushes the red ink along the tube.

(*c*) Turn over and place the open end of the tube beneath the surface of some coloured water in a beaker. Warm the flask with the hand or a flame so as to expel some of the air, and let the liquid rise in the stem (Fig. 89, ii.). This arrangement constitutes an *air thermometer*.

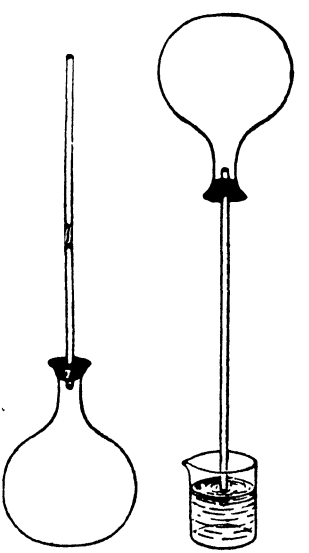


FIG. 89.—Experiments to show the expansion of air by heat.

(*d*) Fasten in an air-tight manner two bulbs or flasks together by a tube bent six times at right angles, and containing some coloured liquid in the middle bends (Fig. 90). Show that the liquid moves if one flask is warmed more than the other. This instrument is known as a *differential thermometer.*

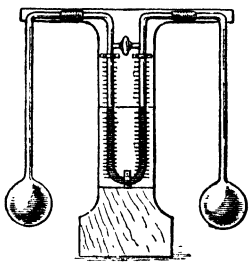


FIG. 90.—A simple form of differential thermometer.

Change of size. Expansion.—As a rule all bodies, whether solid, liquid, or gaseous, get larger when heated, and smaller when cooled.

The change of size which a body undergoes is spoken of as the amount it expands or contracts; or, heat is said to cause expansion in the body. This expansion is regarded in three ways. When dealing with solids, expansion may take place in length (**linear expansion**), in area (**superficial expansion**), and in volume (**cubical expansion**). In the case of liquids and gases we have only cubical expansion. Similar terms can be used with reference to contraction.

The expansion which substances undergo when heated must be allowed for in many engineering projects. Railway lines, for instance, are usually not placed close together, but a little space is allowed between the separate rails, so that they can expand in summer without buckling. Steam pipes used for heating rooms are also not fixed firmly to the walls at both ends, but are left slightly loose or are loose-jointed, so that they can expand or contract without doing any damage. For the same reason the ends of iron bridges are not fixed to the supports upon which they rest. Iron tyres are put on carriage wheels by first heating the tyre, and while it is hot, slipping it over the wheel. As the tyre cools it contracts and clasps the wheel tightly.

The common occurrence in domestic life of the cracking of thick glasses, when boiling water is poured on them, may be explained by this expansion of solids by heating. The part of the glass with which the hot water comes in contact is heated and expands; but the effect is quite local; the heating is confined to one spot, because glass does not allow heat to pass through it readily. It is this local expansion of the glass which results in the cracking of the vessel.

It is not to be supposed, however, that substances always expand on being heated; it will be learned later that water, under certain conditions, shrinks in volume as its temperature rises. A piece of rubber, too, stretched by a weight, contracts considerably on being heated. This result is, however, deceptive. Unstretched rubber acts in the usual way and expands as its temperature rises. The explanation is that hot rubber is not so easily stretched as cold. The amount of stretch due to the weight is lessened, and masks the expansion due to the heat.

Measurement of change of temperature.—Change of temperature means change in the state of hotness or coldness of a body. The change of size which takes place when a thing is heated gives a good way of measuring the change of temperature which it undergoes. Think of the experiment with the coloured water in the flask with a long tube attached to it. Suppose the coloured water in the tube rises through a certain number of inches after the water has been heated, and that when the flask is placed into some other liquid, or some more water, the coloured water is found to rise up the tube to just the same place, we should have every right to say that the second liquid is exactly as hot as the first was. This plan provides a means of measuring temperature. The flask and tube with the water have become a "temperature measurer," that is, a **thermometer.**

32. TEMPERATURE AND THERMOMETERS.

i. The sense of feeling may be deceived.—Arrange three basins in a row; into the first put water as hot as the hand can bear, into the second put lukewarm water, and fill the third with cold water. Place the right hand into the cold water and the left into the hot, and after half a minute quickly put both into the lukewarm water. The left hand feels cold and the right hand warm while in the same water.

ii. Measuring temperature.—(*a*) Place the flask of water, with fitted tube used in Expt. 31 ii. (*a*), in hot water, and notice the height of the liquid in the tube. Transfer it to cold water, and observe that the liquid in the tube sinks.

(*b*) Procure, or make, a thermometer tube with a bulb at one end. With a little practice it is easy to blow a bulb upon a piece of thermometer tubing. One end of the tubing is held in a blow-pipe flame and twirled round until the glass melts and runs together so as to seal up the tube. A small blob of glass is then allowed to

form, and while the glass is molten the tube is taken out of the flame and blown into steadily. To introduce mercury afterwards, heat the bulb and while it is hot invert it and put the open end into mercury. As the bulb cools, mercury will rise in the tube and take the place of the air driven out of the tube by heat. Repeat the operation until the mercury fills the bulb and part of the stem.

(*c*) Place in hot water the bulb of the instrument just constructed, and make a mark at the level of the mercury in the tube. Now place the instrument in cold water, and notice that the mercury sinks in the tube. The mercury is thus seen to expand when heated and contract when cooled.

(*d*) Examine a thermometer. Notice that it is similar to the simple instrument already described, but the top is sealed up, and divisions or graduations are marked upon it, so that the height of the mercury in the tube can be seen easily. These divisions are called *degrees* (Fig. 91).

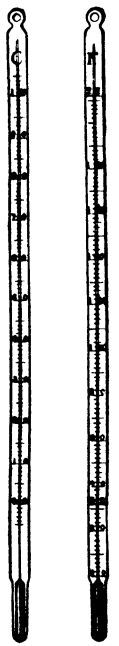


FIG. 91.—Thermometers of the form used for scientific work.

Feeling of heat and cold.—Some people feel cold at the same time that others feel warm. It is easy to understand that the sense of feeling cannot be depended upon to tell us accurately whether the air or any substance is hot or cold. Some instrument is needed which does not depend upon feeling, and cannot be deceived in the way that the senses can. Such an instrument is called a **thermometer**, and it is used to measure temperature, that is, the degree of hotness or coldness of a body.

How expansion may indicate temperature.—It has already been learned that substances usually expand when heated and contract when cooled. A flask filled with water, for instance, and having a stopper through which a glass tube passes, can be used to show the expansion produced by heat and the contraction by cold. But this flask and tube make but a very rough temperature measurer. The water does not get larger to the same amount for every equal addition of heat. Neither is it very sensitive, that is to say, it does not show very small increases in the degree of hotness or coldness, or, as the student must now learn to say, it does not record very small differences of temperature, and for a thermometer to be of value it must

do this. Then, too, as every one knows, if water is made very cold it becomes ice, which, being larger than the water from which it is made, would crack the tube. For many reasons, therefore, water is not a good thing to use in a thermometer.

Choice of things to be used in a thermometer.

1. **The substance used should expand a great deal for a small increase of temperature.**

Gases expand most, and solids least, for a given increase of temperature. Liquids occupy a middle place. The most delicate thermometers are therefore those where a gas, such as air, is the substance which expands. But in common thermometers a liquid, either quicksilver or spirits of wine, is used. Both these liquids expand a fair amount for a given increase of temperature, and, to make this amount of expansion as obvious as possible, they are used in fine threads by making them expand into a tube with a very fine bore.

2. **If a liquid is used it should not change into a solid unless cooled very much, nor into a gas unless heated very much.**

It is difficult to be sure of both these things in the same thermometer. When a thermometer is required for measuring very low temperatures it usually contains spirits of wine, because this liquid has to be cooled a great deal before it is solidified, that is, made into a solid. But this thermometer cannot be used for any great degree of temperature because spirits of wine is soon changed completely into a vapour when heated to only a comparatively small extent. If we wish to measure higher temperatures we use a quicksilver or mercury thermometer, because mercury can be warmed a good deal, or, as it is better to say, raised to a high temperature, without being changed into a gas.

3. **The liquid should be in a fine tube of equal bore with a comparatively large bulb at the end.**

Liquids have to be contained in some sort of vessel or else they cannot be kept together. There must be a fine bore, so that the liquid may appear to expand very much for a small change of temperature. The bore must be equal all the way along, that is, the width or diameter of the inside of the tube must be the same all the way along, so that a given amount of expansion in any part of the tube shall mean the same change of temperature; and, lastly, there must be a large

bulb, so that there is a large surface to take the same temperature as that of the substance the temperature of which is being measured.

Reasons why mercury is used for thermometers.—There are many reasons for selecting mercury as the liquid for an ordinary thermometer in addition to those already mentioned.

It is a liquid the level of which can be easily seen.

It does not wet the vessel in which it is contained.

It expands a considerable amount for a small increment of temperature.

It is a good conductor of heat, and consequently it very quickly assumes the temperature of the body with which it is placed in contact.

Very little heat is required to raise its temperature, and there is therefore very little loss of heat due to warming the thermometer.

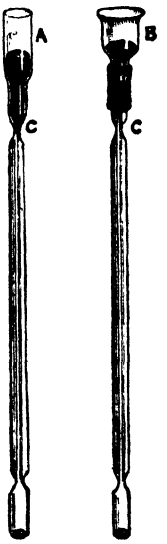


FIG. 92.—Thermometers in course of construction: (i) tube enlarged at top; (ii) tube with small funnel attached.

Construction of a thermometer.—Having selected a suitable piece of thermometer tubing, first a bulb must be blown on one end. The glass is melted at this end and allowed to run together and so close up the bore. While the glass is still molten, air is blown down the tube from the other end, the tube being moved round, so that the bulb may be placed symmetrically with reference to it. The bore of the tube is so fine that it is impossible to pour the liquid down it; some other plan must therefore be adopted. The top of the tube may be enlarged into the shape shown at *A* in Fig. 92, or a small funnel may be connected with it, as at *B*, and the liquid with which the thermometer is to be filled poured in. Let us suppose mercury is being used. Warming the tube makes the air inside it expand, and of course some is driven out. As the tube cools the mercury is forced in, by the weight of the atmosphere, to fill the place of the expelled air. By repeating this alternate

process of warming and cooling, in the circumstances described, enough mercury is soon introduced into the tube. The next step is to seal up the tube, leaving no air above the mercury; to do this the bulb is heated to a temperature slightly higher than the thermometer is expected to register. The mercury expands, and when it has reached the drawn-out part, *C*, of the tube, a blow-pipe flame is directed against the glass, and the tube is thus closed up. This method of closing a tube and keeping the air out is called **hermetically** sealing it. The thermometer at this stage should be put on one side for some days at least, in order that it may assume its final size, which it does very slowly.

33. GRADUATION AND USE OF THERMOMETERS.

i. The temperature of melting ice.—(*a*) Place some pieces of clean ice in a beaker or test-tube and push a thermometer amongst them. Notice the reading of the thermometer; it will be either *no degrees* (0°) or very near it.[1] Warm the beaker or test-tube, and observe that as long as there is any ice unmelted the reading of the thermometer remains the same.

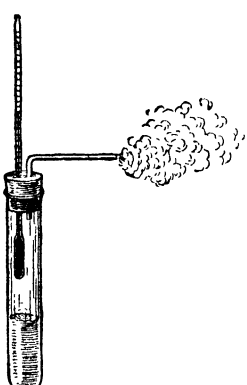


FIG. 93.—Test-tube with thermometer fitted for observing the boiling point of water.

(*b*) Repeat the experiment with pieces of some other blocks of ice, and observe the important fact that the temperature of clean melting ice is the same in all your tests.

ii. Effect of adding salt to the ice.—(*a*) Add salt to the melting ice, and notice that the thermometer indicates a lower degree of temperature.

iii. The temperature of boiling water.—(*a*) Boil some distilled water in a flask, test-tube (Fig. 93), or beaker, and hold a thermometer in the boiling water. Notice the temperature. Raise the thermometer until the bulb is just out of the water and only heated by the steam. Again record the temperature. In both cases the reading is the same. It is either *one hundred degrees* (100°), or very near it, if you use a thermometer with Centigrade divisions.

(*b*) Repeat the experiment with a second lot of pure water, and note that the temperature of boiling water is again about 100°.

[1] A Centigrade thermometer is supposed to be used. If a Fahrenheit thermometer is used the reading will be 32°.

(*c*) Add salt to the water. Hold a thermometer in the steam of the boiling water, and notice that the temperature is the same as before, namely, 100°. Push the thermometer into the water, and notice that a higher degree of temperature is indicated.

(*d*) Again place the thermometer in clean ice in a test-tube or flask. Gently heat the vessel, and notice the following changes:

(1) The mercury remains at 0° until all the ice is melted.

(2) When the ice is melted the mercury rises gradually until it reaches 100°.

(3) The mercury remains stationary at 100° until all the water is boiled away.

iv. A thermometer cannot be deceived.—Arrange three basins of cold, lukewarm, and hot water side by side. Place the thermometer in the cold water and then in the lukewarm water. Notice the temperature indicated in the lukewarm water. Now place the thermometer in the hot water, and when it has been there a minute or two put it into the lukewarm water. Notice that the temperature indicated is practically the same as before. It is thus seen that, unlike our sense of feeling, a thermometer is not deceived by being made hot or cold before using it to indicate temperature.

v. The clinical thermometer.—Examine a clinical thermometer. Notice the large space between separate degrees. Observe that just above the bulb the bore of the thermometer is constricted. Hold the bulb in your hand and watch the expansion of the mercury. On exposing the thermometer to the air and allowing it to cool, notice that the mercury thread breaks at the constriction. The mercury can be shaken back into the bulb (Fig. 97).

The fixed points on a thermometer.—In the graduation of a thermometer the plan usually adopted is to choose "two fixed points" from which to number degrees of temperature. The most convenient lower fixed point to select is the temperature at which ice melts, or water freezes, for this is always the same if the ice is pure, and remains the same so long as there is any ice left unmelted. Whenever the thermometer is put into melting ice, the mercury in it always stands at the same level; or melting ice is always at the same temperature and thus may be used to give one fixed point. The "higher fixed point" chosen is that at which pure water boils at the sea-level. This stipulation must be made, for the boiling point of a liquid is altered when the pressure upon it is changed, being raised if the pressure is greater, and lowered if the pressure is less. When the water boils, the temperature of the steam is the same as that of the water, and remains so as long as there is any water left. The lower fixed temperature is referred to as the "Freezing Point of Water," and the higher as its "Boiling Point."

Marking the freezing point. –For this purpose an arrangement like that shown in Fig. 94 is very suitable. The funnel is filled with pounded ice, which before powdering has been washed carefully: or snow may, if more convenient, be used. The glass dish catches the water which is formed from the melting of the ice or snow. A hole is made in the pounded ice by thrusting in a pencil or glass tube about the size of the

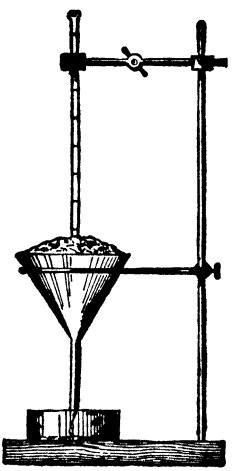


FIG. 94.—Thermometer in ice for the observation of freezing point.

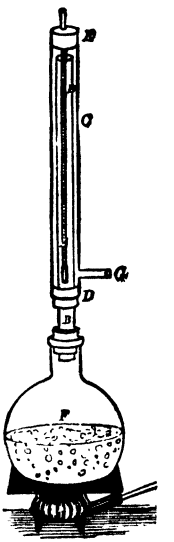


FIG. 95.—Flask fitted for the determination of boiling points.

thermometer, and into this hole the thermometer is put and is so supported that the whole of the mercury is surrounded by the ice or snow. The arrangement is left for about ten or fifteen minutes, until it is quite certain that the tube and mercury are at the same temperature as the melting ice. When this is so the tube is raised until the mercury is just above the ice, and a fine scratch is made with a three-cornered file on the tube at the level of the mercury.

Marking the boiling point.—On account of the condensation of vapour upon the thermometer, the method used in Expt. 33 iii.

to find the boiling point is not a very accurate one. More exact determinations can be made by means of the apparatus shown in Fig. 95. A can or flask *F* is fitted with a cork, through which a glass or brass tube *B* passes. Surrounding this tube is a wider tube *C*, fitted upon the inner tube by means of a piece of thick india-rubber tubing *D*. At the top of the outer tube is a cork *E* having a hole in which a thermometer can be fitted. When the water in the flask is boiled, steam passes up the inner tube *B*, and down the wide tube *C*, and escapes at the outlet *G* into the open air. To use the apparatus, the top of the stem of the thermometer being graduated is pushed gently into the cork which fits in the outer tube, and adjusted so that the point reached by the mercury at the temperature of boiling water is just above the cork. The cork is then fitted into its place, the water boiled, and when steam has been coming off for about a quarter of an hour, an observation is made of the point reached by the top of the mercury in the tube. The observation is repeated after a few minutes, and when two readings obtained at an interval of about ten minutes agree, the point at which the top of the liquid stands is marked upon the stem. The temperature observed is the boiling point of water under the particular conditions existing at the time and place of the experiment.

Precautions necessary in marking the fixed points.—Since experiments have shown that the mixture of common salt with ice or snow lowers the temperature of the ice or snow, great care must be taken that pure clean ice is used when the lower fixed point of a thermometer is being marked. It must also be noted that the presence of substances other than common salt similarly have an effect on the temperature.

The temperature of a boiling solution of common salt in water is higher, too, than that of the steam which is given off from it. Moreover, since it is the temperature of the steam from boiling water which alone remains constant at the sea-level (the nature of the containing vessel and the presence of substances in solution affecting the temperature of the liquid), in marking the higher fixed point of a thermometer the instrument should be surrounded by the steam and not placed in the liquid. It will be seen more fully later that an increase of atmospheric pressure, represented by an increased barometric height, raises the

temperature at which water boils, hence the height of the barometer must be recorded when the higher fixed point is being marked.

Thermometer scales.—Some value must now be given to the two fixed points which have been obtained as previously described, and of course they could be called anything the maker of the thermometer liked, but for the sake of comparing one man's observations and experiments with those of other people, it is most convenient to graduate all thermometers in the same way. The thermometers used in this country are divided up in two ways—(1) the Centigrade scale, (2) the Fahrenheit scale. A third scale—the Réaumur scale is extensively used in Germany.

The Centigrade scale.—Here the freezing point is called **zero** or **no degrees Centigrade**, written 0° C. The boiling point is called **one hundred degrees Centigrade**, and is written 100° C. The space between these two limits is divided into 100 parts, and each division called a **degree Centigrade.**

The Fahrenheit scale.—On thermometers marked in this way the freezing point is called **thirty-two degrees Fahrenheit,** written 32° F., and the boiling point **two hundred and twelve degrees Fahrenheit,** written 212° F. The space between the two limits is divided into 180 parts, and each division is called a **degree Fahrenheit.** The reason of this difference is interesting. The physicist Fahrenheit, after whom the thermometer is named, obtained a very low temperature, by mixing common salt with the pounded ice when marking the lower fixed point, and he wrongly imagined that he had obtained the lowest temperature which could be reached, and called it zero.

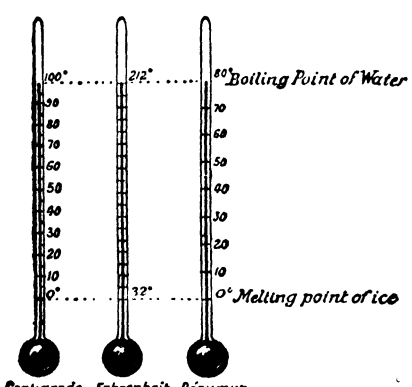


FIG. 96.—Thermometric scales.

The Réaumur scale.—Upon thermometers graduated according to this scale the freezing point is marked 0° and the boiling point 80°. The relation between the three scales is shown in Fig. 96.

The doctor's thermometer.—For the measurement of the temperature of the body, what is termed a **clinical thermometer** is best (Fig. 97). As the temperature of the living human body is never many degrees above or below a temperature of 98° F., a clinical thermometer is only graduated from about 95° F. to 110° F. If the bulb of such a thermometer is put into the mouth, or under the armpit, of a person in health, and left there for two or three minutes, it will be found, on taking it out, to indicate a temperature from 97°·8 F. to 98°·6 F. The thread of mercury in the stem of the thermometer remains in one position, though the air is cooling the mercury while the thermometer is being read. This is because of the constriction at the top of the bulb, which causes the thread of mercury in the stem to be left behind while the mercury in the bulb contracts. To "set" the thermometer for a fresh observation, it is only necessary to jerk it slightly, when the thread of mercury will again join up to the liquid in the bulb.

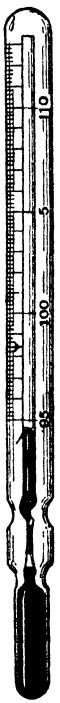


FIG. 97.—A clinical thermometer.

34. COEFFICIENTS OF EXPANSION.

1. Coefficient of expansion of a solid.—Procure and examine the apparatus shown in Fig. 98, if necessary, fitting the parts together. Notice that *AB* is a glass, or metal, rod resting in a groove at *A*, against a weight *W*, and rolling on a needle on a glass bed at *B*. Fix a split straw pointer to the needle to move against the graduated quadrant *E*. Observe that *CD* is a tube of wide bore fixed on the rod by means of corks and having an inlet for steam at *C* and an escape at *D*. When the apparatus is in adjustment take the temperature of the room in the neighbourhood of the apparatus, and then pass steam through *CD* for 5 or 10 minutes. Note what part of a whole turn the pointer moves through. Obtain the diameter of the needle by placing several similar needles in a row, measuring the

total breadth and dividing by the number of needles. Obtain the

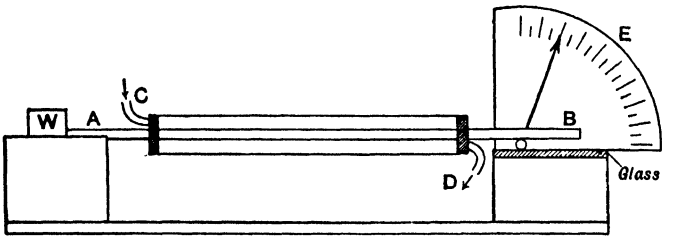


FIG. 98.—Apparatus to show the expansion of a rod.

actual movement of the end *B* of the rod by multiplying together the fraction of a rotation shown by the index and the circumference of the needle. The latter is 3¼th times the diameter of the needle. Assuming the steam to be at 100° C. obtain the rise of temperature of the rod.

Then calculate thus:

A length of..... cm. of the metal or glass rod when raised.....degrees expanded..... cm.

Therefore 1 cm. of the rod when raised 1 degree expands.....cm.

The value so obtained is the coefficient of expansion of the material of the rod used.



FIG. 99.—Apparatus to show the expansion of liquids.

ii. Coefficient of expansion of liquids.—(*a*) Close one end of a glass tube about 30 cm. long and 3 mm. bore. Partly fill the tube with water, and fasten it to a thermometer by means of threads or india-rubber bands (Fig. 99). Place the combination in melting ice, so that the water is surrounded by the ice, and observe the degree on the thermometer, level with the surface of the water in the tube. Repeat the operation with the combination successively in water at 50°, 60°, 70°, 80°, and 90°, taking care that the water in the tube is completely immersed. Now take the combination out of water, and measure the distance from the bottom of the tube to the point at which the surface of the water stood in each case, taking care that the tube does not move. Record the observations thus:

TEMPERATURE.	LENGTH OF WATER COLUMN.	INCREASE OF TEMPERATURE.	INCREASE OF LENGTH.

Find from these results, the average increase of length for 1° rise of temperature, and the fraction which this increase is of the original length.

As the tube is uniform in bore the lengths of the column of water are proportional to the volume of the water, so that the results will show the increase of the volume of water for a rise of temperature of 1°, expressed as a fraction of the original volume.

(*b*) Repeat the preceding experiment, using turpentine, alcohol, or mercury instead of water in the tube, and find in the same way the fraction of its volume at 0° by which the liquid expands for a rise of temperature of 1° C.

iii. Coefficient of expansion of a gas.—Obtain a piece of thermometer tubing of about 1 mm. bore and 20 cm. long. Suck into it a length of about 1 cm. of mercury. Seal one end of the tube and arrange that the index of mercury comes near the middle of the tube when the end has been closed and the tube is cool. Fasten the tube to a thermometer, closed end downwards (Fig. 100). You have in it a certain volume of air, and can find the volume at different temperatures as you did with liquids. Place the combined thermometer and tube in melting ice and notice the position of the air column with reference to the thermometer scale. Repeat the operation for every 10° up to 100° C., taking care that the air column is immersed completely in each case, and giving the tube two or three taps before making an observation, in order to make sure that the mercury is not sticking to the tube. Record your observations thus:

Temperature.	Length of Air Column.	Expansion for 10° C.	Average Expansion for 1° C.

As the tube is cylindrical and uniform in bore, the volume of the air in it is proportional to the lengths of the air column. The average increase of volume for 1° C., expressed as a fraction of the volume at 0° C., is the *coefficient of expansion.* Find from your results the coefficient of expansion of air.

When a gas is heated in circumstances where, as in these experiments, free expansion is possible, it is said to expand *under a constant pressure.* Both at the beginning of the experiment and after the gas has been heated, the pressure to which it is subjected is simply that of the atmosphere.

Measurement of expansion.—While a definite rise of temperature causes most bodies to expand, the amount of the expansion varies within wide limits. In the case of certain

special alloys it is almost negligible, while gases expand more than double their volume on being heated from 0° C. to 300° C.

Having now a means at command of measuring temperatures, a more exact comparison of expansions becomes possible. The **coefficient of expansion** is the term used for the rate of expansion when the temperature is increased. In the case of solids we are usually concerned with the **linear** coefficient of expansion, while for liquids and gases the coefficient of **cubical** expansion is of importance.

The fraction of its length, at 0° C., which a body expands when its temperature is raised through 1° C. is called its coefficient of linear expansion.

The actual expansion being small in the case of solids, it is not necessary to measure the length of a solid at 0° C. A simpler definition is sufficiently accurate. It may be said that the fraction of its length which a solid body expands on being heated 1° C. is called its *coefficient of linear expansion.*

It must, however, be remembered that in the case of gases, owing to the large changes that take place, the strict words of the definition must be adhered to and the expansion compared with the volume at 0° C.

The fraction of its volume, at 0° C., which a body expands when its temperature is raised through 1° C. is called its coefficient of cubical expansion.

Coefficient of linear expansion.—To obtain a measurement of the linear expansion of a rod when heated to a known temperature the form of apparatus shown in Fig. 98 may be used. A rod of glass or metal about 18 inches long is surrounded by a glass tube having an inlet for steam at *C* and an outlet at *D*. The end *A* of the rod rests in a V-shaped groove and against a weight *W*. The other end rests on a needle which is free to roll on a glass base. To the needle is attached a cork having a split straw pointer. Any movement of the needle will be observed against the scale *E*.

When steam is passed through the tube it heats the bar, and the expansion shows itself at *B*—the end *A* being fixed—by causing the needle to roll. To secure good contact the rod may be roughened with emery where it rests on the needle, and pressed to it by means of an elastic band attached to the support.

After the steam has circulated for 5 or 10 minutes, the part of a complete circle the index has moved through is observed. This shows what part of a complete rotation the needle has made. To find the horizontal expansion of the rod corresponding to this rotation it is necessary to determine the diameter of the needle. To do this several similar needles are placed in a row, the total breadth is measured and divided by the number of needles.

The circumference of a circle being equal to its

$$\text{diameter} \times \frac{22}{7}$$

the distance the needle would roll in one complete rotation would be

$$\text{diameter of needle} \times \frac{22}{7}.$$

The actual distance rolled, corresponding to the expansion of the rod,

$$= \begin{matrix}\text{distance corresponding to a} \\ \text{complete rotation}\end{matrix} \times \begin{matrix}\text{the part of a complete rotation} \\ \text{shown by the index.}\end{matrix}$$

Suppose the measured expansion to be E, the length of the rod up to the needle L, and its temperature at the commencement of the experiment 15° C. The fraction of its original length which the rod expands will be $\frac{E}{L}$.

The rise of temperature having been from 15° C. to 100° C., that is 75° C. the fraction of its original length which the rod expands for *each degree of temperature* will be $\frac{E}{L \times 75}$.

This is the **coefficient of linear expansion.**

Coefficient of cubical expansion of a liquid.—The rate of expansion of a liquid with rise of temperature can be obtained with the apparatus shown in Fig. 99. It consists of a glass tube open at one end, and about 30 cm. in length, and with a bore of about 3 mm. It is partly filled with the liquid, the expansion of which is to be determined. The tube, fastened to a thermometer, as shown in Fig. 99, is immersed in a bath of water and observations made from 0° C. to nearly the boiling point of water. The temperature of the bath is obtained from the thermometer, the scale of which also serves for observations of the liquid level. The original length of the column being

measured, and its expansion for a definite rise of temperature obtained as above, all the data exist for the calculation of the coefficient of expansion. It should be carefully noted that it is cubical expansion that is observed. Neglecting the expansion of the glass, the *increased volume* of the liquid due to heating shows itself as an *increase of length* of the liquid column.

Real and apparent expansion of liquids.—Hitherto the expansion of the glass containing a liquid has not been considered. But the glass, like most substances, also expands when heated, though it may not be noticed, as the expansion of the liquid is so much more. The momentary sinking of the liquid levels when the flasks were first placed in the warm water in Expt. 31 ii. (*b*), was caused in each case by the expansion of the flasks. The flasks becoming warm first, increased in size. The liquid level fell. As soon, however, as the contained liquid began to get warm, with its greater rate of expansion, it outstripped the expansion of the glass and the liquid level rose again. The amount which a liquid expands appears less than it actually is because of the expansion of the vessel. This expansion which it appears to possess is called its **apparent expansion.** To obtain the real expansion of a liquid it is necessary to add to the apparent expansion the amount the glass expands, or,

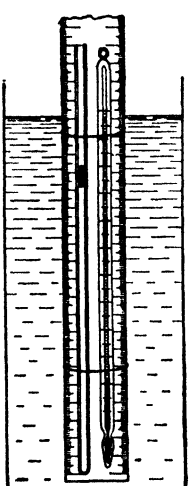


FIG. 100.—Determination of co-efficient of expansion of a gas.

real expansion of a liquid = its apparent expansion + expansion of the glass.

If any two of these values are known, the third can evidently be easily calculated.

It will be seen that in the case of a thermometer and in Expts. 32 and 33 apparent expansion only is observed.

Expansion of gases.—The expansion of gases is very much greater than that of either solids or liquids. Thus, 273 c.c. of dry air at 0° C. become 274 c.c. at 1° C. and 373 c.c. at 100° C. The coefficient of expansion is consequently $\frac{1}{273}$; and this is practically the coefficient for all gases. This law is not, however, strictly obeyed by all

gases, though air and some other gases conform to it with great accuracy. The air thermometer is often used as a standard, since the expansion of air with increase of temperature is both regular and large.

The coefficient of cubical expansion of a gas can be determined by enclosing a definite quantity of air in a long narrow tube closed at one end, the gas being separated from the air by a tiny thread of mercury (Fig. 100). The length of the column of air represents the original volume of the gas. The expansion the air undergoes will push the index outwards, and these movements can be observed. If temperature observations are simultaneously taken, the expansion per unit volume per degree of temperature can be obtained.

Coefficients of linear expansion of solids.

	Expansion Per Degree C.		Expansion Per Degree C.
Brass,	0·000019	Iron,	0·000012
Copper,	0·000017	Platinum,	0·000009
Glass (tube),	0·000008	Zinc,	0·000029

Coefficients of cubical expansion of liquids.

Alcohol,	0·00109	Olive oil,	0·00068
Glycerin,	0·00053	Turpentine,	0·00105
Mercury,	0·00018	Petroleum,	0·00099

Coefficient of expansion of gases.

	Increase of Volume at Constant Pressure.
Hydrogen,	0·00366
Air,	0·00367
Carbon dioxide,	0·00371

CHIEF POINTS OF CHAPTER IX.

Effects of heat.—(1) Change of size. (2) Change of state. (3) Change of temperature. Change of size is known either as *expansion* or *contraction*. The former is generally brought about by heating, the latter by cooling,

Expansion and contraction by heat must be taken into account in (*a*) laying railway lines, (*b*) fixing steam or hot-water pipes, (*c*) building bridges. The effects are utilised in fixing iron tyres to wheels.

Choice of materials to be used in a thermometer.—(1) The substance used should expand a great deal for a small increase of temperature. (2) If a liquid is used it should not change into a solid unless cooled very much, or into a gas unless heated very much. (3) The liquid should be in a fine tube with a comparatively large bulb at the end.

Reasons why mercury is used for thermometers.—In addition to the reasons above, (*a*) its level can be easily seen ; (*b*) it does not wet the vessel in which it is contained ; (*c*) it is a good conductor of heat ; (*d*) very little heat is required to raise its temperature.

The fixed points on a thermometer.—(1) The temperature at which ice melts or water freezes ; (2) the temperature of the steam issuing from boiling water when the barometer stands at 30 inches.

Thermometric scales.—The distance between the freezing and boiling points on the stem of a thermometer may be divided as follows :

	Fahrenheit scale.	Centigrade scale.	Réaumur scale.
Boiling point, - -	212°	100°	80°
Freezing point, - -	32°	0°	0°

Coefficients of expansion.—The fraction of its length which a body expands on being heated through 1° C. is called its *coefficient of linear expansion.*

The fraction of its volume, measured at 0° C., which a body expands on being heated 1° C. is called, its *coefficient of cubical expansion.*

When the total expansion for a large change of temperature is observed, the average change for each degree is called the mean expansion between the extreme temperatures, and the average coefficient the mean coefficient.

EXERCISES ON CHAPTER IX.

1. A flask containing pure water is heated by a single burner and one thermometer is placed with its bulb below the surface of the water, and another thermometer with its bulb just above the surface. When the water boils the readings of the two thermometers are taken. Will the readings be the same ?

What will be the effect on the reading of each thermometer (1) of placing a second burner under the flask, and (2) of dropping some common salt into the flask ?

2. Describe carefully the manner in which the freezing and boiling points on a thermometer are determined.

3. Take a glass tube open at one end and having a bulb at the other. Hold the tube so that the open end dips into water. Heat the bulb gently with a spirit lamp for a minute or two, and then take the lamp away. What will be observed ? How can you account for the facts observed ?

4. Mention conditions that should be satisfied by the bulb and tube of a mercury thermometer. Give the reason for each condition that you state.

5. I take two equal flasks, the mouths of which are fitted with bored corks carrying long glass tubes, and fill one with water coloured blue, and the other with methylated spirits coloured red; I then plunge them both into boiling water. Explain what will take place, giving reasons.

6. Describe carefully how to construct a common thermometer.

7. What is meant by *coefficient of expansion?* Give a method of finding it in the case of (*a*) a solid bar, (*b*) a liquid.

8. One-fifth of a bottle is filled with cold water; the bottle is tightly corked; the cork is pierced by a bent tube, one end of which dips into the water of the bottle and the other into water standing in an open vessel. Describe the results that may be observed if the bottle and its contents are heated up to a temperature of 99° Centigrade, and then allowed to cool.

9. A nurse cleanses a doctor's thermometer which reads up to 105° F. in boiling-hot water. The doctor now finds that the thermometer is useless. Why is this?

CHAPTER X.

CHANGE OF STATE, MELTING POINT, BOILING POINT, VAPOUR.

Change of state.—Substances exist in three states, namely, solid, liquid, and gaseous. By the action of heat a substance may be changed from one state to another. Wax, for instance, is usually a solid, but by heating it becomes a liquid. Butter can in the same way easily have its state altered from solid to liquid. Lead and zinc are also melted when heated, but at a higher temperature than wax or butter.

A good example of the changes of state produced by heat is obtained by heating a piece of ice until it becomes water, and then heating the water until it passes off into steam or water vapour. Here the same form of matter is by heat made to assume three states; in other words, ice, water, and steam are the same form of matter in the solid, liquid, and gaseous state respectively.

Change of state includes changes in the physical condition known as liquefaction or becoming liquid, and vaporisation or becoming converted into vapour. Thus, if we heat ice it is first liquefied or becomes water, and is then vaporised or becomes steam.

35. LIQUEFACTION.

1. **Melting point of wax.**—Melt a little paraffin wax in a beaker, and immerse the bulb of a thermometer in the liquid. When the thermometer is taken out, a thin film of liquid paraffin will be seen upon it. Let the bulb cool, and notice the temperature when the wax assumes a frosted appearance, which shows that it is solidify-

ing. When the wax on the bulb has become solid, place the thermometer in a beaker of water and gently heat the water. Observe the temperature at which the wax becomes transparent again. The average of this result and the preceding one is the *melting point* of paraffin wax.

ii. Melting point of butter.—Place a little butter in a test-tube, and stand a thermometer in it. Place the test-tube in a beaker of water being gently heated on a sand-bath. Notice the temperature at which the butter melts. Take out the test-tube when the butter has all melted and let it cool. Notice the temperature at which it solidifies.

iii. Melting point of ice.—Put some small pieces or shavings of clean ice into a beaker and insert a thermometer into them. Record the temperature indicated. Pour in a little water, stir the mixture, and again record the temperature. Place the beaker on a sand-bath and warm it gently. Notice the reading of the thermometer *so long as there is any ice unmelted.* In all these cases the reading of the thermometer is practically the same, indicating that the temperature of melting ice is constant.

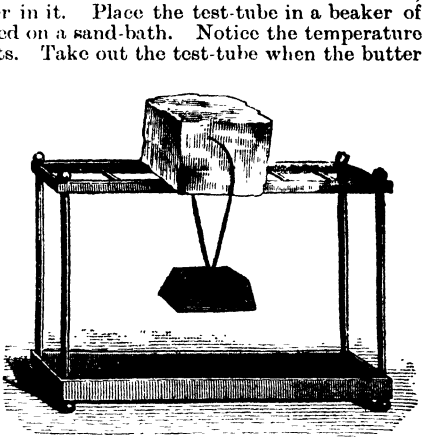


FIG. 101.—An experiment to illustrate regelation.

iv. Regelation.—(*a*) Press together two pieces of ice under water. Notice that they stick together.

(*b*) Support a block of ice as in Fig. 101. Pass over the block a loop of copper wire. Hang on to the wire a 56 lb. weight. Observe that the wire cuts its way through the ice and that the melted ice freezes again behind it.

Temperature of melting.—When a solid is heated, the first effect is usually an increase of size. But if the heating is continued long enough, when the solid reaches a certain temperature, which differs with different solids, melting begins. The solid changes into a liquid. The temperature at which the melting takes place is called the **melting point.** Thus, when a lump of lead is heated, its temperature rises, it gets larger, and as the heating is continued it is converted into a silvery-looking liquid. Wax, ice, and iron are other examples of solids which **melt.** But ice, wax, lead, and iron differ very widely in the

temperatures at which they begin to melt, as the following table shows:

Ice	melts at	0° C.
Bees-wax	„	62° C.
Lead	„	330° C.
Cast-iron	„	1200° C.

So long as any of the solid remains unmelted, the temperature does not rise above the melting point. You can easily satisfy yourself that this is true in the case of ice. If you obtain some small pieces of clean ice, and thrust a Centigrade thermometer into them, you will notice that the thermometer records a temperature of 0° C. Or, if you put some of the ice into a beaker, and pour in some water, you will find after you have stirred the ice and water together for a little while, provided you have put enough ice to be sure that it does not all melt, that the thermometer still records a temperature of 0° C. Even if you put the beaker, with the ice and water in it, over a laboratory burner and warm it gently, you will still find that, so long as there is any ice unmelted, the thermometer still reads 0° C. It is evident, then, that the temperature of melting ice is always the same, and remains the same so long as there is any ice unmelted. Moreover, to change the state of a solid, without changing its temperature, heat is expended.

Regelation.—If two pieces of ice, near the melting point, are pressed together they adhere. The pressure at the points of contact causes the melting point to fall and the ice in the neighbourhood melts. On releasing the pressure freezing again takes place, and the masses are joined together. This phenomenon is called **regelation.** It is in consequence of this that glaciers—rivers of solid ice—curve and flow like water. Fig. 101 shows an interesting example of regelation.

36. VAPORISATION.

1. Cooling produced by evaporation.—(*a*) Sprinkle a few drops of spirits of wine, sal-volatile, or ether on your hand. Notice that the liquid soon disappears, and that its presence in the air can be detected by its smell. The rate at which the liquid evaporates is increased by waving the hand about. Notice that the hand feels cold.

(*b*) Pour a few drops of water upon a dry piece of thin wood, and stand in the water a thin beaker containing a little ether. Blow vigorously down a tube having one end in the ether (Fig. 102), or use a pair of bellows. The ether rapidly evaporates, and *in doing so takes heat from the water* between the beaker and piece of wood. The beaker and wood become frozen together.

(*c*) Heat and boil water in a flask. By means of a thermometer observe the gradual rise of temperature until the water boils. Take the temperature at intervals while the water is boiling. Notice that it remains constant, though heat is being added all the time.

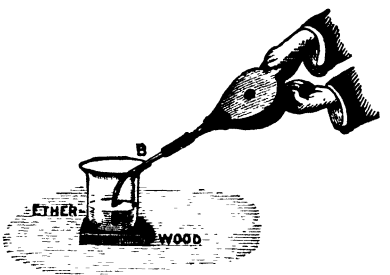


FIG. 102.—Experiment to show that water may be frozen by the rapid evaporation of ether close to it.

Heat is required to change a liquid into a vapour.—When a liquid is changed into vapour a certain amount of heat is used up. It does not matter whether the liquid evaporates or boils; every gram of it requires a certain amount of heat before it becomes converted into vapour. In boiling, this heat is supplied by the flame or fire, and in evaporation it is taken from the objects in contact with the liquid. The faster the evaporation the more quickly is heat absorbed in this way. When a liquid evaporates very rapidly, the cooling produced is very noticeable. For instance, if a few drops of either spirits of wine or ether are sprinkled upon the hand, the liquid soon disappears and the hand feels cold. The heat necessary for the evaporation of these liquids is taken from the hand, or from any other things with which they are in contact, consequently the hand becomes cooler and cooler as the vapour is formed. So much heat may be absorbed in this way that, as illustrated by the experiment shown in Fig. 102, water can be frozen by the evaporation of ether in a vessel in contact with it.

In tropical countries, where the land gets *very hot* during the day, evaporation takes place so rapidly after sunset that the water sometimes becomes so much cooled, by the extraction of the heat required to bring about the change from liquid to vapour, that the water freezes.

You will probably yourself have noticed that not only is the dust laid by watering the roads in summer, but the air is pleasantly cooled by the evaporation of the water.

It has been shown by several different experiments that when once water has started to boil, its temperature gets no higher than the boiling point. So long as there is any water left, no matter how much you heat it, its temperature remains the same.

37. BOILING POINTS.

i. Determination of boiling points.—(*a*) Put a little methylated spirit in a test-tube, and gradually heat it in a beaker of water until the spirit boils. Find the temperature of the boiling spirit and of the vapour, and record the results.

(*b*) A convenient arrangement for determining the boiling point of a liquid is shown in Fig. 95. A can or flask *F* is fitted with a cork, through which a glass or brass tube *B* passes. Surrounding this tube is a wider tube *C*, fitted upon the inner tube by means of a piece of thick india-rubber tubing *D*. At the top of the outer tube is a cork *E* having a hole in which a thermometer can be fitted. When the water in the flask is boiled, steam passes up the inner tube *B*, and down the wide tube *C*, and escapes at the outlet *G* into the open air.

(*c*) To use the apparatus, gently push the top of the stem of the thermometer into the cork which fits in the outer tube, adjusting it so that the 100° point is just below the cork. Fit the cork in its place, boil the water, and when steam has been coming off for about a quarter of an hour, raise the cork and read the thermometer. Repeat the observation after a few minutes, and when two readings obtained at an interval of about ten minutes agree, record the observation. Determine the boiling points of water, turpentine, milk, beer, vinegar and whisky.

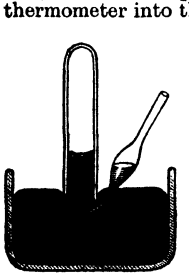


FIG. 103.— Curved pipette to introduce water into barometer tube.

ii. Vapour pressure.—(*a*) Fill a long tube with mercury and invert it in a vessel of mercury (Fig. 105). Keep this tube to show the pressure of the atmosphere. Pass up a second barometer tube three or four drops of water by means of a curved pipette as shown in Fig. 103. Observe the depression of the mercury column and that the water has evaporated. Pass a few more drops of water up the tube. Note that no further evaporation takes place and no further depression of the mercury.

Repeat the experiment with the other tubes shown in Fig. 105, using alcohol and ether. Tabulate your results.

Liquid used.	Water.	Alcohol.	Ether.
Depression of mercury column.			
Temperature.			

(*b*) Introduce some mercury into the bent tube (*A*, Fig. 104). Pour a little alcohol into the longer tube, and by tilting the apparatus allow the alcohol to creep round the bend until some is introduced

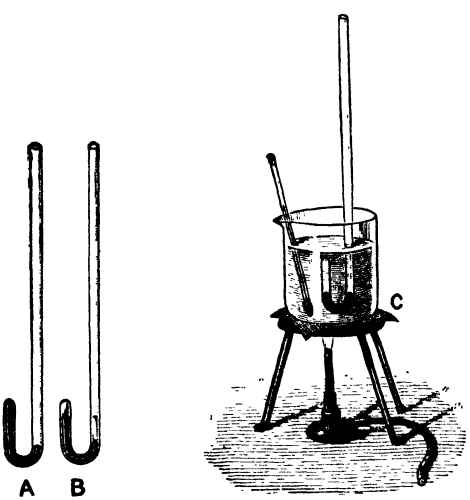


FIG. 104.—Determination of boiling point of alcohol.

into the shorter tube (*B*, Fig. 104). Place the apparatus in a beaker of water with a thermometer. Warm the water and read the thermometer at the instant the mercury stands at the same level in both arms of the apparatus. The reading of the thermometer is the boiling point of alcohol.

Vapour pressure and boiling point.—The apparatus represented in Fig. 95 and described in Expt. 37 i. (*b*) is used in the determination of boiling points. The thermometer is immersed in the vapour from the boiling liquid and is protected from cooling by an outer chamber through which is flowing the escaping vapour. When the temperature as indicated by the thermometer has become stationary, its reading is taken as the boiling point of the liquid. At this temperature the vapour pressure is equal to that of the atmosphere.

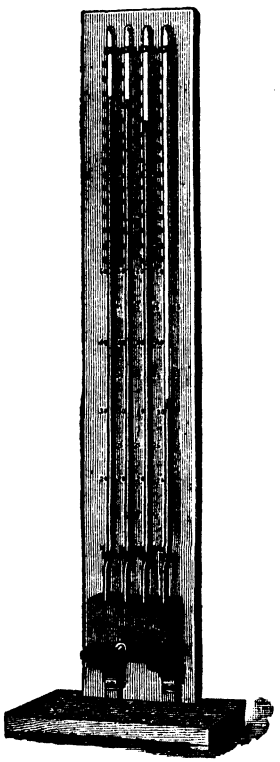


FIG. 105.—Determination of vapour pressure.

If a liquid is introduced into a vacuum it rapidly evaporates. A limit is however reached when, with the vapour in contact with its own liquid, evaporation ceases. The space is **saturated** with the vapour. That the vapour exerts a definite pressure can be shown by the apparatus represented in Fig. 105. The first tube on the left is an ordinary barometer tube, the remaining three have had introduced into them respectively water, alcohol, and ether. The water having evaporated into the Torricellian vacuum causes but a slight fall of the mercury column. The alcohol and the ether exert an increased pressure. The depression of each mercury column measures the pressure of each vapour at the temperature of the experiment. If the liquids and vapours in the tubes are warmed their pressures increase, and the mercury level falls. When each has reached its boiling point the mercury level outside and inside the tube is the same. The

ether will first reach this stage, its boiling point being lowest.

This method of obtaining boiling points, usually employed for liquids which boil at a lower temperature than water, is more conveniently used in the apparatus shown in Fig. 104. The liquid, the boiling point of which is required, is introduced into the shorter arm of the bent tube. The apparatus is heated in a water bath until the mercury level is the same in both arms. The vapour pressure is now equal to that of the atmosphere. The reading of the thermometer gives the **boiling point** required.

38. EFFECT OF PRESSURE ON THE BOILING POINT.

1. **Water boiling under diminished pressure.**—Boil some water in a flask, and let it continue to boil for some minutes until you can be sure all the air is driven out of the flask. Remove the burner and quickly insert a well-fitting cork. Allow the flask to cool for a few minutes, then turn it upside down on a suitable support and throw cold water on to the flask. Notice the water again starts boiling vigorously.

Water boils at a lower temperature under diminished pressure.—Pressure has a great influence on the boiling point of a liquid. The weight of the atmosphere at the surface of the earth is, it will be remembered, equal to that of a weight of 15 lbs. on every square inch. In studying the pressure of the atmosphere it was seen that its amount upon an object depends upon the extent of the air above the object. This pressure will be less at the top of a mountain than at the bottom of a mine, and consequently the pressure of the air will be in the former situation less than in the latter. If we wish to boil a liquid, therefore, in those cases where the pressure of the atmosphere is great, we shall have to heat the liquid more before the bubbles of vapour formed can escape at the surface of the liquid than when the pressure is less. If we heat the liquid more, its temperature will get higher before there is any conversion into vapour, and consequently its boiling point will be higher when the pressure is greater. In finding the boiling point of a liquid we must therefore know the pressure of the atmosphere at that time and place.

Example of water boiling at a lower temperature under diminished pressure.—A very simple experiment convinces us that water may boil at a temperature considerably below 100° C. when the pressure upon its surface is diminished. All it is necessary to do is to take a sound cork which tightly fits the neck of a round-bottomed flask. We then boil some water in the flask and allow it to continue boiling for some minutes so that all the air in the flask is driven out, and its place taken by steam. The burner is then removed and the cork inserted into the neck of the flask as rapidly as possible. After standing to cool for a minute or two, when, owing to cooling, the temperature can no longer be 100° C. the flask is turned over and cold water poured upon its upturned under surface or a cold wet sponge is squeezed upon it as shown in Fig. 106. The cold water causes the steam in the flask to condense, and, as no air can get in, the pressure on the surface of the warm water is now less than it was before, and therefore the water is seen to boil quite briskly again.

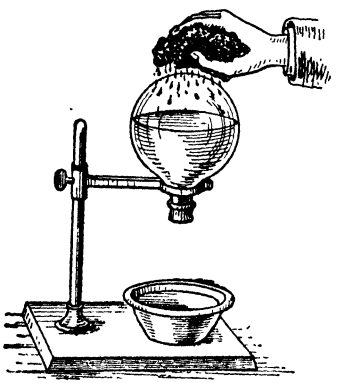


FIG. 106.—Water below 100° boiling under diminished pressure.

39. WATER DOES NOT ALWAYS EXPAND WHEN HEATED.

1. **Anomalous expansion of water.**—Procure or construct an apparatus of the form shown in Fig. 107. The cylindrical bulb should be about 10 cm. long and 1·5 cm. in diameter, and terminate in a capillary tube having a bore of about 0·5 mm. Heat the bulb; and let it cool with the open end of the tube inverted in mercury. By this means sufficient mercury (*M*) may be introduced to occupy about one-seventh the volume of the bulb. Then introduce in a similar way sufficient boiled distilled water to fill the remainder of

the bulb and part of the stem. A short length of oil may also be introduced to prevent absorption of air and evaporation. A paper millimetre scale is then attached to the capillary tube.

Support this apparatus in a wide test-tube containing mercury, so as to secure uniformity of temperature. Place a thermometer in the mercury, and support the wide tube containing it and the apparatus in a beaker of cold water. Notice the position of the top of the liquid in the tube, and read the temperature shown by the thermometer. Add ice to the water, and as the temperature falls notice the level of the liquid in the tube for every degree down to 1° or 2° C.

Then let the water in the beaker gradually rise in temperature, adding a little warm water, if necessary, and again observe the positions at the same temperatures as before. The mean of the two positions observed for each temperature should be taken as the true reading for that particular temperature. By means of squared or section paper, it is easy to construct a curve to represent graphically the observations of the changes of volume of water at temperatures near the freezing point.

The observations, with or without the curve constructed from them, will show the temperature at which the water in the apparatus has the smallest volume, and therefore the maximum density.

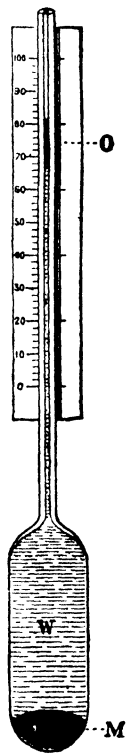


FIG. 107.— Apparatus for the determination of the changes of volume of water near the freezing point.

Changes in volume and density as water is cooled.—It has already been learnt that if the volume of a body gets greater while its weight remains the same, what is called the density of the body must get less and less. It is quite clear that, if the same amount of matter occupies a larger space, it must be less closely packed into that space, and it is the closeness with which matter is packed into a space which we have learnt to call density. What changes in density take place when water is gradually cooled? Since it has been proved by experiments that the same weight of water gradually gets smaller and smaller in volume as it is cooled down to 4° C., we can also express the same fact in another way and say that its density becomes greater and greater as

it is cooled down to 4° C. As from this temperature it gets larger as the cooling is continued, its density must get less and less. On the contrary, if we begin with water at 1° C. and gradually warm it, the density steadily increases up to 4° C., and from that temperature upwards the density regularly diminishes. Because any weight of water has a smaller volume at 4° C. than at any other temperature, or, what is the same thing, has a greater density at this than any other temperature, 4° C. is referred to as the temperature at which water has its **maximum density**.

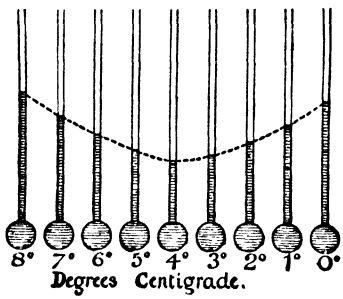


FIG. 108.—Changes in the volume of water between 0° and 8° C.

Hope's apparatus.—An experiment with what is known as Hope's apparatus shows very well that water is at its maximum density at 4° C. A cylinder provided with two side necks in the way shown in Fig. 109 is filled with water at the same temperature as the air. Into the side necks, corks with thermometers passing through them are fitted. A freezing mixture, which can be made by mixing salt with pounded ice, is applied to the middle of the cylinder. This is done by filling a vessel, fixed round the middle of the outside of the cylinder, with the mixture in a way which the illustration makes quite clear. The freezing mixture, of course, at once cools the water in the middle of the cylinder. On watching the thermometers it is found that the first effect of the cooling is to cause the temperature of the lower thermometer to fall.

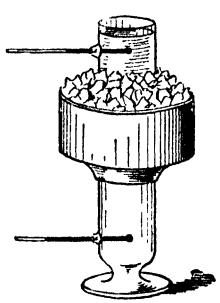


FIG. 109.—Hope's apparatus for the observation of the temperature of maximum density of water.

The temperature of the upper thermometer, however, remains unaltered. The only way in which this can be explained is by

supposing that as the water in the middle of the cylinder is cooled it gets colder and sinks to the bottom. As the cooling proceeds it is found that the water at the bottom of the cylinder *never gets below* 4° C. But soon after the water at the bottom of the cylinder has reached 4° C., the temperature of the upper thermometer begins to fall, and goes on getting lower till it actually reaches 0° C. But all this time the water at the bottom remains at 4° C. Now it is quite clear that the densest water will sink to the bottom, and as the temperature of the water there remains at 4° C. it may be concluded that water at this temperature is denser than at any other.

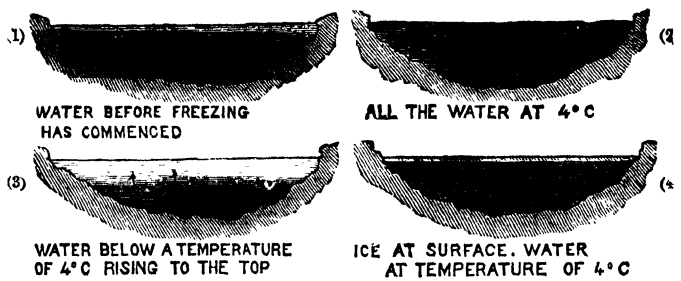


FIG. 110.—Stages in the freezing of a pond of water.

These considerations are summed up in the statement that **water at a temperature of 4° C. expands whether it is heated or cooled.**

Results in nature of the peculiar expansion of water.—From the results of an experiment with Hope's apparatus, or a consideration of the expansion and contraction of water, it is easy to understand what happens when the water of a pond is gradually cooled on a frosty night. As the temperature of the water at the surface gets lower and lower, the water there contracts and is consequently denser. It therefore sinks, and its place is taken by warmer water from below. The same cooling and sinking of the surface water continues until the temperature of the whole of the water is 4° C., at which temperature it has its maximum density, and consequently when the water at the bottom of the pond reaches this temperature it remains where it is. After the temperature of the water at the surface

it is cooled down to 4° C. As from this temperature it gets larger as the cooling is continued, its density must get less and less. On the contrary, if we begin with water at 1° C. and gradually warm it, the density steadily increases up to 4° C., and from that temperature upwards the density regularly diminishes. Because any weight of water has a smaller volume at 4° C. than at any other temperature, or, what is the same thing, has a greater density at this than any other temperature, 4° C. is referred to as the temperature at which water has its **maximum density**.

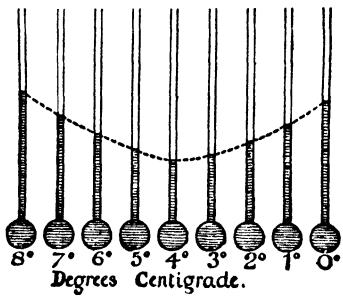


FIG. 108.—Changes in the volume of water between 0° and 8° C

Hope's apparatus.—An experiment with what is known as Hope's apparatus shows very well that water is at its maximum density at 4° C. A cylinder provided with two side necks in the way shown in Fig. 109 is filled with water at the same temperature as the air. Into the side necks, corks with thermometers passing through them are fitted. A freezing mixture, which can be made by mixing salt with pounded ice, is applied to the middle of the cylinder. This is done by filling a vessel, fixed round the middle of the outside of the cylinder, with the mixture in a way which the illustration makes quite clear. The freezing mixture, of course, at once cools the water in the middle of the cylinder. On watching the thermometers it is found that the first effect of the cooling is to cause the temperature of the lower thermometer to fall.

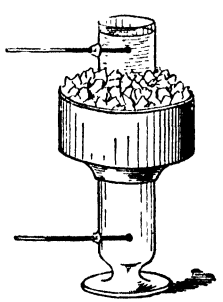


FIG. 109.—Hope's apparatus for the observation of the temperature of maximum density of water.

The temperature of the upper thermometer, however, remains unaltered. The only way in which this can be explained is by

has reached 4° C., any further cooling causes it to expand and get lighter, and this result continues until 0° C. is reached and the water at the surface is changed into ice, which, being considerably lighter than water, remains on the surface. Ice is, moreover, a very bad conductor of heat, and consequently the temperature of the water below the ice gets cooler very slowly, and the thickness of the ice increases at but a small rate.

This condition of things prevents several disastrous consequences which would of necessity follow if ice were denser than water. If ice were denser than water, it would sink to the bottom of the pond at the moment it was formed, and as the frost continued, the ice would spread throughout the mass of the water, and not only would this result in the destruction of all the water animals in the pond, but it is more than probable that the heat of summer would be insufficient to melt it completely.

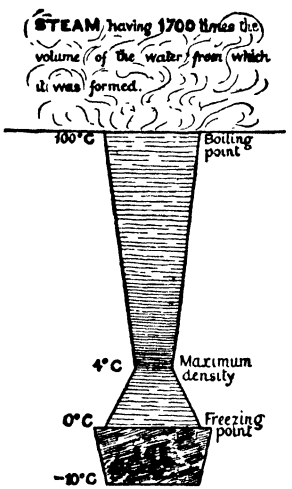


FIG. 111.—Graphic representation of the changes of volume and temperature accompanying changes of state of water.

Summary of results.—When a piece of ice at a temperature below 0° C. is warmed it expands, like most other solids, until the temperature of 0° C. is reached. At a temperature of 0° C., we know that it melts and changes into water at 0° C. While this change is taking place, though heat is continuously being added, there is no rise of temperature. The heat is used up in changing the state of the ice. After all the ice has been changed into water at 0° C., each successive addition of heat has two effects. First, the temperature is raised, and secondly, the size or volume of the water is altered. Whilst, however, the temperature rises regularly, the alterations in size are not regular; the size of the water gets smaller and smaller to begin with for every degree increase of temperature. This is continued until the temperature of 4° C. is reached, at which

temperature the water has a smaller size than at any other, or is at its **maximum density.** From 4° C. onwards the temperature and volume increase together as the water is more and more heated, and this holds true until a temperature of 100° C. is reached, when the water boils and is changed into steam. When once the water has commenced to boil, the temperature remains at 100° C., or, as it is called, the *boiling point* of water, so long as any water is left. If steam is placed in an enclosed vessel away from water, its temperature can, of course, be raised above 100° C.; in this case, what is known as *superheated steam* would be obtained.

40. FREEZING MIXTURES.

A Freezing Mixture.—Mix five parts of pounded ice with two parts of common salt in a mortar. Place a test tube partly filled with water in the mixture. After a few minutes the water will be found to freeze. Find the temperature of the mixture with a thermometer

We have seen that melting takes place at a temperature which is fixed according to the nature of the solid. The heat used to produce the melting does not raise the temperature and it is therefore called **latent** or hidden heat. If a solid melts, heat is therefore absorbed, and if the heat is not supplied by a gas flame, for instance, it is obtained from the vessel holding the solid, which will consequently fall in temperature. When salt is mixed with broken ice, the ice melts and the vessel containing the mixture, and the mixture itself, fall in temperature. The ice is caused to melt by being mixed with the salt. Such a mixture is called a **freezing mixture,** for it is used to freeze or to cool greatly substances placed in the mixture.

Examples of freezing mixtures.

Snow, or pounded ice	produce a fall of 20°C.	Snow	produce a fall of 42°C.
Salt		Hydrochloric Acid	
Snow	produce a fall of 45°C.	Sodium Sulphate	produce a fall of 28°C.
Calcium Chloride		Hydrochloric Acid	

CHIEF POINTS OF CHAPTER X.

Melting point.—The temperature at which a solid changes into a liquid is called its *melting point.*

The melting point of a solid is lowered by pressure. Two pieces of ice, if pressed together with sufficient force, become united

through melting and the re-freezing on the release of the pressure. The phenomenon is called **regelation.**

Boiling point.—When a liquid is changed into a vapour by the formation of bubbles in its interior which escape at the surface of the liquid it is said to boil. The temperature at which this takes place is called the *boiling point.*

The boiling point of a liquid is always raised by an increase of pressure.

Changes in volume as water is cooled.—Water steadily contracts as it is cooled to 4° C. If the cooling is continued below 4° C., the water begins to expand, and does so until 0° C. is reached.

The density of the water in the same circumstances increases until 4° C. is reached, from which temperature it diminishes as the cooling proceeds; 4° C. is known as the temperature of *maximum density of water.*

In its conversion into ice, water expands very much and with great force. Ice expands with a rise and contracts with a fall of temperature.

Freezing mixtures.—Certain solids when mixed together suffer a great fall in temperature. This fall is due to the absorption of heat by the mixture on liquefying.

EXERCISES ON CHAPTER X.

1. A vessel of water at the freezing point contains two small glass bulbs. One is at the bottom, the other floats, but is almost wholly below the surface. The water is gradually heated; soon the bulb that was at the bottom rises, but after a while sinks again, and remains sunk. What is the meaning of this behaviour? How will the other bulb behave during the heating of the water?

2. Describe how you would graduate a thermometer. Would any correction be necessary if you did it on the top of a mountain, or at the bottom of a coal mine?

3. What do you understand by the temperature of maximum density of water? How has the temperature been determined?

4. By means of an india-rubber tube the steam from a boiling kettle is passed into a mixture of ice and water in which a thermometer is placed. The experiment is continued for a considerable time, the mixture being kept well stirred. Describe the results which may be observed and the behaviour of the thermometer.

5. A few drops of water are poured into a flask: the flask is boiled for a minute or two over a spirit lamp, and then quickly plunged, mouth downwards, into cold water. What results may be observed? How do you account for these results? What results will be obtained if the flask is kept empty, but allowed to stand in boiling water before being plunged into the cold water?

6. Describe experiments which you have seen illustrating the conversion of (*a*) a solid into a gas, (*b*) a liquid into a solid, (*c*) a liquid into a gas. Describe what you have observed during each process. [Reference to any of the forms of water must be excluded from your answer.]

CHAPTER XI.

QUANTITY OF HEAT AND ITS MEASUREMENT; SPECIFIC HEAT: LATENT HEAT.

41. QUANTITY OF HEAT IN RELATION TO TEMPERATURE AND WEIGHT.

i. Distinction between temperature and heat.—Place a can or beaker containing water over a burner. Place in the can a small test-tube containing water. After the can has been heated for a little time, observe the temperature of the water in the test-tube and surrounding it; it will be the same. Take away the burner, and lift the test-tube out of the can. You now have a small quantity of water and a larger quantity both at the same temperature; but there is more heat in the large amount than in the small amount. Prove this by pouring the hot water from the test-tube, and that from the can, into the same quantity of cold water from the tap in separate large beakers. The large amount of hot water will thus be found to have a greater heating effect than the small amount; hence it must have possessed more heat than the small amount.

ii. Result of mixing hot and cold equal weights of the same substance.—(*a*) Put a certain weight of warm water in a beaker, and the same weight of cold water in another beaker. Observe the temperature of each by means of a thermometer. Pour the cold water into the hot. It will be found on stirring them together with the thermometer (taking care not to break the thermometer), that the temperature of the mixture is midway between the two original temperatures.

(*b*) From the observations construct a table like the one below, to show that the temperature, produced by mixing equal weights of the same liquid at different temperatures, is equal to half the sum of the temperatures:

Temperature of Water *A*.	Temperature of Water *B*.	$\frac{A+B}{2}$.	Temperature of Mixture.

iii. Equality of loss and gain of heat.—(*a*) Weigh about 200 gms. of cold water into a beaker, and observe its temperature. Put the same weight of water into another beaker; heat it to about 45° C Now place the beaker of hot water on your table, with a thermometer in it, and observe its temperature. When the temperature has fallen, to say 40° C., take hold of the beaker with a duster, and quickly pour the hot water into the cold. Stir up the mixture with the thermometer, and observe the temperature after mixing. Record your observations as below:

Weight of cold water, - - - - .. .. gms.
Temperature ,, - - - - ..... ° C.
,, of mixture, - - - - ..... ° C.
Number of degrees through which the temperature of the cold water was raised, - - - - - - ..... ° C.
Weight of hot water, - - - - ..... gms.
Temperature of hot water, - - - .. ... ° C.
Number of degrees through which the temperature of the hot water fell, - ..... C.

Tabulate the gain and loss of heat that occur, as shown below:

Gain.	Loss
Weight of cold water. × its rise of temperature ..... ×	Weight of hot water × its fall of temperature . . . . \ ....

The gain will be found to be slightly less than the loss. This is not really the case, and it only appears so because the amount of heat required to raise the temperature of the glass of the beaker containing the cold water has not been taken into consideration.

(*b*) Repeat the experiment, using unequal weights of hot and cold water. Notice that in each case the weight of hot water × the fall of temperature is approximately equal to the weight of cold water × the gain of temperature. The difference shows the amount of heat absorbed by the glass of the cold beaker.

The amount of heat gained by 1 gram of water when its temperature is raised 1° C., or lost when its temperature falls 1° C., is adopted as the unit quantity of heat.

Difference between heat and temperature.—Temperature is not heat; it is only a state of a body, for the body may be cold one minute and hot the next. A hot body is one at a high temperature, a cold body one at a low temperature. If a hot body and a cold body are brought into contact there is an exchange of heat until they are both of the same degree of hotness or coldness, that is, at the same temperature. Hence, temperature may be defined as **a condition or state of a body which is changed by the gain or loss of heat.**

Analogy of temperature with water level.—It is well known that if two vessels containing water and arranged at different levels are connected by means of a piece of india-rubber tubing, there is a flow of water from the vessel of water at the higher level towards the vessel at a lower level. This is a consequence of a property possessed by all liquids which makes them, as we say, *seek their own level*. This flow of water continues until the water in the two vessels is at the same level. Evidently this is a similar state of things to that which we have in the case of a hot and cold body in contact. In one case there is a flow of water until the level is the same in the two vessels. In the other there is a passage of heat until the temperature of the two bodies is the same. *Temperature corresponds to water-level.*

Temperature changes when hot and cold liquids are mixed.—Temperature may be regarded as heat-level, so that a hot substance is at a higher heat-level than a colder one. Now suppose that a certain weight of hot water is put into one vessel and an equal weight of cold water into another. We shall then have equal weights of water at different heat-levels If the two liquids are mixed together, the temperature or heat-level of the hot water will fall, and the temperature of the cold water will rise. The loss of level of one will be equal to the gain by the other, so that the temperature of the mixture will be midway between the two original temperatures Thus, if the weights of water are equal, and the temperatures at first are 60° C. and 20° C., then the temperature of the mixture will be 40° C. The temperature of the hot water would fall 20° C. and the temperature of the cold water would rise 20° C.

The actual temperature of the mixture would be slightly less than the calculated temperature, because some heat would be lost while the liquids were being mixed. The loss may be regarded as a leakage of heat, and it would of course reduce the heat-level of the mixture in the same way that a leak in a water-level apparatus would cause the level after mixing to be less than it would be if the apparatus were perfect.

Quantity of heat in water at different temperatures.—Quantity of heat may be measured by heating effect, so that we can say that the quantity of heat in a certain quantity of water depends upon the *weight* of the water and its *temperature*. For

any temperature, say 60° C., we may regard the amount of heat in 100 grams of water as twice as great as in 50 grams of water, if we imagine for the sake of simplicity that water at 0° C. contains no heat. When equal or unequal weights of water at different temperatures are mixed, the quantity of heat lost by the hot water is the same as the quantity gained by the cold water. The fall of temperature multiplied by the weight of hot water is equal to the rise of temperature multiplied by the weight of cold water.

Unit quantity of heat.—Now that it has been shown that a substance may possess different quantities of heat, it is time to consider how such quantities of heat are measured. As in all other cases of measurement, a unit or standard quantity is required with which to compare quantities of heat. The unit quantity of heat generally adopted is the **amount of heat necessary to raise the temperature of one gram of water through one degree Centigrade.** This unit is called a **calorie** or **therm.** The amount of heat required to raise the temperature of 2 grams of water through 1° C. is thus 2 units or 2 calories. Similarly, if 1 gram of water at 0° C. is heated in a test-tube over a burner until its temperature is 1° C., it will have received from the burner 1 unit of heat, or 1 **calorie.** When the temperature of this 1 gram of water reaches 3° C. it will have received 3 units of heat. If the tube contains 10 grams of water at 0° C., and its temperature is raised to 12° C., it will have received 10 times 12 units of heat, the number of units being equal to weight (in grams) × increase of temperature (in degrees Centigrade).

It will thus be seen that the number of units of heat taken up by any weight of *water* as its temperature rises, or the amount given out by any weight of *water*, the temperature of which is falling, may be found by multiplying the number of grams of water used by the number of degrees, as measured by a Centigrade thermometer, through which the temperature rises or falls. This rule may be written as follows:

Number of heat-units = weight of water in grams × number of degrees Centigrade through which its temperature rises or falls.

42. QUANTITY OF HEAT IN RELATION TO SUBSTANCE AS WELL AS TO TEMPERATURE AND WEIGHT.

i. The same quantity of heat may produce different changes of temperature.—Weigh out equal quantities of water and turpentine at the same temperature in two beakers of the same size. Pour equal quantities of hot water at the same temperature into the cold water and into the turpentine. Observe the rise of temperature produced in each case. Though the equal amounts of hot water contain the same quantity of heat, the rise of temperature of the turpentine will be found to be more than the rise of temperature of the cold water; in other words, the *capacity of turpentine for heat is less than the capacity of water for heat.*

ii. Comparison of rates at which water and mercury gain heat.—Weigh out equal quantities of cold water and mercury at the same temperature in two test-tubes or flasks. Support the two vessels side by side at the same distance above a flame, or in a large beaker of boiling water. Let them remain for a few minutes; then observe their temperatures. The rise of temperature of the mercury will be found to be greater than the rise of temperature of the water; in other words, mercury gets hot more quickly than water under the same conditions.



FIG. 112.—Equal weights of water and mercury do not become hot at equal rates, though they both have the same opportunity.

iii. Different quantities of heat in equal weights of different substances at the same temperature.—Place equal weights of lead and water in test-tubes standing in the same beaker, and heat them over a laboratory burner until the water boils; the temperature of both the lead and water will then be about 100° C. Provide two beakers containing equal weights of cold water at the temperature of the room. Put the hot lead into one of these and the hot water into the other. Stir both mixtures and note the temperature in each case. The water into which the heated lead is plunged is not at so high a temperature as that into which the hot water is poured.

Equal weights of water at the same temperature are thus shown to be heated to different degrees by equal weights of lead and water at the same high temperature.

iv. Capacity for heat.—Place some iron nails in a beaker, and the same weight of cold water in another beaker. Let the two beakers stand for a while so as to assume the temperature of the room.

Boil water in a kettle or other vessel, and pour equal quantities into the two beakers. Observe the temperature of the mixture in the two beakers. The iron nails will be found to be hotter than the water in the other beaker, because iron takes less heat to raise its temperature than is required by an equal weight of water at the same temperature.

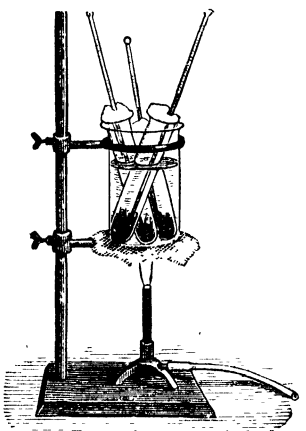


FIG. 113.—Method of heating metals in test-tubes for the determination of their capacity for heat. Each test-tube has a loose plug of cotton wool at the top.

v. Heat capacity of iron and other metals.—(*a*) Weigh out about 50 grams of cold water and observe its temperature. Put into a test-tube an equal weight of iron tacks; stand the test-tube, with a thermometer surrounded by the tacks, in a beaker of water, and boil the water (Fig. 113). Observe the temperature of the tacks, and when the water has been steadily boiling for some time, take out the thermometer and cool it under the tap. Quickly pour the heated tacks into the cold water, and observe the temperature of the mixture. Notice that it is not so high as when the hot water is added.

Comparison of heat quantities.—It has been seen that the quantity of heat in water depends upon (i.) the weight of the water, and (ii.) its temperature. It might be supposed, therefore, that as any weight of water at a certain temperature contains a certain quantity of heat, the same weight of another substance at the same temperature contains the same quantity of heat. This, however, is not the case. 100 grams of water at a temperature of 50° C. always contain 5000 units of heat,[1] but 100 grams of turpentine, mercury, lead, iron, or any other substance at the same temperature as the water, namely 50° C., do *not* contain this number of units of heat. The quantity of heat in a substance thus not only depends upon the weight and the temperature, but also upon the substance itself.

Capacity of water for heat.—Of all known substances, water has the greatest capacity for heat; consequently a larger amount of heat is required to raise the temperature of a

[1] Assuming for simplicity that water at 0° C. contains no heat.

given weight of water through a certain number of degrees than is needed by an equal weight of any other substance.

Thus, suppose a pound of water be put into one flask and a pound of mercury into another, and that these flasks are then heated for five minutes by two laboratory burners, which, as far as we can tell, give out the same quantity of heat. The temperature of the two liquids at the commencement of the experiment is, say, 15° C. If at the end of the experiment the temperature of the water was 20° C., that of the mercury would probably be about 180° C., and in order to raise the water to this temperature, if that were possible by this means, much more heat would be required. Similarly, and for the same reason, in cooling through any number of degrees of temperature a definite weight of water will give out a larger amount of heat than an equal weight of any other substance, the temperature of which falls through the same number of degrees.

Results in nature of the high capacity of water for heat.—The results in nature of this great capacity for heat which water possesses are important.

Though water takes a large amount of heat to warm it and is consequently heated by the sun's rays only slowly, yet when it cools it parts with its heat just as slowly. The effect of this on the climate of islands is very marked. The winter temperature is never very low, and the climate never very severe, because the water surrounding the country acts as a great storehouse, slowly giving up heat to the land. Similarly, the summer temperature is never unbearably hot, because the surrounding water takes so long to warm, and, being always cooler than the land, keeps the temperature of the latter from getting very hot.

Temperatures produced by mixing various hot and cold substances.—If equal weights of lead and water be heated to the same high temperature, say 100° C., and the lead be plunged into a given weight of water at a lower temperature, say 20° C., and the hot water be mixed with another equal weight of water at 20° C., and the resulting temperatures in the two cases be determined, it is found that the temperature of the weight of cold water into which the hot water was poured is higher than that of the equal weight of cold water into which the lump of lead was plunged. Hence, equal weights of lead and water at

the same high temperatures cannot give out the same amount of heat when cooled, because they contain unequal amounts. The water at 100° C. contains a larger quantity of heat than an equal weight of lead at 100° C., because its capacity for heat is greater.

Or, if 1 lb. of water at the temperature of the air be mixed with 1 lb. of iron at 100° C., the resulting temperature is not so high as that obtained by mixing 1 lb. of water at 100° C. with 1 lb. of iron at the atmospheric temperature. This evidently means that 1 lb. of water at 100° C. contains more heat than 1 lb. of iron at 100° C., or the capacity of iron for heat is less than that of water. In the same way, similar experiments with water and mercury show that the capacity of mercury for heat is less than that of water.

Comparison of capacities for heat of different metals.—When equal weights of water, tacks, copper-wire, and mercury at the same high temperature, that of boiling water, for instance, are each in turn stirred up with equal weights of cold water at the same temperature and in separate beakers, it is found that the hot water raises the temperature of the weight of cold water in which it is placed through a larger number of degrees than any of the other substances raise the temperature of the weight of water into which they are placed. This is because the capacity for heat of water is greater than that of any of these (or any other) substances.

If the temperature is observed of the mixture formed in each of the cases supposed, namely, tacks and water, copper-wire and water, and so on, and then the number of degrees through which each has raised the temperature of the water into which it was put is calculated, a series of numbers is obtained which enables a comparison to be made of the capacities for heat of each of the substances experimented with. The substances arranged in the order of their capacities for heat stand thus: (1) Iron (tacks); (2) Copper-wire; (3) Mercury; (4) Lead.

The amount of heat required to raise the temperature of one gram of a substance through 1° C. or the amount of heat given out by one gram of a substance the temperature of which falls through 1° C., in comparison with the amount of heat taken up (or given out) by an equal weight of water, **is known as the specific heat of the substance.**

43. SPECIFIC HEAT.

1. Specific heat of a solid.—Weigh out into a copper calorimeter about 30 grams of water and observe its temperature. Introduce into the test-tube of the steam-heater (Fig. 114) about 20 grams of small iron tacks. Boil the water in the heater and observe the temperature of the tacks.

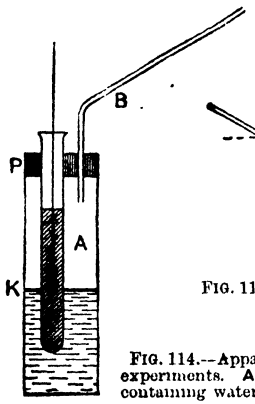


FIG. 114.—Apparatus for heating substances for specific heat experiments. A, test-tube; B, outlet for steam; P, cork; K, can containing water.

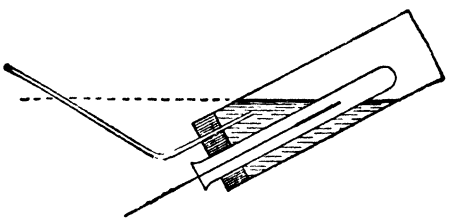


FIG. 115.—Method of pouring substances out of the steam-heater.

Take hold of the test-tube, or of the whole steam-heater, with a duster; quickly pour the tacks into the cold water, tipping the heater as shown in Fig. 115, and observe the temperature of the mixture.

Weight of water, - - - - - - - ... ...gms.
Temperature of water, - - - - - - - ..... °C.
,, mixture, - - - - - - ... ... °C.
Number of degrees through which the temperature of the water was raised, - - ..... °C.
Weight of iron tacks, - - - - - - ..... gms.
Temperature ,, - - - - - - ..... °C.
Number of degrees through which the temperature of the tacks fell, - - - - ... ... °C.
The heat given out by the tacks in cooling ..... °C. is equal to the weight of water × its rise of temperature
..... gms. × ..... °C.
..... calories.
..... grams of tacks in falling ..... °C. gave out ..... calories.
1 gram of tacks in falling ..... °C. gives out ..... calories.
1 gram of tacks falling 1° C. gives out ..... calories.

This is the specific heat required.

No account has been taken here of the heat given to the calorimeter. It is equivalent to a certain extra weight of water.

ii. The water equivalent of a calorimeter.—Determine the weight in grams of a copper calorimeter. Observe the temperature of the air and consequently of the calorimeter.

Pour into the calorimeter a convenient quantity of warm water at a temperature of from 35° C. to 40° C. Enough to one-third fill the calorimeter is a good amount. Notice with a thermometer, which you should carefully use as a stirrer, that, on pouring the warm water into the cold calorimeter, its temperature falls. When its temperature becomes stationary, which it will soon do, record the temperature again. Determine the weight of the calorimeter and water. Subtract the weight of the calorimeter, and so obtain the weight of water used.

Weight of calorimeter,	..... gms.
Temperature of calorimeter,	.....° C.
Weight of water,	..... gms.
Temperature of water,	.....° C.
Resulting temperature,	.....° C.

The exchange of heat which takes place may be considered as follows:

Weight of hot water × fall of temperature
..... × .....
..... calories.

This gives the number of heat units used in increasing the temperature of the calorimeter by an observed number of degrees. Find from the result the number of calories required to raise the temperature of the calorimeter through 1° C., that is, the water equivalent or water value of the calorimeter.

iii. Determination of the specific heat of solids.—Determine the weight of the copper calorimeter, the water equivalent of which you have already found. Pour in enough water to one-third fill it. Again weigh. Put a thermometer into the water and leave it to take the temperature of the water. When the temperature is stationary, record it. Weigh out about 50 grams of short pieces of copper wire. Heat the copper in the steam-heater provided, and record the temperature of the copper with a second thermometer. Quickly introduce the hot copper into the cold water, stir, note the rise in temperature of the water, and, when constant, record.

Set down your observations thus:

Weight of calorimeter and water,	..... gms.
,, ,, alone,	..... ,,
Therefore weight of water in calorimeter,	..... ,,
Water value of calorimeter,	..... ,,
Total water,	..... ,,
Temperature of mixture,	.....° C.
,, ,, water,	..... ,,
Therefore rise of temperature,	..... ,,

Quantity of heat gained, - - - - -	... calories.
Weight of copper, - - - - -	..... gms.
Temperature of copper before mixing, -	..... °C.
,, ,, mixture, - - - - -	..... ,,
Therefore fall of temperature, - - -	.. ... ,,

..... grams of copper in falling ... degrees gave out ..... calories gained by cold water and calorimeter ;

therefore

1 gram of copper in falling ..... degrees would give out ..... calories ;

and

1 gram of copper in falling 1° C. would give out ..... calories.

The result thus obtained is the specific heat of copper.

iv. Specific heats of liquids.—(*a*) Weigh a calorimeter. Half fill it with turpentine, and find the weight of the turpentine. Observe the temperature of the turpentine. Observe also the temperature of some boiling water. Pour boiling water into the turpentine ; keep the two liquids well stirred, and observe the temperature of the mixture. Find the weight of the water added. From these observations calculate the specific heat of turpentine.

(*b*) Determine in the same way the specific heat of mercury.

Examples of the determination of specific heats.—To obtain the specific heat of a substance, a convenient quantity of the substance is usually heated to a definite temperature and then allowed to give up its heat to a known weight of water. If losses through radiation and other causes are avoided as much as possible, the heat lost by the substance in cooling may be taken as equal to that gained by the water in having its temperature raised. The weight and rise of temperature of the water having been observed, this gain of heat can be calculated by multiplying the weight of water by its rise of temperature. The heat lost by each gram of the substance, the specific heat of which is being determined, in cooling 1° C., can then be calculated, and the result is the specific heat required. The following actual experiment will show the calculation necessary.

A weighed bunch of bronze coins were heated in a steam heater (Fig. 114) until they acquired a constant temperature near 100° C. They were then quickly dropped into a known weight of water, the temperature of which had been determined. After stirring thoroughly, the highest temperature attained by the mixture was observed.

The following are the observations recorded:

Weight of water and calorimeter,	105·15 gms.
Weight of calorimeter,	38·87 ,,
Weight of water,	66·28 ,,
Temperature of water,	16·7° C.
,, warmed water,	23·5° C.
Therefore rise of temperature,	6·8° C.
Quantity of heat gained by the water,	66·28 × 6·8 calories.
Weight of bronze coins,	67·27 gms.
Temperature of bronze before mixing,	99·8° C.
,, warmed water,	23·5° C.
Therefore fall of temperature of coins,	76·3° C.

Thus 67·27 grams of bronze in cooling 76·3° C. gave out 66·28 × 6·8 calories, which were gained by the water.

Hence 1 gram of bronze cooling 76·3° C. gave out

$$\frac{66{\cdot}28 \times 6{\cdot}8}{67{\cdot}27} \text{ calories.}$$

Therefore 1 gram of bronze cooling 1° C. gave out

$$\frac{66{\cdot}28 \times 6{\cdot}8}{67{\cdot}27 \times 76{\cdot}3} \text{ calorie} = 0{\cdot}087 \text{ calorie.}$$

This by definition is its specific heat.

Specific heat of bronze = 0·087.

Water equivalent of a calorimeter.—In this calculation, no account has been taken of the heat spent in warming the calorimeter. Its presence is equivalent to an extra quantity of water. The amount of water the calorimeter is equivalent to is called its water value. The following is an actual experiment made to determine this.

A calorimeter was taken, weighed, and placed in cotton wool in a large beaker. It was then partly filled with a weighed quantity of cold water. Into this was poured a quantity of warm water, and the whole stirred until a constant temperature was reached, which was observed.

Temperature of cold water,	14·6° C.
,, warm ,,	63·0° C.
,, mixture,	34·7° C.

Therefore fall of temperature of warm water,	28·3° C.
And rise of temperature of calorimeter and cold water,	20·1° C.
Weight of calorimeter,	38·87 gms.
,, ,, and cold water,	90·33 ,,
,, ,, and mixture,	129·76 ,,
Therefore weight of cold water alone,	51·46 ,,
And ,, warm water alone,	39·43 ,,
Therefore the heat given out by the warm water,	$39{\cdot}43 \times 28{\cdot}3$ calories.

This would raise the temperature of $39{\cdot}43 \times 28{\cdot}3$ grams of water through 1° C.

Therefore this would raise $\frac{39{\cdot}43 \times 28{\cdot}3}{20{\cdot}1} = 55{\cdot}5$ grams of water through 20·1° C.

It actually raised 51·46 grams of water through this temperature.

The calorimeter was therefore equivalent to

$$55{\cdot}5 - 51{\cdot}46 = 4{\cdot}04 \text{ grams of water.}$$

Therefore water value of the calorimeter = 4·04 grams.

This result can now be applied to improve the calculation of specific heat in the first experiment.

Weight of calorimeter and water,	105·15 gms.
,, ,, alone,	38·87 ,,
Weight of water in calorimeter,	66·28 ,,
Water value of calorimeter,	4·04 ,,
Total water,	70·32 ,,
Temperature of warmed water,	23·5° C.
,, water,	16·7° C.
Therefore rise of temperature,	6·8° C.
Quantity of heat gained,	$70{\cdot}32 \times 6{\cdot}8$ calories.
Weight of bronze coins,	67·27 gms.
Temperature of bronze before mixing,	99·8° C.
,, ,, in water,	23·5° C.
Therefore fall of temperature,	76·3° C.

67·27 grams of bronze in falling 76·3 degrees gave out $70{\cdot}32 \times 6{\cdot}8$ calories.

Therefore 1 gram of bronze in falling 76·3 degrees would give out

$$\frac{70{\cdot}32 \times 6{\cdot}8}{67{\cdot}27} \text{ calories.}$$

And 1 gram of bronze in falling 1° C. would give out

$$\frac{70{\cdot}32 \times 6{\cdot}8}{67{\cdot}27 \times 76{\cdot}3} \text{ caloric} = 0{\cdot}093 \text{ calorie.}$$

The result thus obtained is the specific heat of bronze.

44. LATENT HEAT.

i. Heat required to melt ice.—(*a*) Let a few lumps of ice stand in a beaker until some of them have melted. Notice that the temperature is 0° C. Counterpoise two empty beakers of the same size in the pans of a balance, and put a small lump of the ice into one, and the same weight of water from the melted ice in the other. You have thus equal weights of ice and water at 0° C. Pour equal weights of hot water into the two beakers. When the ice is melted, observe the temperature of the water in each beaker. The temperature of the water in the beaker in which the ice was placed will be found much lower than that of the water in the other beaker, owing to the ice using up a large quantity of the heat in melting into water.

(*b*) Take equal weights of hot water in two large beakers of the same size. Place a piece of ice in one of the beakers, and observe the temperature of the water when it has melted. Pour ice-cold water into the other beaker until the same temperature is reached. Find the weights of ice and ice-cold water which have been added. It will be found that a small weight of ice has as much cooling effect as a large weight of ice-cold water.

ii. Heat required to melt one gram of ice.—Weigh about 300 gms. of warm water into a beaker, and observe its temperature. Put a few small pieces of ice into the water, stir them round with the thermometer, and, as soon as they have melted, again observe the temperature of the water. Now make another weighing, and find out by subtraction the weight of the ice added. From these results calculate in the way shown on p. 173 the number of heat units required to melt one gram of ice.

Latent heat.—The experiments which have just been described are of the greatest importance, and should be understood clearly. It is certain that when a mixture of ice and water is heated over a laboratory burner, heat is being continually given to the mixture. Yet the temperature as recorded by the thermometer gets no higher. The question arises, what becomes of this heat, as it has no effect upon the temperature of the mixture? The ice is gradually melted, and if the heating is continued long enough it is all changed into water. As soon as this has happened, every further addition of heat raises the temperature of the water. These considerations lead to the conclusion that the heat previously given to the mixture is all used up in bringing about the change of ice into water. Further, it is found that not only in the case of ice, but when any solid is turned into a liquid, there is no increase in temperature, even while heat is being added, until the whole of the solid has been changed to a liquid.

This amount of heat which is necessary to change a solid into a liquid is spoken of as **latent heat.** The word latent comes from a Latin word, meaning "lying hidden," and refers to the fact that the heat used up in changing a solid to the liquid condition has no effect upon a thermometer, but appears to be hidden away in the liquid.

How the latent heat of water is measured.—To find out how many units of heat are required to melt a gram of ice we mix together known weights of warm water and ice, the temperature of both being known, and then record the temperature of the mixture at the instant the last piece of ice disappears. The facts which in this way become known are as follows:

i. Weight of warm water in grams.
ii. Weight of ice in grams.
iii. Temperature of warm water.
iv. Temperature of the ice.
v. Temperature of the mixture at the instant the ice finally disappears.
vi. Number of degrees through which the temperature of the water falls.

The observations can be used to determine what number of heat units have been lost by the water, and what number have been gained by the ice and by the water into which the ice is changed as it melts.

The loss of heat is at once calculated by knowing that the temperature of a certain number of grams of water has fallen through a given number of degrees, and if these numbers are multiplied together the result shows the number of units of heat lost by the warm water.

The gain of heat consists of two parts; first, the number of units of heat necessary to melt a known number of grams of ice, and this number we do not know. Secondly, the number of units of heat required to raise the temperature of the water at 0° C. (formed by melting a known number of grams of ice) up to the temperature of the mixture. This number of units of heat can be found by multiplication.

We also know that the total loss of heat and the total gain or heat are equal. Consequently, it should be plain that the

difference between the two known results thus obtained tells the number of units of heat used up in melting the ice.

Latent heat of water.—The number of units of heat which are required to change the state of a gram of ice, converting it from the solid to the liquid condition, without raising its temperature, is called the **latent heat of water**, or the **latent heat of fusion of ice.** To melt one gram of ice requires 80 heat-units. That is to say, as much heat as would raise the temperature of a gram of water through 80° C., or would raise the temperature of 80 grams of water through 1° C., is used up in changing a gram of ice into a gram of water at the same temperature. Similarly, to melt 1 lb. of ice requires as many heat-units as are necessary to raise the temperature of 1 lb. of water from 0° C. to 80° C., or as much heat as is wanted to raise the temperature of 80 lbs. of water through one degree Centigrade.

Natural consequences of the latent heat of water.—Just as it is necessary before a pound of ice can be changed into a pound of water to supply an amount of heat which would raise the temperature of a pound of water through 80° C., so before a pound of water can be changed into a pound of ice, we must take from it precisely the same amount of heat. This is why it requires several cold nights to cover a pond with ice, for not until every pound of water at the surface has had this large amount of heat taken from it can it change into ice. For just the same reason, it takes a very long time to melt completely the snow in the roads and the ice on the ponds, even after a thaw has set in.

45. HEAT ABSORBED IN THE CONVERSION OF WATER INTO STEAM.

Latent heat of steam.—Arrange a flask with the connections shown in Fig. 116. The short length of wider glass tube is a trap to catch condensed steam. Put some water into the flask and boil it. While the water is getting hot, weigh out about 300 grams of water in a beaker or a thin metal vessel, and observe its temperature. After steam has been issuing from the glass tube for a few minutes, place the vessel so that the end of the tube is well under the water, and let it stay there until the thermometer records a

temperature of about 40° C. Then weigh the water again to find the weight of steam condensed. Enter the observations as follows :

Weight of water - - ..... gms.	Temperature of water at beginning - ..... °C.
Weight of water + weight of condensed steam ..... gms.	Temperature of water at end - - - ..... °C.
Weight of condensed steam - - - ..... gms.	Rise of temperature - ..... °C.

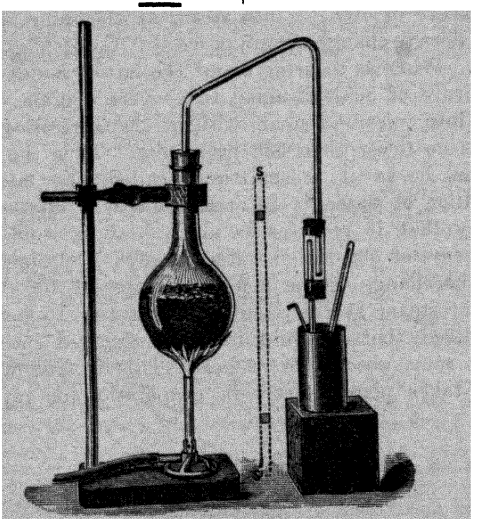


FIG. 116.—Apparatus for determining the heat given up when steam condenses into water.

As before, the changes of temperature can be arranged under two heads :

GAIN.	LOSS.
The temperature of ..... gms. of water was raised through .....° C. The quantity of heat used equals weight of water × rise of temperature. = ... heat units.	..... gms. of steam were condensed to water at 100° C. The temperature of ..... gms. of water at 100° C. fell to .....° C., that is, through .....° C., and the quantity of heat thus given up equals weight of condensed steam (water) × fall of temperature. = ... heat units.

Find the number of units given out by 1 gram of steam at 100° C, in condensing to water at 100° C.

Latent heat of steam.—It is familiar now from many experiments that heat is required to convert water into steam. After what has been learnt about the latent heat of water, you will have no difficulty in understanding the reason for this. All the heat is absorbed, or used up, in bringing about the change from the liquid state to that of vapour. It requires a great many more heat-units to convert one gram of water at a temperature of 100° C. into steam at the same temperature, than it does to change a gram of ice at 0° C. into a gram of water at 0° C. Whereas to bring about the latter change requires an expenditure of 80 heat-units, to convert a gram of water at 100° C. into a gram of steam without changing its temperature requires no fewer than 536 heat-units. Thus the **latent heat of steam**, or as it is sometimes called, the **latent heat of vaporisation of water**, is 536. Expressed in another way, we may say that it requires as much heat as would raise the temperature of 536 lbs. of water through 1° C. simply to bring about the change of one pound of water at 100° C. into one pound of steam at the same temperature. It must also be remembered that a liquid is never changed into a vapour without some absorption of heat. This is true whether the change takes place quietly in evaporation or rapidly as in boiling.

Specific heats.

Brass,	0·0939	Lead,	0·0315
Gas coal,	0·3145	Marble,	0·2158
Copper,	0·0933	Paraffin,	0·622
Glass, crown,	0·161	Steel,	0·118
Gold, flint,	0·117	Sulphur,	0·234
Iron,	0·1124	Zinc,	0·0935

Melting points and latent heats of fusion.

	Melting Point.	Latent Heat.
Ice,	0° C.	79·2
Beeswax,	62	42·3

Boiling points and latent heats of vaporisation.

	Boiling Point.	Latent Heat.
Steam,	100° C.	536
Alcohol,	78	205
Turpentine,	159	74
Sulphuric acid,	338	—
Hydrochloric acid,	110	—
Nitric acid,	86	—
Glycerin,	290	—

CHIEF POINTS OF CHAPTER XI.

Temperature may be defined as a condition, or state, of a body which is changed by the gain or loss of heat. It corresponds to water-level, and may be regarded as heat-level.

The unit quantity of heat is the amount of heat necessary to raise the temperature of one gram of water through one degree Centigrade. This unit is called a *therm*, or a *calorie*.

The number of heat-units taken up or given out by a quantity of water being heated or cooled is equal to the weight of water in grams × number of degrees Centigrade through which its temperature rises or falls.

The capacity for heat of a substance is analogous to the capacity of a vessel for fluids. Water has a greater capacity for heat than any other substance. Its high capacity for heat has a beneficial influence on the climate of islands.

In determining the number of heat-units lost or gained by a substance, its capacity for heat must be taken into account as well as its weight and its temperature. In fact:

Number of heat-units = weight of substance × its rise or fall of temperature × its capacity for heat.

The specific heat of a substance may be defined as the amount of heat required to raise the temperature of one gram of the substance through 1° C. Or, the amount of heat given out by one gram of a substance, the temperature of which falls through 1° C., in comparison with the amount of heat given out by an equal weight of water, the temperature of which falls through 1° C.

Latent heat.—The heat used up in changing a solid into a liquid, or a liquid into a gas, without change of temperature, is known as *latent heat*.

Latent heat of water.—The number of heat-units required to convert one gram of ice at 0° C. into water at the same temperature is known as the *latent heat of water*. Its numerical value is 80.

Latent heat of steam —The number of heat-units required to change one gram of water at 100° C. into steam at the same temperature is known as the *latent heat of steam.* Its numerical value is 536.

EXERCISES ON CHAPTER XI.

1. One hundred grams of boiling water are poured upon one hundred grams of ice. What results may be observed ?

2. Four ounces of hot lead filings and four ounces of water at the same temperature are poured upon separate slabs of ice. Will the lead or the water melt more ice ? Give reasons for your answer.

3. An ounce of water at 0° C. is mixed with ten ounces of water at 70° C. What is the temperature of the mixture ?

An ounce of ice is dissolved in ten ounces of water at 70° C., and the temperature of the mixture is found to be something over 56° C. What can be learnt from this experiment ?

4. Suppose that it requires 80 times as much heat to melt one ton of ice as would be required to warm one ton of water one degree of temperature on the centigrade scale, how much of the ton of ice would be melted by pouring into a cavity in its surface a gallon of boiling water ? (A gallon of water weighs 10 lbs.)

5. How would you propose to prove by experiments that to boil away a gallon of water requires about 5½ times as much heat as is needed to raise it from the freezing to the boiling point ?

6. A silver tea-pot weighs 300 grams. One gram of silver requires as much heat to warm it as would be required by 0·056 gram of water to warm it equally. The tea-pot contains 20 grams of tea-leaves, and each gram of tea-leaves requires as much heat to warm it as would suffice to warm equally 0·5 gram of water. If 600 grams of boiling water be poured into the tea-pot, calculate the highest temperature of the tea, assuming that tea-pot and tea-leaves were originally at a temperature of 15° C.

7. By what experiments would you show that different amounts of heat are absorbed when equal masses of different substances are heated through the same range of temperature ?

8. By what experiments could you show (*a*) that copper conducts heat better than iron ; (*b*) that iron gives out more heat than an equal mass of copper would do in cooling through the same given range of temperature ?

CHAPTER XII.

TRANSFERENCE OF HEAT.

46. TRANSFERENCE OF HEAT. CONDUCTION.

i. Relative conductivities of metals.—Obtain wires or strips of copper, iron, brass, German silver, and of any other metals available. Let the diameters be the same as nearly as possible, and the lengths about 15 or 20 cm. Place the wires upon a clay tile or other suitable support, as shown in Fig. 117. Support the tile in a horizontal position and heat the wires with a flame where they meet. After a few minutes slowly move a safety match along each wire in succession, commencing at the ends away from the flame, and notice the points at which the matches will light. Repeat the experiment several times; then take away the flame and measure the distance of these points from the heated ends. Find the average distance for each wire.

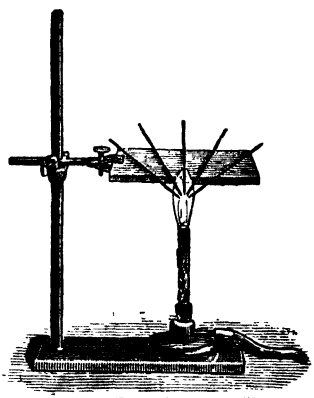


FIG. 117.—Experiment to illustrate the conduction of different metal wires.

Make a list of these distances in order of magnitude, putting beside each measurement the substance to which it refers. Note what this order suggests as to the relative conductivities of the substances.

ii. Lowering of temperature by conduction.—(*a*) Make a short coil of stout copper wire ¼-inch internal diameter. Pass it over the wick of a candle without touching the wick. The candle is extinguished. Is the flame simply muffled out? Convince yourself as to what really happens.

(*b*) Turn on, but do not light, a gas jet. Hold over it a wire gauze, and light the gas above the gauze. Notice that the flame does not strike through (Fig. 118). Why? Vary the experiment

by lowering a piece of cold wire gauze upon an ordinary Bunsen flame. What happens ? (Fig. 119.)

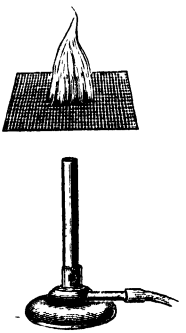


FIG. 118.—To illustrate Expt. 39 ii (*b*).

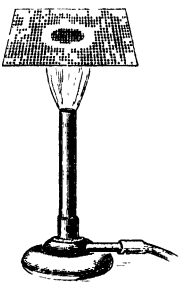


FIG 119.—To illustrate Expt. 39 ii. (*b*).

(*c*) Wrap a piece of paper smoothly round a brass tube and hold in the flame of a gas burner. The paper is not scorched. Wrap the paper around a wooden rod of the same size, and heat

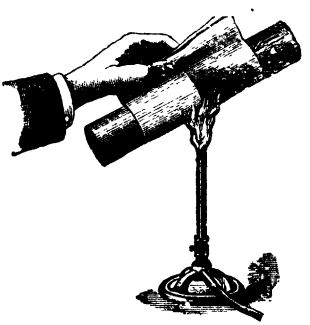


FIG. 120.—Experiment to illustrate the different conductivities of wood and metal for heat.

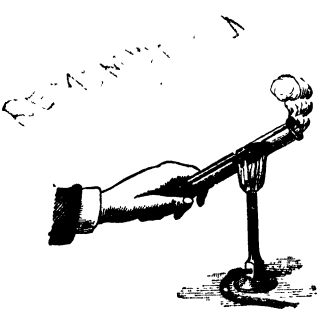


FIG. 121.—Illustration of the fact that water is a bad conductor of heat.

as before : the paper is scorched (Fig. 120). Brass is a good conductor, wood but a poor one. How does this explain what you have noticed ?

iii. Water is a bad conductor of heat.—Fill a test-tube three-quarters full with cold water, and having weighted a small piece of ice by winding wire round it, or in some other way, drop it into the test-tube. Hold the test-tube near the bottom where the piece of ice is, and warm the top of the water in a Bunsen flame, as shown in

Fig. 121. The water at the top can be heated until it boils vigorously and yet the ice is not melted. Why has the ice received insufficient heat to melt it?

iv. Gases are bad conductors of heat.—(*a*) Examine the shadow of a red-hot poker. Notice that the heating of the air as exhibited by its flickering extends but a little way downwards, thus showing that air is a bad conductor of heat.

(*b*) Place a little lime in the palm of the hand and bring the point of a hot poker upon it. The air enclosed in the lime does not conduct the heat of the poker, so the hand is not burnt.

Conduction of heat.—Heat may pass from one particle of a body to the next, travelling from the hotter to the colder parts, and causing no visible motion of the particles of the body. This mode of transference is called **conduction**, and is the process by which solids are heated. By touching a succession of things in a room, say the marble mantel-piece, the fender, the back of a chair, the hearth-rug, we obtain a succession of sensations; the first two we say are cold, the chair-back not quite so cold, while the rug feels quite warm, and yet they are one and all under the same conditions and there is no reason why they should not be at the same temperature. The explanation of these different sensations is really simple. In all those cases where the hand *receives heat* we feel the sensation of warmth, while in those where the hand *gives out* heat we say the thing is cold or cool. Now it can be seen why the fender feels colder than the hearth-rug. The fender takes more heat from the hand than the hearth-rug, and it does so because it is a better **conductor of heat.**

It is worth while to consider this expression a little. Put one end of a short metal rod in a fire and hold the other. Soon the rod begins to feel warm, and as time goes on it gets warmer and warmer, until at last it can be held no longer. Heat has passed from the fire along the rod, or has been **conducted** from the fire by the rod.

The process by which heat passes from one particle of a body to the next is called conduction, and the body along which it passes is known as a conductor.

Good and bad conductors.—Those substances which easily transmit heat in this way are called **good conductors**, while those which offer a considerable amount of resistance to the passage of heat are called **bad conductors.**

Metals are, as a rule, good conductors of heat, but some metals conduct heat better than others.

Most liquids are bad conductors of heat, though quicksilver, being a metal, is an exception. If liquids were heated only by conduction, water would boil throughout just as quickly when the source of heat was placed in contact with the top layer of liquid as it does when the heating takes place from below.

Gases are even worse conductors of heat than liquids. In reckoning the conductivities of solids the proportion of heat carried by the air may be neglected. It is very small.

Everyday uses of bad conductors.—To keep ice in the warm days of summer the custom is to wrap it up in flannel and put it into a refrigerator. The flannel, because of its loose texture, encloses a quantity of air, which, being a bad conductor of heat, prevents the passage of heat from the warm outside air to the cold ice inside. Similarly, ice which has to be conveyed by rail or boat is packed in sawdust.

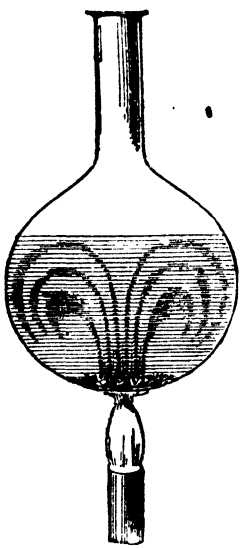


FIG. 122.—Convection currents in water.

The refrigerator itself, too, depends upon much the same facts. The common form consists of a double-walled box with a space between the walls. This is either left "empty," as it is called when it is full of air; or, it is filled with some other bad conductor, such as the mineral substance *asbestos*.

If we wish to lift a hot plate we hold it with a folded cloth which does not readily conduct heat. Cylinders of engines are sometimes encased in a packing of some badly conducting material in order to prevent loss of heat.

47. CONVECTION.

1. Convection in a liquid.—Heat over a small flame a round-bottomed flask full of water, as in Fig. 122. Throw into the water some solid colouring matter, like cochineal, aniline dye, litmus, etc. Notice how the hot, coloured water ascends.

ii. Convection currents in gases.—(*a*) Place a short piece of candle in a saucer, light it, put a lamp glass over it, and pour sufficient water into the saucer to cover the bottom of the lamp glass (Fig. 123). Watch how the light of the candle is affected. Next cut a strip of card less than half the height of the lamp glass, and nearly as wide as the internal diameter of the top. Insert the card into the lamp glass so as to divide the upper part into halves. Now light the candle again, and see whether it will burn with the divided chimney over it. Test the direction of the currents of air at the top of the chimney by holding a smoking taper or match over it.



FIG. 123.—Convection currents in gases.

Process by which liquids are heated.—The process by which water and other liquids are heated may be studied easily by heating water into which some solid colouring matter, like cochineal, aniline dye, litmus, etc., has been thrown, in a round-bottomed flask over a small flame as in Fig. 122. The water nearest the flame gets heated, and consequently expands and gets lighter. It therefore rises, and causes a warm ascending current of coloured water. But something must take the place of this water which rises, and the cold water at the top, being heavier than the warm water, sinks to the bottom and occupies the space of the water which has risen. This water in its turn gets heated and rises, and more cold water from the surface sinks. Upward currents of heated water and downward currents of cool water are thus formed, until by-and-by the whole of the water is heated. These currents are known as **convection currents**, and the process of heating in this manner is called **convection**. Eventually the whole of the water gets so hot that the bubbles of vapour which are formed near the source of heat are not condensed again in their upward passage through the liquid, and coming to the surface they escape as steam.

Gases are similarly heated by the process of convection, which may be thus defined :—**Convection is the process by which fluids**

(liquids and gases) become heated by the actual movement of their particles due to difference of density.

Ventilation.—The ventilation of ordinary dwelling rooms is easily possible because of the way in which gases become heated by convection. The air in a room becomes warmed and is rendered impure at the same time. Consequently there is a tendency for the impure air to rise, and if a suitable place near the ceiling is made for it to get out, as well as a place near the floor for the colder, purer air from outside to enter, a continuous circulation of air is set up which will keep the atmosphere of the room pure and sweet.

To show convection currents in air, place a short piece of candle in a saucer, light it, put a lamp glass over it, and pour sufficient water into the saucer to cover the bottom of the lamp glass (Fig. 123). In this case the light of the candle is affected and eventually goes out. But if a strip of card is cut less than half the height of the lamp glass, and nearly as wide as the internal diameter of the top, and is inserted into the lamp glass, so as to divide the upper part into halves, and the candle is again lit, it will be seen to continue to burn with the divided chimney over it. The simple change has secured a well directed current of air which feeds the flame. The direction of the current through the top of the chimney can be shown by holding a smoking taper or smouldering brown paper over it.

48. RADIATION.

1. Heat transmitted by radiation.—(*a*) Place the differential thermometer (Fig. 90) about a foot away from the flame of a laboratory burner so that both its arms and the flame are in one straight line. Notice that the bulb of the thermometer nearer to the flame is hotter than the one more remote. How does the heat of the flame travel to the thermometer?

(*b*) Arrange the bulbs of the thermometer a foot above the flame and observe that the nearer bulb is warmed much more than before. The bulb in this case is heated by convection as well as radiation.

(*c*) When you have an opportunity, focus the rays of the sun upon the back of your hand by means of a reading glass. This can be done by placing the reading glass between the sun and the hand and moving the glass until the brightest image of the sun is obtained Notice that the heat is very intense and burns you. Notice that the glass itself is not heated to the same extent.

ii. Effect of surface upon radiation and absorption.—(*a*) Obtain two small bright tin cans or canisters, and fit into each a cork having a hole through which a thermometer will pass. Cover the outside of one of the vessels with lamp-black by holding it over a candle or luminous gas flame, or over burning camphor. Put the same quantity of hot water at the same temperature in each, and then cork up the vessels, each cork having a thermometer through it so that the bulb is well immersed in the water. Observe the temperature of each vessel of water, and if the temperature of one is higher than that of the other, cool the vessel until the temperatures are equal. Then put the vessels in a cool place where there are no draughts, and after 20-30 minutes again read the temperatures.

The blackened vessel will be found to have lost or radiated more heat than the bright one.

(*b*) Similarly pour equal amounts of cold water of the same temperature into a blackened and a bright vessel, and hang them for 20-30 minutes before an even fire or closed stove, or at the same distance above an iron plate, supported on a tripod stand and heated by a laboratory burner so that they may be in a position to receive heat equally. At the end of this time observe their temperatures. The blackened vessel will be found at a higher temperature than the bright one. Which vessel *absorbed* more heat? Compare with the vessel which *radiated* more in the last experiment.

Radiation of heat.—The fact that you feel warm in a summer sun, or that bread can be "toasted" by holding it near the fire, is sufficient to convince you that heat can travel from one place to another in a third way which is neither conduction nor convection. The respects in which **radiation** differs from the other ways in which heat moves from one place to another are:

(1) it travels in straight lines, and

(2) it does not warm the medium through which it travels.

Although you may not have thought of the fact that radiation travels in straight lines, you have made use of it when you have screened your face from the heat of the sun or of a fire. When you wish to protect yourself from the glare of the summer sun you seek a shady space because then some object, it may be a tree or a house, is in the straight line between you and the sun.

Curtains have sometimes been burnt by the sun's rays being concentrated upon them by a bottle of water, though the water is not warmed much by the passage through it of the radiations from the sun. Evidently, then, the water in such a case does not pass on its heat after first becoming warm itself. It

does not act as a conductor. Yet something must pass through it which can make bodies hot. This something is called **radiation.** Its nature is simply a wave motion in the medium through which the rays pass. This medium is known as "ether," but it is in no way connected with the liquid ether used for scientific purposes. The ether is little more than a name, for, though something must exist to transmit waves of light and heat, nothing is certain as to its constitution.

49. DEW.

i. Condensation of moisture.—(*a*) Breathe on various cold objects such as a mirror or polished metal. Note what happens.

(*b*) Bring a glass of cold water into a warm class-room. Observe the dull film of moisture deposited on the glass.

(*c*) Does dew form on one species of plant more than another, or on any one part of a leaf in particular?

(*d*) Choose a clear, still evening, and arrange stones, pieces of slate, and sheets of paper on grass; examine them the next morning as soon after sunrise as possible. Which surface, the under or the upper, is more bedewed?

(*e*) Invert a few tumblers, earthenware jars, etc., some on grass and some on soil, both on clear nights and on cloudy nights. As in the last exercise, examine them as soon after sunrise as possible. Is there a deposit of dew inside the jars, etc.?

(*f*) Repeat the last exercise, but place the jars, etc., on metal plates, slates, or tiles. Examine the inside and outside of the jars, etc., as before, and record the results.

Dew.—Dew differs from the forms of condensed moisture seen in mists, clouds, rain, and snow, in being formed *upon* the surface of the earth. After sunset, the surface of the earth, which has been receiving heat throughout the day, begins to radiate this heat in the manner already described. Different parts of the earth possess differing powers of radiation. Those which during the day absorb heat to the greatest extent radiate it most abundantly after the sun has disappeared, and consequently become cooled before those the radiating power of which is small. Similarly, the air in contact with these bodies also becomes cooled and is then unable to hold as much water vapour as before and the surplus is deposited in the form of **dew.**

For an abundant formation of dew several conditions are necessary. First, radiation must go on freely, and this happens

on *bright clear evenings* when there are no clouds to obstruct the radiation. The air which is being cooled by contact with the body from which free radiation is taking place must not be disturbed before the dew-point is reached, or no dew will be thrown down, that is, the *evening must be still.* A breeze will constantly renew the air above the body which is being cooled by radiation and will prevent the dew-point being reached. Good radiating surfaces are those of leaves—whether of grass or other plants—also stones.

Side by side with this simple condition of things another process is going on which considerably augments the amount of dew formed. Throughout their life, plants continually give off water in the form of vapour, which is exhaled through the numerous apertures spread over their leaves, especially the under surfaces. This process, which is known as *transpiration*, supplies a very large amount of water vapour to the air. When the cooling referred to above has gone on for some time, and the dew-point has been reached, the transpired moisture, instead of diffusing into the atmosphere in the state of vapour, is condensed at the *stomata*, as the little apertures are called, as soon as it comes into contact with the cooled air. Thus all the dew is not obtained directly from the atmosphere.

Hoar-frost—or as it is sometimes called *white-rime* or simply *rime*—is deposited instead of dew on those evenings when the radiation cools the overlying air to the temperature of freezing water before any deposition of moisture takes place. Hoar-frost is not frozen dew. It does not first assume the liquid condition, but is precipitated at once in the solid form. In these circumstances the **dew-point** is at or below the freezing-point.

50. DETERMINATION OF THE DEW-POINT. HYGROMETERS.

i. Mason's hygrometer.—Take two precisely similar unmounted thermometers. Simply hang one from a suitable support, such as the ring of a retort stand. Cover the bulb of the other with a square of muslin tied up round it to form a bag. The muslin is best tied just above the bulb by a piece of thread. To this piece of thread attach several other long pieces of thread and let them dip into a glass of water. When the muslin has become thoroughly damp,

compare the readings of the two thermometers (Fig. 124). Notice that the temperature recorded by the one with the wet muslin round it is lower than that shown by the other.

The two thermometers used in this way form what is known as a *hygrometer*, or a *wet- and dry-bulb thermometer*.

ii. **Regnault's hygrometer.**—Fit up a large test-tube in the manner shown in Fig. 125, where *a* is a right-angled glass tube which dips into some ether in the test-tube; *b* is a second glass tube bent at right angles, which just passes through the india-rubber stopper; *e* is a delicate thermometer dipping into the ether; *d* is a piece of india-

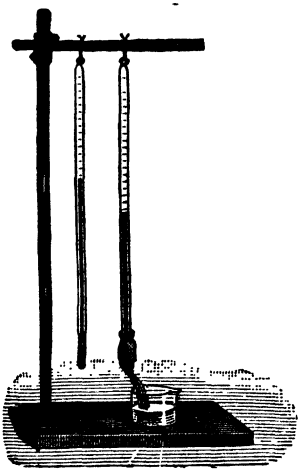


FIG. 124.—The bulb of one thermometer is kept moist, and the evaporation of this moisture causes this thermometer to show a lower temperature than the other.

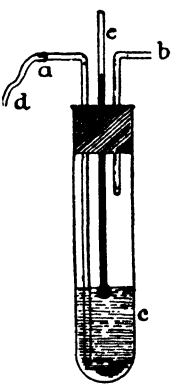


FIG. 125.—Experiment to illustrate the action of Regnault's hygrometer.

rubber tubing attached to the tube *a*. A second thermometer is supported in the neighbourhood of the apparatus for recording the temperature of the air. Blow through *d*. This causes the ether to become vaporised, the vapour escaping through *b*. This vaporisation is effected at the expense of the heat in the test-tube, which consequently becomes cooled, and after a time moisture is found to be deposited on the outside of the test-tube. At the instant such deposition occurs read the thermometer *e*. Stop blowing and read the temperature again at the instant the moisture outside the test-tube disappears. The mean of these two temperatures is the *dew-point*.

Mason's hygrometer.—Mason's instrument consists of two precisely similar thermometers, suitably attached to a frame as

in Fig. 124. Round the bulb of one of the thermometers is tied a piece of muslin, to which cotton threads are attached, these hang down into water kept in a glass, supported as shown in the figure. The instrument depends for its use upon two facts which have been brought before the student's attention already. The first is that water is vaporised only at the expense of a certain amount of heat; and, secondly, the quantity of water vapour which air can take up at any temperature depends upon the amount already contained by it. Water rises up the cotton threads by the force known as capillary attraction, and consequently keeps the muslin moist. The water on the muslin evaporates, getting the heat necessary for evaporation from the bulb of the thermometer which it surrounds. The thermometer is thereby cooled, and the column of mercury sinks. This process continues until the air round the bulb is saturated and evaporation ceases. Thus the wet-bulb thermometer records a lower temperature than the one with a dry bulb. The difference between the readings is greater the drier the air at the commencement of the observation, and it provides a means of estimating the amount of water-vapour present by seeing how much more must be added to saturate the air.

The **wet- and dry-bulb thermometer**, as Mason's hygrometer is called, is usually employed to indicate the relative amount of moisture in the air, but the readings may also be used to determine the **dew-point** (that is, the temperature at which dew would be deposited from the air at the time of observation) by a simple calculation in connection with a set of hygrometrical tables prepared for the use of practical meteorologists. When the dew-point has been found, the relative humidity of the air or the percentage of saturation can be determined.

Regnault's hygrometer.—Regnault's hygrometer depends upon the same principle as that exemplified by Expt. 50 ii. The construction of this instrument is shown in Fig. 126. *D*, *D* are two polished silver thimbles, in which are arranged two test-tubes. The one on the right is half full of ether, and passing down into this ether is a right-angled tube *t* and a delicate thermometer *T*. There is a side tube in connection with that in the right-hand thimble which puts this test-tube in connection with a hollow tube, *U V*, which by means of a piece of india-rubber

tubing can be placed in connection with the aspirator *A*. The tube in the left-hand thimble is *not* in connection with *UV*, the thermometer in it being used only to read the atmospheric temperature. The stopcock shown is turned on, and air is drawn

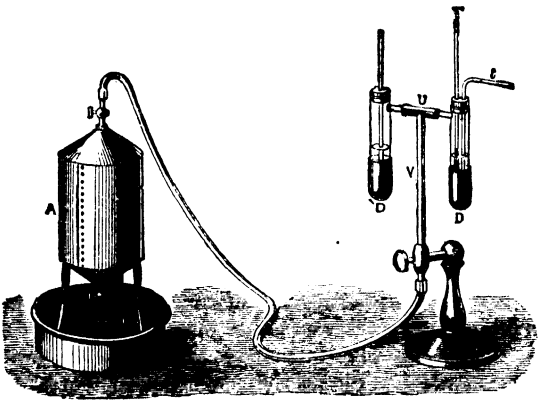


FIG. 126.—Regnault's hygrometer.

through the ether by means of the aspirator, resulting in a deposition of moisture on the outside of the right-hand thimble *D*. The moment at which this deposition occurs the thermometer *T* is read. The blowing is stopped and the temperature is read again at the instant dulness disappears. The mean of these two temperatures gives the **dew-point**.

CHIEF POINTS OF CHAPTER XII.

Heat is transmitted in three ways (1) by conduction, (2) by convection, (3) by radiation.

Conduction is the process by which heat passes from one particle of a body to the next.

Gases are worse conductors than liquids, and liquids are usually worse conductors than solids.

Convection is the process by which fluids (liquids and gases) become heated by the actual movement of the particles of the fluid due to difference of density.

Heating buildings by hot water and methods of ventilation are applications of heating by convection.

Radiation differs from conduction and convection in two ways (1) it travels in straight lines, (2) it does not warm the medium through which it travels.

If air containing water vapour is cooled sufficiently it will yield up its excess of moisture. The temperature at which this deposition takes place is called the **dew-point.**

When air has as much water vapour as it can hold it is said to be **saturated** with it.

The amount of water vapour actually present in air, compared with the maximum quantity it could take up, gives a measure of the **Hygrometric state** of the air.

EXERCISES ON CHAPTER XII.

1. What is meant by convection?

Illustrate your answer by sketches, taking the case of a vessel filled with water and heated from below and explain why it is that convection is set up.

2. Why is a vessel of water heated more quickly if heat is applied at the bottom than if it is heated at the top?

Draw a diagram to illustrate the movements of a liquid heated from below.

3. Point out the difference between the conduction and convection of heat. Describe an experiment showing that water is a bad conductor of heat.

4. Water sometimes spurts from the spout of a kettle standing upon a fire. How do you account for this, and how would you prevent it without taking the kettle off the fire.

5. On a cold morning a gardener grasps the iron part of his spade with one hand and the wooden part with the other. Explain why one hand feels colder than the other.

6. If a spoon made of solid silver and one made of brass and only silver plated are placed in bowls in some boiling water, the handle of the silver spoon becomes much hotter than that of the plated one. Why is this?

Describe an experiment by which you would show that your explanation is correct.

7. How is the reading of a thermometer altered by wrapping a wet rag round the bulb? What will happen if the rag is wetted with (1) ether, (2) oil instead of water? How do you explain the various results?

8. Two test-tubes *A* and *B* are filled with water. A small piece of ice is allowed to swim in *A*, and a similar piece of ice is sunk by a weight to the bottom of *B*. Heat is applied to the closed end of *A* and to the open end of *B*. In which test-tube may we expect the ice first to melt? and in which may we expect the water first to boil? Give reasons for your answer.

9. On some days a steam-engine leaves a long white cloud behind it, and on some other days only a short one. Explain this by describing in each case why the cloud forms, and why it disappears.

10. A saucer containing water is left to evaporate on a window sill. Explain the atmospheric conditions which will favour or retard the disappearance of the water.

11. Explain what happens to the steam issuing from the funnel of a steam-engine: (*a*) on a fine warm day; (*b*) on a damp day; (*c*) in a tunnel.

CHAPTER XIII.

ATMOSPHERIC PHENOMENA. OCEAN CURRENTS.

51. MIST, FOG, CLOUD, RAIN, SNOW, AND HAIL.

Mist and fogs.—The general features of the forms of condensed moisture referred to under these names are familiar to every one. We naturally associate mists with rivers or other water surfaces; most often they make their appearance after the sun has set. They seem to be caused as follows: the air over the land by the side of the river gets cold more quickly than that over the river itself, because land radiates heat very much better than water does. The air over the water will show a tendency to rise, and the cold air will move towards the water to take its place. But the air over the water will thus be cooled and will be unable to hold as much water-vapour as before, and the excess of moisture assumes the nature of a mist. It is impossible to say when a mist becomes a fog. A fog is generally regarded as something denser and blacker than a mist. The condensation which forms a fog, at all events, and probably a mist also, takes place round the small particles of dust in the air, which act as nuclei for the minute water drops which make up the fog. In London or the large manufacturing towns, where there is so much carbon in the air, the fogs become very bad indeed—or, as is commonly said in the former place, they are thick enough "to cut."

Clouds.—Clouds are sometimes formed in a similar manner to mists and fogs. They differ chiefly in their position, and can be very correctly regarded as fogs high up in the air. When from any cause an upward current of warm air laden with invisible moisture is cooled in these higher atmospheric regions,

a cloud will be produced. This can be brought about by a warm, moist stratum coming into contact with a cold current of air, whereby it becomes cooled and some of its moisture is condensed into minute particles, which Tyndall called "water-dust." Or, in its upward passage, it naturally comes to a colder zone, and as naturally some of its moisture takes the visible form of a cloud. Moreover, in ascending to levels where the pressure is less, air expands, and is thereby cooled; in fact, this cause of cooling is far more effective than the former.

Rain.—The particles constituting a cloud, that is, the water-dust spoken of above, continually tend to coalesce or unite together to form larger drops. When these reach a certain size the air can no longer support them and in consequence of their increased mass they fall, being attracted more strongly to the earth. They do not always reach the surface of the earth, however, for in their downward course it may happen that they have to traverse a layer of dry, unsaturated air, when the drops will become wholly evaporated again. In their passage through very moist air, then, rain-drops continually get larger, whereas in falling through dry air they become smaller, until they may even eventually disappear.

Snow.—Sometimes the temperature of a cloud is below the freezing-point of water, when it is manifestly impossible for the moisture to assume the liquid state, and it becomes condensed in a solid form. If, in addition to this, the temperature of the air through which the descending solid particles pass is below the freezing-point, we shall have a fall of solid particles in the form of **snow**. The falling particles unite continually to build up larger masses which we know as **snow-flakes**, which under favourable conditions, assume the most beautiful forms. Ice crystallises in the system called the *hexagonal*, and it is found on examining snow-flakes that they are combinations of minute crystals belonging to this system. Their shapes are seen to perfection in the Arctic regions, and observers in these high latitudes have recognised over a thousand different forms. In this country they will be seen best when a fall of snow takes place in a still, quiet atmosphere, with the temperature at zero or lower to prevent a partial thaw ruining their exquisite beauty. If the snow in its descent upon the earth becomes

partly melted, and perhaps later on partly frozen again, instead of snow-flakes reaching the surface we shall have a mixture of half-melted snow and rain, which is called **sleet.**

Hail.—The mode of formation of hail-stones has never been satisfactorily explained. In our country it falls more commonly in summer and spring than during the winter months. Its existence seems in some way to be connected with the electrical state of the atmosphere. It has been suggested that it is quite sufficient to imagine a very cold current of air impinging against

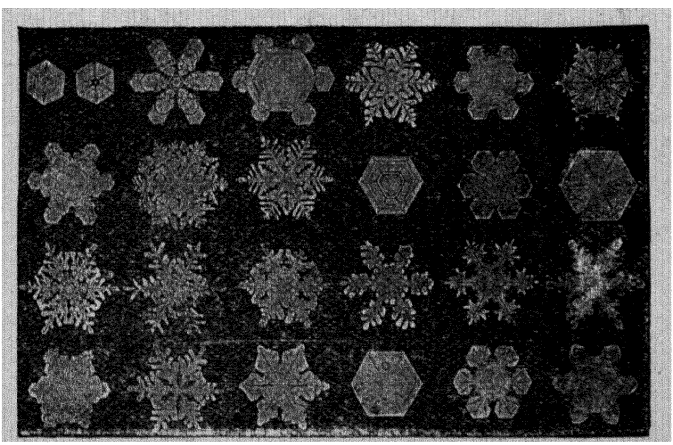


FIG. 127.—Snow Crystals. (From photographs by Mr. W. A. Bentley.)

a moist cloud, and suddenly lowering its temperature as a whole to below the freezing-point, to account for its formation. However it may be formed, it is another instance of the condensation of atmospheric vapour in the solid state. Hail-stones take the form either of hard or soft pellets which vary in size from a small pea to a small hen's egg. Just as rain-drops and snow-flakes commonly increase in size as they approach the earth's surface, so hail-stones aggregate together in their downward flight, becoming larger and larger as they reach their destination. An examination of hail-stones at different times and in various places has shown that they vary in their nature. Sometimes, on cutting through a hail-stone, it is seen to be built up round a speck of dust as a nucleus and to take the appearance of having

been built up gradually, and not formed *en masse*, exhibiting, as it does, a more or less stratified structure. Soft hail-stones seem to be small lumps of snow without any central dust particle.

52. CIRCULATION OF THE ATMOSPHERE.

General remarks.—That the air is in movement is a fact of common knowledge. We see the results of its motion as we watch the branches of trees swaying to and fro. We feel the impact of the air particles upon our face when we turn towards a strong breeze. We have now to consider several questions having reference to this circulation of the atmosphere. Is there any regularity or order in the way in which winds blow ? What causes them ? and a host of other questions which will present themselves to the student. Before attempting an answer to such queries as these, it will be convenient to call attention here to the way in which winds are named. In describing the direction of ocean currents (p. 200), a northerly current is one which flows toward the north, or similarly with any other. But the contrary is true of winds. **The direction of an ocean current is always given as that point of the compass to which it flows, whereas that of a wind is always spoken of as that from which it blows.**

The cause of winds.—In discussing the movements of liquids we found that water always flows from a place of high pressure to one where the pressure is lower. We said it sought its own level. Similar movements take place in all fluids ; there is in every case a movement from a point of higher to one of lower pressure until the pressure is equalised. But as has been learnt, the pressure of the atmosphere varies considerably from place to place, and being a fluid there will naturally be a disturbance of the whole, resulting from the endeavour to bring about an equilibrium of pressure. **The air will move from the places where the pressure is high towards those spots where the pressure is low.** These movements constitute **winds.** The winds are **permanent** if the difference of pressures causing them are constant throughout the year. They are **periodic** if the pressure differences only arise at stated intervals. **Variable** winds result from any pressure disturbance which may ensue from local peculiarities of situation or from any other cause. It has been seen that variations in pressure are the result of alterations in

temperature and of the increase or decrease of the amount of water vapour held by the air. These causes are, in consequence, to be regarded as the primary factors in bringing about winds.

It has been found that the pressure of the atmosphere is least in polar regions and along the equator, while the districts of greatest pressure occur approximately along the Tropic of Cancer and also along the Tropic of Capricorn, in the northern and southern hemispheres respectively. From what has been said, it is manifest that there will be a movement of air from the Tropics of Cancer and Capricorn towards the poles on one hand and towards the equator on the other. If the earth were at rest, that is to say, we should have in the northern hemisphere a north wind blowing between the Tropic of Cancer and the equator and a south wind from near the same parallel of latitude to the north pole; while in the southern hemisphere there would be a south wind between the equator and the corresponding parallel of latitude and a north wind extending from the same circle to the south pole.

The trade winds.—But the earth is not at rest. It is spinning round like a top. The poles are at rest while places on the equator are performing a journey of 25,000 miles in 24 hours, that is, are moving with a velocity of over a thousand miles an hour. Other places on the surface have velocities intermediate between these two extremes and dependent upon their latitude. Bearing this in mind, we will consider again the wind in the northern hemisphere, which would be a north wind were the earth at rest, and would blow between the Tropic of Cancer and the equator. The air moving towards the equator is subjected to two velocities—(1) that which it has in a southerly direction, depending upon the actual pressure difference between the place from which it starts and that towards which it moves; (2) that which it has in a direction from east to west, due to the earth's rotation from west to east. The resultant is obtained in the manner described on p. 41. By applying this method the student will understand that the direction which the wind will appear to have will be from the **north-east**, and this movement will give rise to the north-east winds which are more or less permanent in the low latitudes we have specified. They are known as the **Trade Winds**, and were so called because of the assistance they rendered to the navigation of trading vessels

before the introduction of steamers. They blow with great constancy across the oceans, but are more or less interfered with over the continents, since the local conditions on the land vary from place to place.

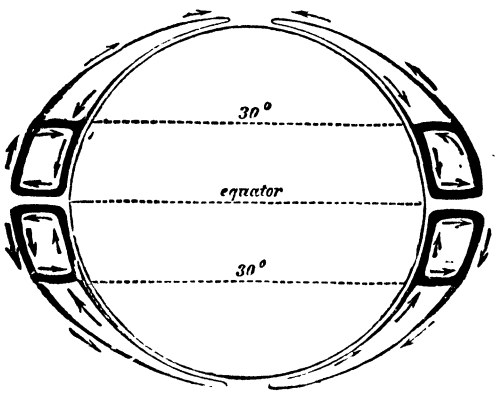


FIG. 128.—Explanation of the trade winds and atmospheric circulation.

By applying the same reasoning to the southern hemisphere it will be clear that the direction of the trade winds south of the equator will be south-east.

Land and sea breezes.—Near the sea, especially in the tropics, there are well-marked breezes, which result from the

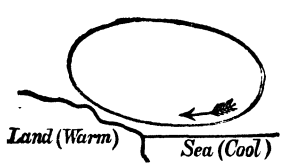


FIG. 129.—Sea breeze.

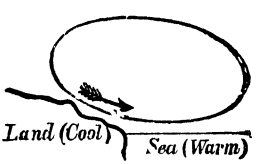


FIG. 130.—Land breeze.

different thermal properties of land and water. Water has a higher specific heat, and is also a poorer absorber of heat than land, and consequently during the day the air above the land will become warmer than that above the water, and an upward current of air will be set up over the land. The cooler air from over the sea will flow in to take the place of the air which rises

and will constitute a **sea-breeze.** After sunset both the sea and land begin to radiate heat; the land being a better radiator becomes cool quickly, but the sea remains warm. The air over the water consequently becomes warmer than that over the land, and the pressure above the sea will be lower than that over the land, causing a current of air from the land out to sea, which is known as a **land-breeze.**

Monsoons.—In describing the trade winds no reference was made to the Indian Ocean, and this was because the conditions here periodically change as a result of the proximity to the great land-masses of Asia as well as of the apparent annual motion of the sun. The name *monsoon* is itself derived from a Malay word meaning a *season*, to mark the fact of the periodical change in the direction of the winds, that is, the seasonal variation which they undergo. In summer, the Asiatic continent is shone upon nearly vertically by the sun, and therefore it becomes intensely hot compared with the waters of the Indian Ocean to the south. As a natural result the air over the land rises because of the decrease in its density. The air from the relatively cold ocean flows in to take the place of the upward current, and in consequence of the earth's rotation flows from the *south-west* across India, but in directions in other parts depending upon the nearest place of high pressure. This south-west monsoon blows from April to October. In the southern hemisphere, in a manner somewhat similar to that causing an alteration in the direction of the trade winds, the direction of the monsoon blowing in these months is south-east. In winter the conditions have become changed. The sun is now more or less vertical over the latitudes in the vicinity of the Tropic of Capricorn. The air over the Asiatic continent becomes cooled and denser, while that over parts of Africa has become warmed and rarer. There is a wind caused which blows from Asia to Africa in a north-easterly direction, and is felt from October to April.

It must not be supposed that monsoons blow in no places besides the Indian Ocean, for wherever the local conditions interfere with the regular trade winds, these winds will partake of a seasonal character. Monsoons occur for these causes in Madagascar, Guinea, Australia, Brazil.

53. OCEAN CURRENTS.

i. Circulation of water.—Place a piece of ice in a trough of water, *A*, *B*, *C*, *D*, and at the other end of the trough arrange a metal rod *E*, kept hot by a flame *F*. Pour a little coloured liquid into the trough, and notice the general movement of the water as shown in Fig. 131.

Ocean currents. Causes.—There are several causes at work tending to produce movements in the waters of the oceans. They have already been referred to incidentally, and can be satisfactorily enumerated at once.

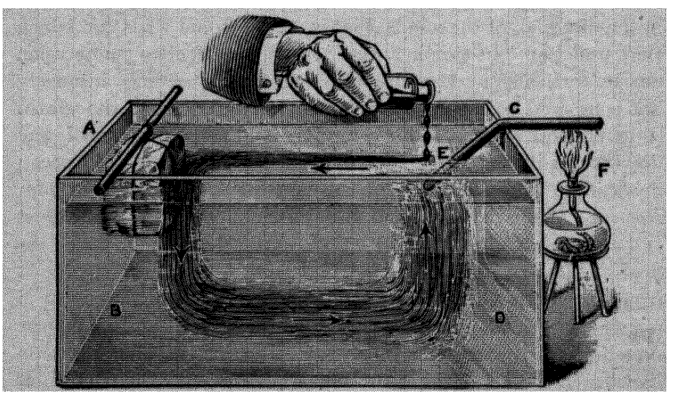


FIG. 131.—Circulation of water.

(1) **Action of the prevailing winds.**—In this connection the comparative shallowness of the ocean must be insisted upon, for it is only by bearing this in mind that any conception of the power of the wind can be attained.

(2) **The heating effect of the sun in tropical regions.**—The effect of heat upon liquids is to make them expand, causing a given mass to occupy a larger volume, and so become lighter bulk for bulk. The result is a rising of the lighter waters and a sinking of the heavier colder waters to take their place, precisely as has been described under convection currents (p. 182).

(3) **The increase in saltness and consequently of density of the water as a result of evaporation.** The waters of the ocean contain solid substances in solution. When heated they give off pure water in the form of vapour, the saltness consequently increases and with this the waters become heavier bulk for bulk thereby disturbing the equilibrium.

It is evident that the second and third of these causes produce contrary effects and tend to neutralise one another. It is most probable that the winds constitute the motive force which results in the production of the great regular movements of the water which are referred to under this heading. The winds are set up by the sun and the amount of evaporation depends upon temperature, so that really the great cause of oceanic circulation is the energy of the sun.

The general result of the difference of temperature in equatorial and polar regions is a tendency for warm surface currents to flow into higher latitudes, while a cold under current creeps along the ocean bottom from the poles towards the equator. This action has been illustrated in Experiment 53 i.

CHIEF POINTS OF CHAPTER XIII.

Fog consists of minute droplets of water, each having a small particle of dust or other matter as a nucleus, and formed near the earth's surface.

Mist is similar to fog, but not so dense, and the particles of water in it are believed to be larger than in fog.

Clouds mostly consist of innumerable particles of water suspended high up in the air; they are elevated mists or fogs. Some clouds, however, are composed of ice particles.

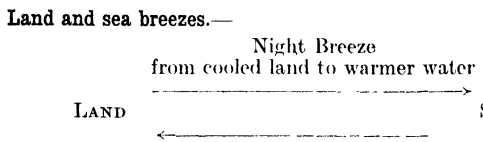
Rain consists of drops of water, formed by the union of the minute water particles of clouds, which fall towards the earth in consequence of their increased mass.

Snow is the solid form in which the moisture of clouds is precipitated when the temperature is below the freezing-point of water. Snow-flakes always have a regular crystalline form.

Hail consists of pellets of ice or snow, frequently made up of concentric layers formed round a nucleus. This stratified structure proves its formation in stages and not all at once.

Wind is air moving from places where the pressure is high towards those spots where the pressure is lower. The differences of pressure are produced by variations in the temperature and moisture of the air.

Land and sea breezes.—



Monsoons are periodical winds strongly felt in and about the Indian Ocean and China Sea.

Monsoons. { Northern hemisphere { North-east, October to April / South-west, April to October. Southern hemisphere { South-east, April to October. / North-west, October to April

Ocean Currents.—The great ocean currents are mainly due to the action of the winds. Other minor causes are the heating effect of the tropical sun, and the increase of saltness and consequently of density of salt water by evaporation.

EXERCISES ON CHAPTER XIII.

1. Distinguish between fog and mist. Why is the former more common in towns ?

2. Describe carefully how clouds are formed. Explain how rain, snow, or hail, falls according to the condition of the atmosphere.

3. What is the cause of wind ? Account for the directions of the Trade Winds in the two hemispheres.

4. Explain what is meant by (*a*) Land and Sea Breezes, (*b*) Monsoons.

5. Describe an experiment to illustrate the circulation of water in consequence of a difference of temperature.

6. Give a short account of the great ocean currents.

CHAPTER XIV.

PROPAGATION AND REFLECTION OF LIGHT.

Light is a form of radiation.—In considering, in a part of the previous chapter, the ways in which heat can be transferred from one place to another, it has been seen that the heat of the sun reaches the earth by radiation. These solar radiations comprise what is collectively called sunlight; they are conveyed in the form of waves through the medium ether, which is believed to pervade all space, and may be conveniently referred to as ether-waves. These ether-waves are of various lengths and can produce different effects. If they fall upon our bodies the longer waves may be absorbed, and the energy of the wave-motion become converted into **heat**; if they fall upon the retina of an eye, the shorter waves may produce a sensation of light and the waves are then spoken of as **light**; falling upon a photographic plate or upon a green leaf, the shortest ether-waves may produce chemical effects, and are then referred to as **actinic**. But, in their passage through the ether, these ether-waves do not give rise to any of these effects; they are simply waves transferring energy by wave-motion.

54. RECTILINEAR PROPAGATION OF LIGHT.

1. **Light travels in straight lines.**—Take three cards and make a small hole in each with a fine needle. Fix the cards upon wooden blocks, so that all the holes are at the same height and in a straight line. Place a lighted candle or a lamp in front of the first card, and look through the third (Fig. 132). So long as the holes are in a straight line you can see the light from the candle shining through. Move one of the cards aside, and notice that you can no longer see

the light. What is true of light applies equally to all other kinds of radiation.

ii. **Pin-hole camera.**—Construct a pin-hole camera as follows: Make two pasteboard tubes by rolling pasted paper on a wooden cylinder, so that one fits inside the other. For the wider tube previously cover the cylinder with dry paper. Cover one end of the narrower tube with tissue paper, and thrust this end into the wider tube. Line with black paper. Place the tube with the pin-hole

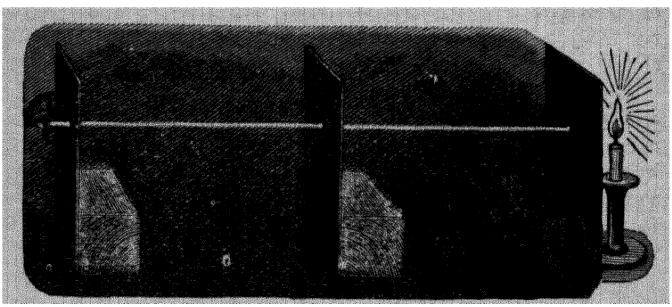


FIG. 132.—The light of the candle can only be seen when the three holes in the screens are in a straight line.

facing a luminous object, *e.g.* a candle. Notice that the image of the candle seen upon the tissue paper is upside down. Reason out how this image is formed. Many toy-shops sell a pin-hole camera at a penny, and one of these may be used instead of constructing one from tubes.

iii. **Overlapping of images**—Make several pin-holes near the first one in the pin-hole camera. For every pin-hole there is an image formed on the screen. Make the pin-holes more and more numerous, and nearer together, till the images overlap and become confused. At last diffused light is produced, which is an overlapping of images.

This experiment explains why the image becomes blurred, and eventually disappears if the size of the single pin-hole is increased.

Light travels in straight lines.—That light rays travel in straight lines can be shown at once by examining the paths of the rays as they pass through a hole in the shutter of a darkened room. Though the light-waves are not themselves visible, yet the path of the light becomes apparent, because the minute particles of dust in the air are rendered luminous by the vibrations of the ether being reflected by them. If there were no dust particles in the room the beam of light would be invisible. When the path of a beam is made visible by smoke or dust it is

seen to be a straight line. That light travels in straight lines may, indeed, be inferred from several everyday experiences. We cannot see round a corner; if light travelled through a uniform medium in lines that were sometimes bent, there is no reason why we should not. Or, again, everyone knows that it is only necessary to put a small obstacle in the path of the light from a luminous body to shut out completely our view of it. The light from the setting sun, when the sky is cloudy, is often seen to travel in straight lines.

The images produced by a pin-hole are inverted.—When an object is viewed through a pin-hole camera, it is seen to be upside down upon the screen. Similarly, all images produced by a small aperture are inverted. This inversion is a direct consequence of the fact that light travels in straight lines. That this is really the case can be understood fully by the following simple considerations. Let *H*, in Fig. 133, be the pin-hole, and *AB* the candle. Rays are sent out in all directions by every point of the candle, but of all the rays from a point, such as *A*, only that in the direction *AH* can pass through the hole and form an image *A'*. Similarly, the only ray from *B* which can get through the hole is *BH*, so an image of *B* is formed at *B'*. The same reasoning applies to any part of the candle, hence a complete inverted image is produced.

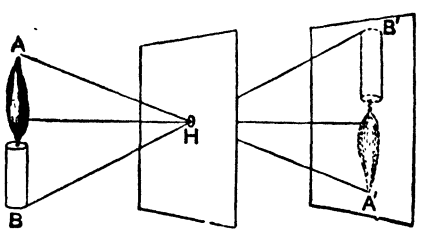


FIG. 133.—Explanation of the inversion of images seen through a pin-hole.

If the light which passes through a small hole in the shutter of an otherwise dark room be caught by a screen of cardboard, a coloured, inverted image of the sky and landscape will be seen. By using a pin-hole camera, a photograph of the view can be obtained (Fig. 134). The bright circles of light seen under trees in summer are really images of the sun formed by the small spaces between the leaves.

Size of image produced by a pin-hole.—That the size of the image depends upon the distance of the screen from the pin-hole

is proved practically by varying the distance of the screen from the pin-hole and measuring the length of the image. The

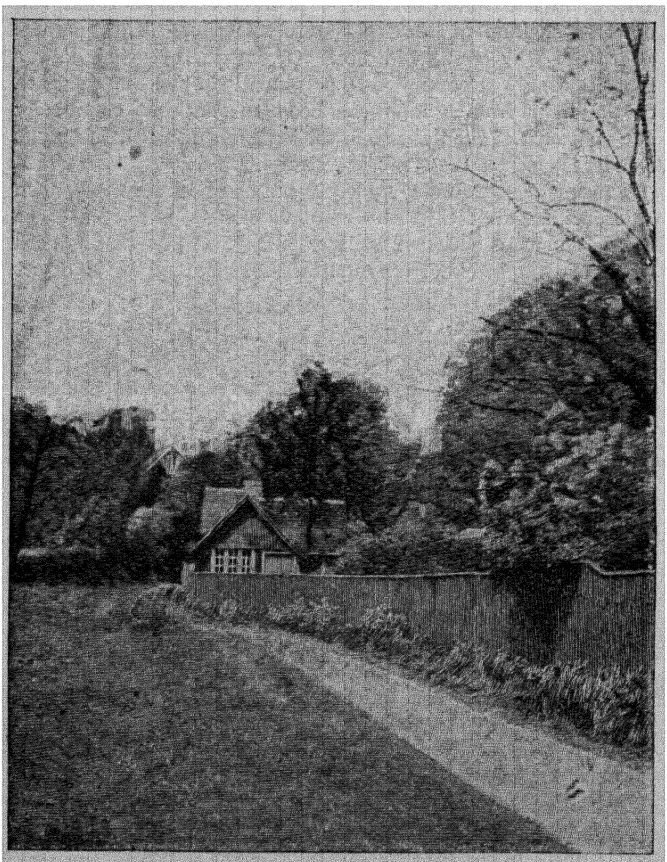


FIG. 134.—Picture taken by a pin-hole camera.
From a photograph by Mr. F. Butterworth.

greater the distance of the screen the longer the image. The reason for this alteration in the size of the image is a simple one. The rays of light from the top and bottom of the

object travel through the pin-hole, and since one is travelling upwards and the other downwards, they will be farther apart the greater the distance they travel. Consequently, the image is longer the more the screen is moved from the pin-hole.

The relation between the sizes of the object and image according to their distances from the aperture is

$$\frac{\text{length of object}}{\text{length of image}} = \frac{\text{distance of object from aperture}}{\text{distance of image from aperture}}.$$

The larger the image the less bright it is, because the small amount of light is spread over a greater area in the case of the enlarged image.

Illumination due to overlapping of images.—When a pin-hole is made in the front face of a pin-hole camera, an image of

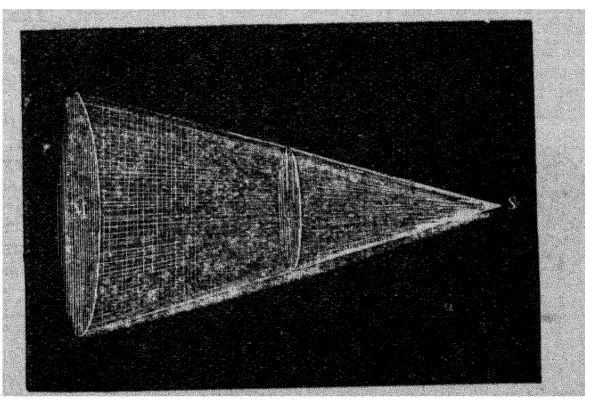


FIG. 135.—To illustrate why the intensity of light diminishes as the distance from it increases.

the bright object looked at is formed on the screen in the manner described in the preceding paragraphs. If a second hole be pierced, a second image is obtained. When the number of holes is increased steadily, one at a time, the images, it is observed, start overlapping, at the same time becoming blurred. When the number of images has become considerable, no separate image can be distinguished; **diffused light**, as it is called, is produced, and the screen is illuminated in the ordinary way.

Intensity of light.—In proceeding from the source of illumination, light spreads out as indicated in Fig. 135, so that

though each ray retains its original intensity the number of rays which illuminate a given area depends upon the distance of that area from the luminous source *S*. At twice the distance the rays are spread over four times the area, so their illuminating effect, as at *M*, is only one fourth of what it is at *m*. The amount of light received from a luminous source is thus inversely proportional to the square of the distance from the source.

55. SHADOWS.

i. Shadows produced by small sources of light.—(*a*) Place a stick vertically between a screen and an ordinary fish-tail gas burner, so that the flat flame and the stick are in the same plane. The shadow of the stick on the screen is defined sharply. Turn the flame through a right angle, so that it is now parallel to the screen. The dark shadow is fringed by another less intensely black.

(*b*) Cast a shadow of a sphere on to a screen, using a small source of light, such as a candle-flame. Notice that the shadow cast on the screen is distinct, circular, and of equal darkness throughout.

ii. Shadows produced by large sources of light.—(*a*) Substitute a lamp, with a ground glass globe larger than the sphere, for the candle in the last experiment. Notice that the shadow on the screen is made up of two parts, an inner dark circular patch called the *umbra*, while concentrically arranged round it is a partially illuminated shadow, forming a ring, called the *penumbra* (Fig. 136).

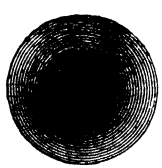


FIG. 136.—Umbra and penumbra.

(*b*) Using the lamp with a large globe, as in the last experiment, cast a shadow of a very small sphere. Notice that the shadow comes to a point, as can be shown by moving the screen slowly from the sphere, when the shadow gradually becomes smaller and disappears. This is a *converging* shadow, while those of two previous experiments are *diverging* shadows.

Shadow of a rod.—When a thin rod is illuminated by the edge of the flame of an ordinary fish-tail gas burner, the shadow of the rod thrown on a screen has sharp edges, and it is equally black throughout. This, in common with all the phenomena of shadows you have studied experimentally, is the result of the fact that light travels in straight lines. The light from the edge of a flame is in the place of the illuminated pin-hole, and the explanation offered for the formation of the image by a small aperture applies here also—except that there is no crossing of the rays and no inversion.

Umbra and penumbra.—When in experimenting with the rod the flame is arranged so as to be parallel with the screen, the dark distant shadow—the umbra—is fringed on each side by a less distinct shadow—the penumbra. Similarly, when a

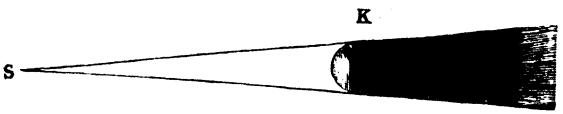


FIG. 137.—When the source of light is relatively small the shadow of an object has no penumbra.

sphere is illuminated by a small source of light, such as a candle flame, it casts a clearly defined shadow on the screen, or the umbra only is present (Fig. 137). When, however, the source of light is larger, like the ground glass globe of the lamp, the umbra is surrounded by a partially illuminated concentrically

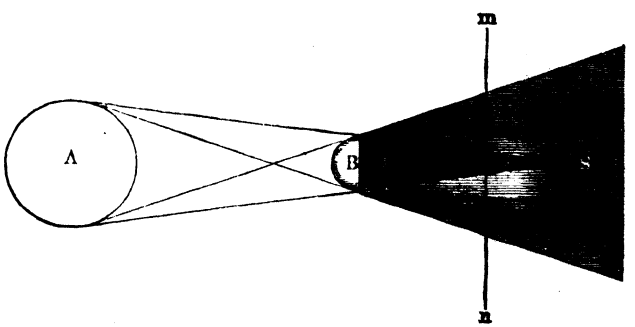


FIG. 138.—Formation of umbra and penumbra of a shadow.

arranged shadow—the penumbra. In Fig. 138, *A* represents the illuminated globe, *B* the sphere, and *mn* the screen. By following the course of the rays of light the student will have no difficulty in understanding

(1) The formation of umbra and penumbra.

(2) How these result from the propagation of light in straight lines.

56. PHOTOMETRY.

i. Law of inverse squares.—Pin a piece of white paper upon a drawing board to act as a screen. Fix the drawing board at right angles to a table in a darkened room. In front of the screen place a vertical rod about 1 to 2 cm. in diameter (a retort stand will do). Beyond this, place to one side, a candle fixed on a piece of wood, and to the other side two candles, one immediately in front of the other, fixed on a block of wood. Notice that two shadows of the upright appear on the screen. Move the candles near each other so that the two shadows of the upright touch but do not overlap. Notice that one shadow, that cast by the two candles, is darker than the other. The latter shadow is illuminated by one candle, the other is illuminated by two candles. Now, move back the two candles until the shadows appear equally dark, that is, in equal contrast with the bright part of the screen. Then the two candles give just as much light to the screen as the one. Measure the distance of the one candle, and the mean distance of the two. Compare these distances; are they in the ratio 2 : 1? Compare also the squares of the distances.

Do the experiment with different distances, and again compare the squares of the distance. Hence prove that the *illumination is inversely proportional to the square of the distance.*

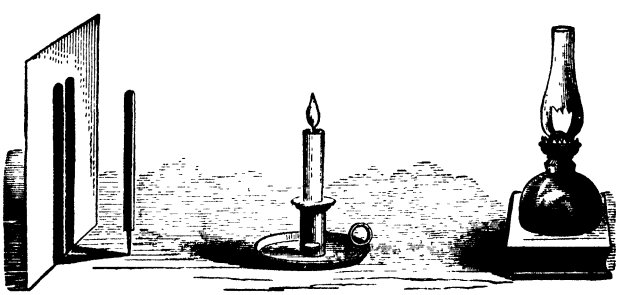


FIG. 139.—Comparison of the intensities of the flame of a candle and lamp.

ii. The shadow photometer.—Using the same screen and rod as before, compare the illuminating power of a candle flame with that of a fish-tail gas burner or a lamp (Fig. 139). Place the candle, which can be conveniently fixed on a flat piece of wood, at a measured distance from the screen, say 30 cm. Move the lamp away from the screen until the shadows cast by the candle flame and the lamp flame are of equal intensity. To estimate this accurately you must arrange the flames so that they are the same height from the table, and in such positions that the two shadows on the screen just touch but do not overlap one another.

The relative darkness of the shadows is more easily seen if the eyes be contracted or half closed, especially when, as here, the light is of somewhat different colours.

Measure the distance of the lamp flame from the screen. Vary the distance of the candle from the screen, and find the corresponding proper position of the lamp flame. Record thus :

The shadow photometer (Rumford).

Distance of Candle from the Screen.	Distance of Lamp Flame from the Screen.
1	.....
2	.....
3	.....
4	.....

Compare the squares of these distances. These are evidently the ratios of the relative illuminating powers of the two sources of light; and hence the comparison shows the *candle power* of the lamp or gas flame used.

Bunsen's photometer.—(*a*) Obtain a piece of white paper. Make a grease spot in the centre. Allow a light to shine on the paper. Observe that the grease spot is darker than the surrounding surface. Observe the paper by transmitted light. Notice that the grease spot is now brighter than the general surface.

(*b*) Use the paper as a screen and illuminate one side of it by means of a candle and the other with a lamp. Move the candle and lamp until the grease spot is barely distinguishable in point of brightness from the white surface near it. Measure the distance of the candle and lamp from the grease spot. Using the law of inverse squares, calculate the luminosity of the lamp in terms of the candle.

Photometry.—The principle already noticed, that the amount of light received from a luminous source is inversely proportional to the square of the distance, is of great value in determining the relative brightness of two sources of light, or of measuring any light in terms of a standard light.

In **Rumford's shadow photometer** (Fig. 139), the shadow cast by one source of light is illuminated solely by the other source of light. When the two shadows are equally dark, the illumination, due to each light, at the position of the screen on which the shadows are thrown, is the same. By comparing together the squares of the distances of each light from the screen, the relative intensities of the two lights are determined. If the distance of the candle is 10 inches and that of the lamp

20 inches, the intensity of the candle would be represented by $10 \times 10 = 100$, and the lamp by $20 \times 20 = 400$, or the lamp is four times as bright as the candle.

In **Bunsen's grease-spot photometer** the two sources of light are compared by placing them on either side of a screen having a grease spot in it. Its use depends on the fact that a grease spot equally illuminated on either side has the same brightness as

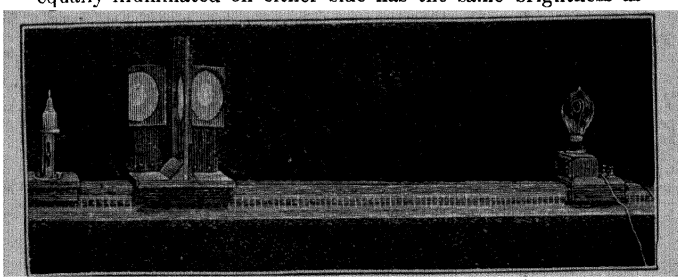


FIG. 140.—Bunsen's grease-spot photometer.

the general surface. The light lost by transmission through the translucent, or semi-transparent spot in one direction is compensated for by the equal transmission in the opposite direction. The eye is indifferent whether the light it receives is due to a reflection from white paper or to transmission through a translucent spot. It observes that the brightness of the grease-spot is equal to that of the rest of the surface. The intensities of the two lights are again proportional to the squares of their distances from the screen.

57. LAWS OF REFLECTION.

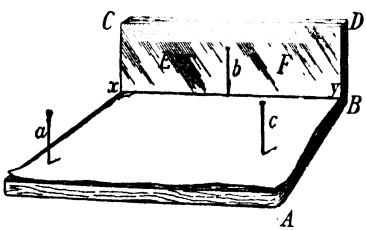


FIG. 141.—Arrangement of pins and mirror for proving the laws of reflection of light.

1. Pin method of proving the laws of reflection.—Fix two slabs of wood at right angles, as in Fig. 141, *AB*, *CD*. Against the upright slab put a piece of glass *EF* with blackened back, so that reflection only takes place from its front surface. Upon the horizontal slab lay a sheet of white paper.

Stick a pin *b* in the wood against the glass, and place another pin in the position *a*. Now procure a third pin and stick it into the wood at *c* in such a position that *c*, *b*, and the image of *a* are in a straight line. Draw with a finely pointed pencil a line along the edge of the glass *xy*; then take glass and pins away.

The paper will be marked by the pin-holes and the line *xy*. Draw lines through the pin-holes, and at *b* erect a normal *bd* to *xy*, that is, a line perpendicular to *xy*. Measure the angles *abd*, *cbd*, and compare them (Fig. 142.) Repeat the experiment two or three times, with the pins in different positions, and so determine that the angle of incidence and the angle of reflection are equal.

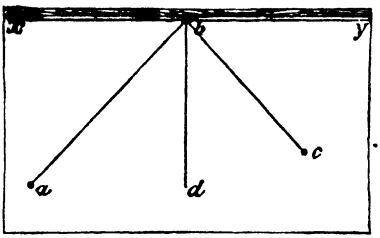


FIG. 142.—The angle of reflection *cbd* is equal to the angle of incidence *abd*.

Observe that since the holes made by the pins are all on the same piece of paper with the normal, *the incident ray, the normal, and the reflected ray are all in the same plane.* Moreover, the reflected ray is on the opposite side of the normal to the incident ray.

ii. Laws of reflection illustrated by a mirror.—Fasten with wax a little whitened, wooden stick perpendicularly at the centre of a

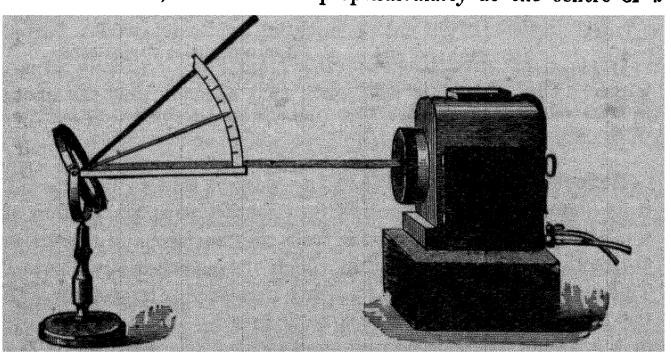


FIG. 143.—To illustrate Expt. 57 ii.

plane looking-glass. Cast upon the mirror, at the foot of the rod, a beam of parallel rays from the lantern, or a sunbeam coming through a hole in a screen. Notice (*a*) that the reflected beam always makes the same angles with the mirror and the stick as the incident beam does, and (*b*) that the incident beam, the stick, and the reflected beam all lie in one plane (Fig. 143).

iii. Reflection at two surfaces.—Place a lighted candle in front of a thick plate-glass mirror, as in Fig. 144. Notice that two images can be seen when viewed a little from one side. One is due to reflection at the front surface, and the other is produced by reflection at the silvered back surface.

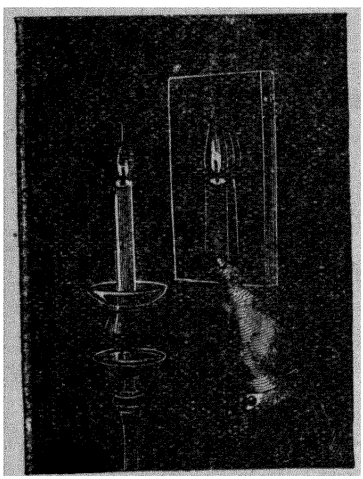


FIG. 144.—Multiple images produced by a thick mirror.

iv. Images formed by plane mirrors.—(*a*) Place a knitting needle in front of a plane unsilvered sheet of glass fixed vertically in front of a dark background. Arrange another such needle behind the mirror in such a position that wherever the eye be placed, the needle behind the mirror always appears in the position of the reflected image of the first needle.

It is easy to arrange the needle in the right position behind the mirror, by observing that when this needle apparently moves more in the direction in which the eye is moved than the image does, the needle is too far behind the mirror, and *vice versa*.

Measure the distances of the two needles from the back of the mirror. They should be the same, thus showing that the image is situated as far behind a plane mirror as the object is in front of it.

Reflection of light.—When any wave is said to be **reflected,** it is understood that it comes into contact with the surface of some body, and is thrown back from that surface, and travels in a direction opposed to that in which it was originally moving. This may happen in two ways, either **regularly** or **irregularly.** In the first case, it is turned back according to simple rules, while in the second there is no uniformity about the direction. The page on which this explanation is printed appears to be white because—owing to the roughness of the paper—of the irregular reflection of the light which falls upon it. Or, if we powder a sheet of glass, the powder seems to be white for a similar reason; there are many surfaces formed from which irregular reflection takes place. The manner in which a sheaf of rays is

broken up and leaves the surface of a body when reflected irregularly can be seen in Fig. 145.

Laws of reflection of light.—Light is reflected regularly from a plane mirror—that is, a flat reflecting surface. Such a mirror can be made from a variety of substances, but the most common is bright metal or silvered glass.

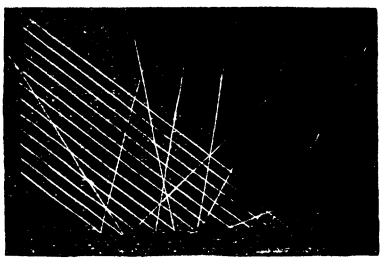


FIG. 145.—Irregular reflection of light.

The angle at which the light, or any sort of wave, strikes the reflecting surface is called the **angle of incidence**, and the wave an **incident wave.** The angle at which the wave leaves this surface is known as the **angle of reflection,** and the wave as it leaves the **reflected wave.**

There is a definite connection between the angles of incidence and reflection, and it can be expressed as follows:

1. **The line representing the reflected wave is in the same plane with the normal and the line representing the incident wave, and is on the opposite side of the normal to the incident line.**

2. **The angle of incidence is equal to the angle of reflection.**

It has also been learnt by experiment that when a wave strikes a reflecting surface *normally*, *i.e.* having travelled along the normal, it is reflected back upon the same line.

Formation of an image by a plane mirror.—The two rules just enunciated enable the formation of an image by a plane mirror to be understood easily.

Let MM (Fig. 146) be the plane mirror, and A a bright spot of light like the head of a pin in Expt. 57 i. First see what happens to the light ray which leaves A and strikes the mirror normally. It is reflected back along the same line, and the reflected ray appears to come from a point A' along AA. In the case of any other rays, such as AB, the light ray is reflected in such a way that the angle of reflection CBD is equal to the angle of incidence ABC, and appears to an eye, placed as in Fig. 146, to come along BD from a point A', where BD produced meets AA'. If the same construction is made for any other ray AB' it will

be reflected and appear to be coming along $B'D'$, which produced backwards will pass through the same point A'. A' is therefore the image of A, and it can be proved easily by geometry that A' is as far behind the mirror as A is in front of it.

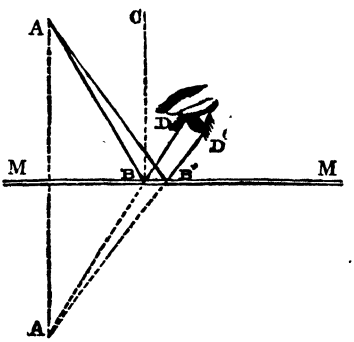


FIG. 146.—Representation of the course of rays producing an image in a plane mirror.

These facts apply equally to the formation of the images of *objects*, which can be considered as accumulations of small material particles to which the construction given above for a point may be applied.

When a mirror rotates, the angle through which the image moves is twice the angle through which the mirror moves.—Knowing the laws of reflection, this truth can be arrived at easily by a knowledge of geometry. Thus let MM (Fig. 147) be a mirror, capable of rotation, when in a vertical position. After a brief interval it will, when rotating, have arrived at a new position M_1M_2. PO is a ray of light which strikes the mirror in its vertical position at right angles.

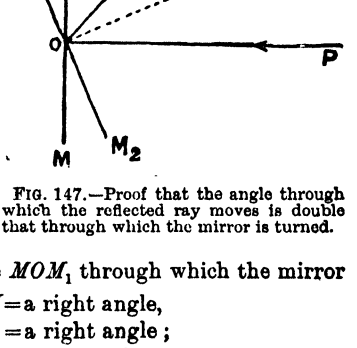


FIG. 147.—Proof that the angle through which the reflected ray moves is double that through which the mirror is turned.

The mirror rotates and the direction of the reflected ray is along OR. It is required to compare the angle POR through which the ray of light moves, with the angle MOM_1 through which the mirror moves.

Angle POM = a right angle,
angle QOM_1 = a right angle;
$\therefore$ angle POM = angle QOM_1.

The angle through which the mirror moves

$$= M_1OM = QOM_1 - MOQ$$
$$= POM - MOQ$$
$$= QOP.$$

But since angle of incidence = angle of reflection

$$QOP = QOR;$$
$$\therefore \quad POR = 2QOP.$$

But *POR* is the angle through which the ray of light moves, and it is twice that through which the mirror moves.

58. SPHERICAL MIRRORS.

i. Principal focus of a concave mirror.—Procure a concave mirror and cover it with black paper, except a small part at the centre or round the pole. That is, let the aperture of the mirror be small. Allow rays of sunlight to fall upon it (these come from so great a distance that they can be considered *parallel*). Move a very small

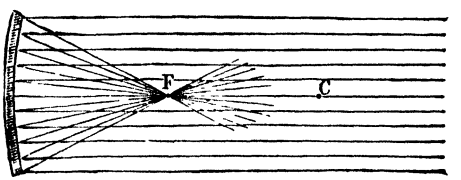


FIG. 148.—The principal focus of a spherical mirror.

paper screen up and down in front of the reflecting surface so as not to cut off the incident rays. Notice that at a certain point a clear image of the sun is formed, and probably the screen will be burnt.

ii. Concave mirrors ; law of distances.—(*a*) Place a lighted candle in front of a concave mirror so that the flame is on the principal axis. Move a small screen of white cardboard to and from the mirror, so that it does not intercept all the light passing between the candle and the mirror. At a certain distance from the mirror a clear image of the flame will be found upon the card screen.

(*b*) Now move the flame a short distance away or towards the mirror. It will be found necessary to move the card towards or away from the mirror in order to again obtain a sharp image.

The distances from the mirror to the candle flame u and from the mirror to the card v should be carefully measured. They will be found to depend upon the focal length f according to the equation:

$$\frac{1}{u} + \frac{1}{v} = \frac{2}{r}.$$

Reflection from Spherical Mirrors.—A spherical mirror is a part of a spherical surface which has the power of reflecting. It may be either *concave* or *convex*, the former if the reflection takes place from the hollow side, the latter if from the bulging side. The centre of the sphere of which it forms part will evidently be the centre from which the part of it constituting the mirror was struck, and this point is called the **centre of curvature**. The distance from this point to the reflecting surface is the **radius of curvature**. Thus in Fig. 149 *c* is the centre of curvature and *cM*, *cd*, *cM'*, are all radii of curvature. *MM'* is called the diameter or aperture of the mirror and *d* is called by many names, perhaps *pole* of the mirror is the best. A line going through the pole and centre of curvature is the **principal axis** of the mirror, any other radius produced being a *secondary axis*. We know from geometry that every radius is at right angles to the tangent at the point where it cuts the circle; and, since we can consider the tangent and circle as coincident at this point, from what has been already said it will be clear that the radii are normals to the mirror. Evidently, then, if we place a luminous object at the centre of curvature we shall have all the rays of light from it reflected back along the lines of incidence, or the image will be formed at the same place as the object.

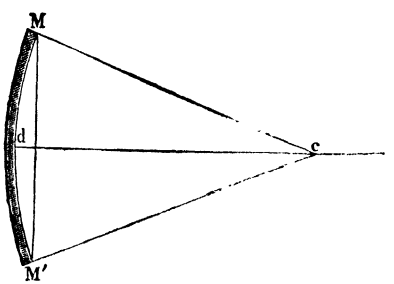


FIG. 149.—Reflection from a spherical mirror.

If parallel rays fall on a concave mirror, *e.g.* rays from the sun, they will be reflected and brought to one point, called a focus. The point so obtained is called the **principal focus** of the mirror. In Fig. 148, *F* represents this point and *C* the centre of curvature. The parallel lines show the direction of the sun's rays. This point *F* is midway between the pole and *C*, or the *focal length* is half the radius of curvature.

CHIEF POINTS OF CHAPTER XIV.

Light, like every kind of radiation, is a form of energy. It is a process of transference of energy by ether-waves. Those ether-waves which affect the retina are known as light.

Rectilinear propagation of light.—Light travels in straight lines when propagated through any one medium, but often has its direction changed when passing from one medium to another (see Refraction).

Consequences of rectilinear propagation of light.—(*a*) The images produced by a pin-hole camera are inverted.

(*b*) The size of the image formed by such a camera can be found:

$$\frac{\text{length of object}}{\text{length of image}} = \frac{\text{distance of object from aperture}}{\text{distance of image from aperture}}$$

(*c*) Illumination is due to the overlapping of images.

(*d*) The formation of the umbra and penumbra.

A photometer is an instrument for comparing the intensities of different sources of light.

Reflection.—Light, in being reflected from suitable surfaces, obeys the following laws:

1. The reflected ray lies in the same plane as the incident ray and the normal; and is on the opposite side of the normal to the incident ray.

2. The angle of reflection is equal to the angle of incidence.

Spherical mirrors.—When parallel rays, or those from a very distant object, fall on a concave mirror, they meet, after reflection, at a point half-way between the mirror and its centre of curvature. This point is called the principal focus. The law of the distances of the object and of the image respectively from the mirror in other cases is expressed by a simple equation.

EXERCISES ON CHAPTER XIV.

1. A candle is placed close to a looking-glass, and its reflection in the glass is viewed a little from one side. What will be seen? As far as you can, explain the appearances which you describe.

2. Describe a pin-hole camera, and explain, illustrating your answer by a diagram, how the image of a luminous object is formed by it.

What experiment would you perform to show why it is that the image first becomes blurred and then disappears when the size of the hole is gradually increased?

3. Three candles are placed quite close together in a row at the centre of a room, and a wooden rod is held in a vertical position at a distance of about a foot from the candles. Explain, giving

diagrams, why it is that as the rod is moved in a circle round the candles the shadow cast on the walls is in some positions sharp and in others very ill defined.

4. The sun shines through a crack in the shutter of a darkened room. A person inside the room says that he sees a ray of light entering the room. Put his statement in a more accurate form. What can he really see ?

5. A small opaque sphere is placed between a gas burner and a white screen, and when the gas is turned down so that the flame is very small it is found that the shadow cast on the screen is quite sharp, but on turning up the gas so that the flame is large, that the edge of the shadow is blurred. Explain the reason for this change, illustrating your answer by means of diagrams.

6. State the two laws in accordance with which a ray of light is reflected by a smooth surface, and describe experiments by which you would demonstrate the truth of each of these laws.

7. What is an inverted image ? If the capital letter F were drawn on paper and held in front of a mirror, how would you have to draw the letter on the paper and how hold the paper, in order that the image of the letter in the mirror should present its ordinary aspect ?

8. A bright object is placed at a short distance in front of an ordinary looking-glass. An eye looking into the mirror sees, in general, a number of images of the object, the second of the series in order of nearness being usually the most brilliant. Explain this.

9. A shadow is usually divided into an "umbra" and a "penumbra." Explain these terms. How can you readily obtain (1) a shadow which is practically all umbra, (2) a shadow which is all penumbra ?

10. By moving a fragment of looking-glass, a boy finds that he can throw images of the sun up and down the walls and ceiling of a room. Where must he stand to be able to do this ? Show by a diagram that the angle through which the image moves is twice as large as the angle through which the boy moves the glass.

11. Describe how parallel rays are reflected by a concave mirror. What is the law of distances in the case of such a mirror ?

*

CHAPTER XV.

REFRACTION OF LIGHT.

59. REFRACTION AT PLANE SURFACES.

1. Refraction by water.—(*a*) Procure a rectangular metal box, such as a cigarette box, and put a wooden or metal scale on the bottom. In a darkened room let sunlight fall slantwise against the

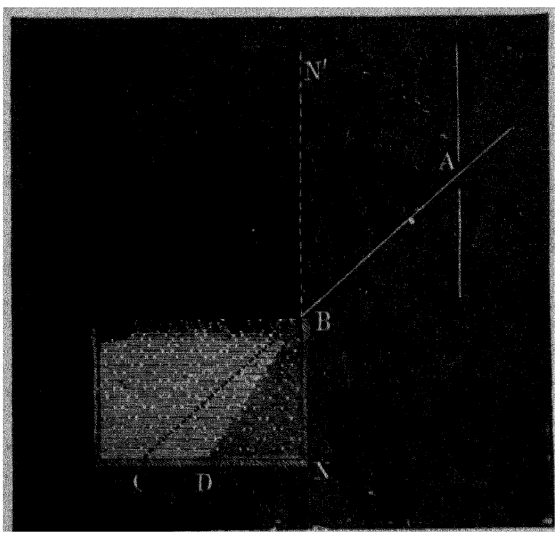


FIG. 150.—Experiment to show refraction of light by water.

edge. The side of the box throws a shadow which reaches, say, to *C*, which, since light travels in straight lines in the same medium, is a continuation of the ray of sunlight *AB*. Without disturbing anything fill the box with water. The shadow no longer reaches to *C*, but only as far as *D*. Clearly the light waves have been bent or *refracted* out of their original course. Notice that *NN'* is the

normal, and that the rays of light travelling from the comparatively rare air into the denser water are refracted towards the normal.

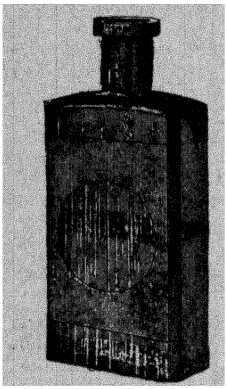


FIG. 151. — Bottle for illustrating refraction of light by water.

(*b*) Obtain a medicine bottle with flat faces, and paste over one face a piece of paper having a circular hole cut in it (Fig. 151). On the glass draw a vertical and a horizontal line on the clear space as shown. Pour water into the bottle until it reaches the level of the horizontal line. Let a narrow beam of light enter the side of the bottle and strike the surface of the water where the two lines cross; it will be found to be bent in the water towards the vertical line.

ii. Pin method of proving laws of refraction.—(*a*) Upon a piece of board *ABCD* (Fig. 152) place a sheet of paper, and upon the paper put a piece of fairly thick glass with parallel sides (a thick piece of glass from a box of weights, a paper-weight, or a number of slips of microscope glass will do very well). Rule along the edges of the glass with a finely-pointed pencil. Place two pins, *a*, *b*, as shown in the illustration, and then, *looking through the glass* from the other side, stick in the pins *c*, *d*, so that all four appear in a straight line.

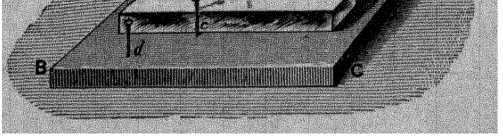


FIG. 152.—Pin method of illustrating refraction of light.

(*b*) Now take away the glass and pins and join the pin-holes on the paper as shown in Fig. 153. Draw the normal *ebf* and the circle *elfg*. Draw *lm* and *gk* perpendicular to *ebf*, and compare the lengths *lm*, *gk*. Obtain the ratio $\frac{lm}{gk}$ for different positions of the pins; it will be found practically the same in all cases in which the same material is used.

Notice that the direction of the ray *cd*, emerging from the glass, is parallel to that of *ab*.

iii. Phenomena caused by refraction.—(*a*) Place a bright object, say a coin, on the bottom of an empty basin, and arrange your eye

so that the object is *just* hidden by the edge of the basin (Fig. 157). Get somebody to pour water into the basin. You will now be able to see the coin without any movement of your eye having taken place. Evidently there has been a bending of the direction of the light rays somewhere.

(*b*) Place a glass cell in front of a white surface brightly illuminated. Let the surface of the water be visible. Put a lump of ice on the water, and observe the streaky appearance of the

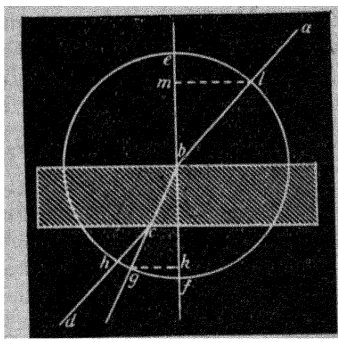


FIG. 153.—Construction for proving the laws of refraction by the pin method.

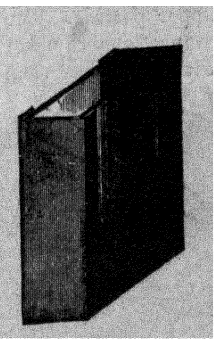


FIG. 154.—Cell with glass sides for refraction experiments.

illuminated part of the screen. Add syrup, alcohol, and hot water by a pipette. Notice that the light has not passed through uniformly.

(*c*) Observe the shadow of the disturbed air above a stove or hot poker. Notice the apparent shifting of position of objects seen through such air.

(*d*) Fill a glass cylinder with water, and place a coin at the bottom. On looking straight down through the water the coin appears nearer the surface than it really is. Hold another coin near the outside of the cylinder, and place it at such a height that the two coins appear at the same level. The amount by which the coin in the water is apparently elevated by refraction can thus be found. The length of the column of water through which the coin is observed, divided by the distance from the top of the water to the outside coin, gives the index of refraction of water.

(*e*) Repeat the experiment with methylated spirit.

Refraction of light.—Up to the present the light rays have been supposed to be moving through a uniform medium. When this is so, as has been seen, light travels in straight lines, and, if it meets a reflecting surface, it is turned back, according to the laws which have been studied. If, however, the light passes

from one medium into another of a different density the propagation of the wave is no longer rectilinear: the passage from one medium into the other is accompanied by a bending of its path. This bending is known as **refraction**, and the ray is said to be **refracted**.

Rules of refraction.—In Fig. 155 the shaded lower part of the diagram represents a denser medium than the unshaded upper portion. The word denser is used here, and in similar connections, to mean optically denser, and must not be confused with what has been said of the density of bodies in Chapter II. Let *RI* represent a ray passing from the rarer to the denser medium, or the ray incident on the surface of the denser medium at *I*. The angle *RI* makes with the normal at *I* is the angle of incidence. The ray is bent; instead of continuing its course in a straight line along *IR'*, it is refracted and travels in the direction of *IS*, which represents the refracted ray, the angle *SIP* being the angle of refraction. The angle *R'IS*, which represents the amount the ray has been turned out of its original path, is called the **angle of deviation**. With the centre *I* and any convenient radius describe a circle, and from the points where it cuts the incident and refracted rays, perpendiculars on to the normal are drawn as in Fig. **155**. A perpendicular is also dropped from the point *R'*. It is clear from geometry that *R'P'* is equal to the perpendicular let fall on to the normal from the point where the incident ray cuts the circle. The ratio between the lengths of *R'P'* and *SP* is constant for the same two media, *e.g.* air and water, whatever the angle of incidence. This ratio is called the **index of refraction**. Its value for air and water is about $\frac{4}{3}$; for air and glass roughly $\frac{3}{2}$, depending upon the kind of glass.

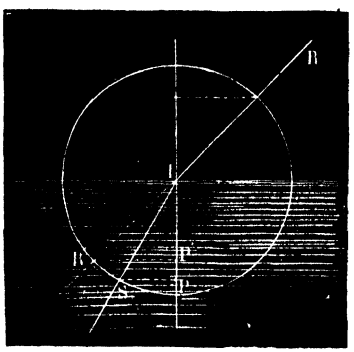


FIG. 155.—Geometrical representation of laws of refraction.

The laws of refraction are then :

1. **The incident ray, the normal, and the refracted ray are all in the same plane. The incident and refracted rays are on opposite sides of the normal.**

2. **If a circle be described about the point of incidence, and perpendiculars be dropped upon the normal from the intersections of this circle with the incident and refracted rays, the ratio of the lengths of these perpendiculars is constant for any two given media.**

Refraction through a plate with parallel sides.—In the case of a ray of light passing completely through a plate of glass, having its sides parallel, the ray is bent towards the normal when entering the glass, and away from the normal when emerging, so that its course is as shown in Fig. 156. The ray is thus displaced laterally, but it emerges in a direction parallel to its original direction. The effect of refraction is in this case to cause the point S to be seen at S'. In constructing a diagram for any similar case, it should be borne in mind that when a ray passes from a less dense to a more dense medium, it is bent *towards* the perpendicular to the separating surface, and when a ray passes from a dense to a rarer medium it is bent *away from* the perpendicular to the separating surface.

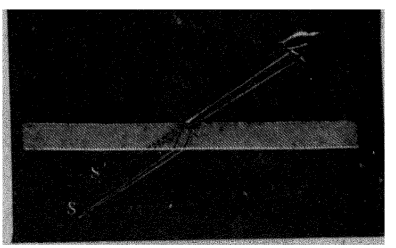


FIG. 156.-Refraction causes the object S to be seen at S'.

Various effects of refraction.—The common experiment with a coin at the bottom of a basin, where, being first hidden from view by the edge of the basin, it becomes visible again when water is poured into the basin, is explained easily by tracing out the path of the light.

In Fig. 157 let c be the position of the coin, in which it is just hidden so far as the eye is concerned by the edge of the empty basin. If the rays from the coin c be continued in straight lines these lines will evidently pass above the eye. Now when the water is put in, these rays, which before miss the eye, are refracted in passing out of the water and just enter the eye,

making the coin appear to be in the position c'. The right side of the vessel, if continued upwards, represents the normal, and evidently in passing out of water into air the light-waves are bent away from the normal.

The same kind of construction applies to the case when an object is seen obliquely through a thick glass plate (Fig. 156). The object is displaced and also made to appear nearer than it really is.

The streaky appearance seen when the light from a lantern is passed through water containing ice is due to the repeated refraction of the light on passing into layers of water of

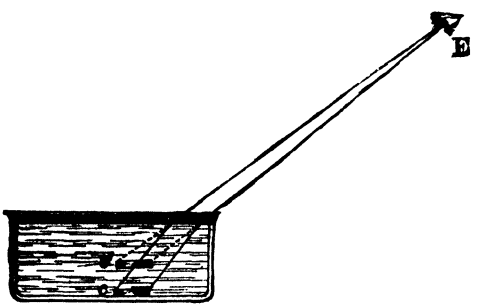


FIG. 157.—When the trough is empty the coin c is invisible to an eye at E, but when it is full of water the coin is seen at c'.

different densities (Expt 59 iii. (b)). This reason, too, explains the results obtained when alcohol or syrup are mixed with water.

The refraction of light in its passage from one medium into another of different density explains several other very familiar observations. A stick held in anything other than a perpendicular position in water appears to be bent upwards. If a straight stick is fixed upright in water and looked at from a point a few feet above the end, it will appear shorter than it really is, in the proportion of three to four, so that if a length of four feet is under water, it appears to be only three feet long. In the same way standing bodies of water always appear shallower than they really are on account of refraction. A pool of clear water when viewed from a point vertically above the surface only appears $\frac{3}{4}$ of its actual depth.

60. REFRACTION THROUGH A PRISM.

1. Pin method of tracing deviation by a prism.—Stand a prism upright, that is, upon one of its ends, upon a piece of white paper. Stick two pins into the paper in the positions Q, S (Fig 158), place two more P, R, on the opposite side of the prism, so that the four

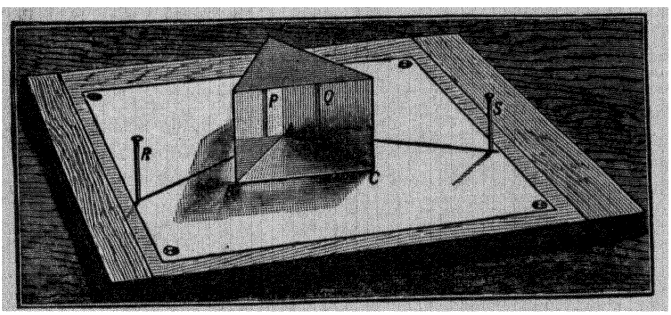


FIG. 158.-Pin method of tracing deviation of light by a prism.

appear in a straight line when looking through the prism. Draw the outline of the prism ABC, and then take away the prism and the pins and connect the pin-holes as shown in the diagram. It will be found that the ray is bent towards the base of the prism both when it enters and emerges.

Refraction through a prism.—When a wedge-shaped piece of glass, or a prism, as it is called in optics, is interposed in the path of a ray of light from a small hole in the cap of a lantern, it is easy to see by watching the image of the hole on a screen that the image moves in a direction towards the base of the prism. This is because the ray is bent by its passage through the prism, so that on its emergence from the glass it continues in a new path inclined towards the base of the prism. The amount of bending experienced by the ray of light depends upon :

(1) The angle between the inclined sides of the prism meeting in its edge, or the angle of the prism, as it is called ;

(2) the material of which the prism is made, and

(3) the nature of the incident light.

If two prisms of the same angle and material are so arranged that the edge of one adjoins the base of the other, the second

prism undoes the bending of the first, and the ray as it leaves the combination of prisms continues its path in a direction parallel to the incident ray, though not, of course, in the same straight line with it.

Path of a ray of light through a prism.—In Fig. 159 let the triangle *abc* represent a section of the prism at right angles to its faces, such as we should see by looking at the end of it. Suppose *DE* is a ray of light striking the face *ab* of the prism. The light on entering the prism passes from the air into the glass, or *from a rarer into a denser medium*, and is bent *towards* a line drawn perpendicular to the face of the prism at the point where the ray of light strikes it. It consequently travels along the line *EE'* until it reaches the face *ac* of the prism. Here it passes from the glass into the air, *i.e. from a denser into a rarer medium*, and is, in such circumstances, bent *from* the perpendicular, and travels along the line *E'D'*. In every such passage through a prism it is noticed that the light is always bent or refracted towards the thick part of the prism.

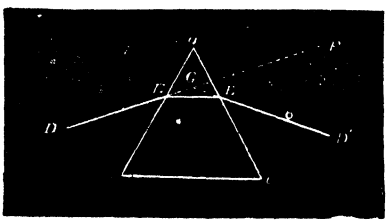


FIG. 159.—Construction to show the path of a ray of light through a prism.

61. REFRACTION THROUGH A LENS.

i. The principal focus of a lens.—To find the distance of the principal focus from the centre of the lens, that is, its *focal length*, it is only necessary to form an image of the sun by it on a screen and to measure the distance between the lens and the screen.

ii. Convex lens ; law of distances.—Place a candle flame on one side of a convex lens and arrange a cardboard screen to catch the image on the other side. When the image is quite sharp, measure the distances of flame and screen from the lens. Using letters u, v, and f as with concave mirrors (p. 217), obtain several pairs of values for u and v, the lens being moved between each determination. Measure the focal length f as in Expt. 61 i.

It will be found that u, v, and f have values which satisfy the equation:

$$\frac{1}{u} + \frac{1}{v} = \frac{1}{f}.$$

It should be remembered that u is the distance from the flame to the lens.

iii. The simple microscope.—If the flame be brought towards the lens, the screen must be moved away from it. At last the image cannot be obtained on the screen no matter how far off it is carried. This happens when the flame is at the principal focus of the lens. In this case and when the flame and lens are still nearer one another, no image is formed, for the rays of light now emerging from the lens are *diverging*. This is what happens when a lens is used as a magnifying glass or simple microscope and therefore held close to the object to be examined. The enlarged appearance is due to the divergence of the emerging rays. The image is not real, but is *virtual*, and appears to be situated on the same side of the lens as the object. A virtual image is thus one which cannot be caught on a screen.

Refraction through a lens.—Most lenses are of glass, with curved surfaces, which are portions of spheres. In some lenses one surface is quite plane. All lenses can be divided into two

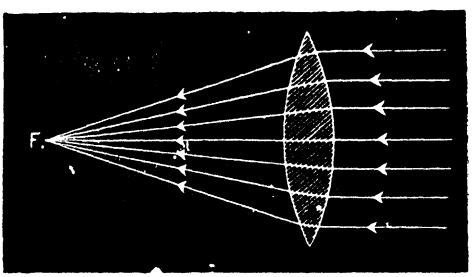


FIG. 160.—Parallel rays falling on a convex lens.

classes—**convex** or **converging**, and **concave** or **diverging**. Converging lenses can be distinguished by their power of forming an image of a distant object like the sun, or by that of magnifying. Concave lenses form no image in this way, and, moreover, instead of magnifying, they make objects appear smaller when viewed through them.

To understand their action upon the course of rays of light through them it is simplest to regard them as being built up of parts of prisms in contact, as shown in Fig. 161, where a convex lens is built up in this way. A ray of light falling upon any one of these prisms is refracted towards its thicker part, and consequently they all converge towards a point, which, if the incident rays are parallel, is known as the **principal focus** of the lens, as *F* in Figs. 160 and 161.

The photographic camera in its simplest form consists of a

convex lens and ground glass screen which can be moved towards or away from the lens until a sharp image of the object to be photographed is obtained. The screen is now replaced by a prepared glass plate, coated with a substance, generally a

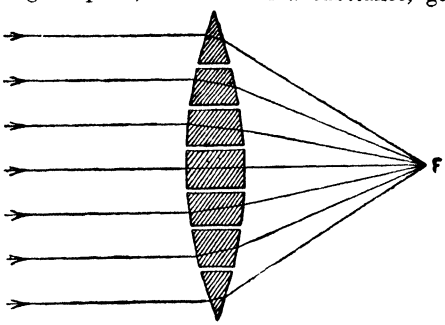


FIG. 161 —Lens built up of parts of prisms.

compound of silver, which is extremely sensitive to the action of light. The plate is exposed to the light passing through the lens for a space of time varying from 1000th of a second to several minutes, according to the intensity of the light forming the image. The image is not visible until the plate has been "developed" and "fixed." In this way a **negative** is obtained in which light-coloured objects appear dark and *vice versa*. The

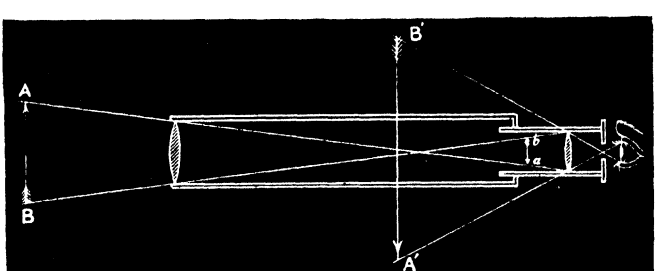


FIG. 162.—Formation of image by telescope.

positive or ordinary photograph is made by printing through the negative on sensitised paper.

The telescope.—The principle of the refracting astronomical telescope can now be described. Notice in Fig. 162, the larger

double convex lens forming an image *ab* of the object *AB*. This image becomes the object to a second smaller, though in other respects similar, lens, placed near the eye. The second lens forms an enlarged image *A'B'* of this first image *ab*, in the way shown in the figure. A lens placed in the same relation to the smaller one as the larger in the figure is called the *object glass*, while the smaller one which magnifies the image formed by the first constitutes the *eye-piece*.

CHIEF POINTS OF CHAPTER XV.

Refraction of light.—A ray of light passing from a less dense to a more dense medium is bent towards the perpendicular to the separating surface, and conversely. The laws of refraction can be stated thus:

1. The incident ray, the normal, and the refracted ray are all in the same plane. The incident and refracted rays are on opposite sides of the normal.

2. If a circle be described about the point of incidence and perpendiculars be dropped upon the normal, from the intersections of this circle with the incident and refracted rays, the ratio of the lengths of these perpendiculars is constant for any two given media.

Refraction by a prism.—The deviation of a beam of light caused by the action of a prism depends upon (*a*) the angle of the prism, (*b*) its material, (*c*) the nature of the light.

Refraction by a lens.—Convex lenses make rays falling upon them tend to become converging; concave lenses produce divergence. A lens may be considered as built up of prisms, which are in opposite directions in convex and concave lenses. The focus for parallel rays is called the principal focus. The law of distances for lenses is expressed in a simple equation very similar to that for concave mirrors.

EXERCISES ON CHAPTER XV.

1. A bright bead is placed at the bottom of a basin of water, and a person stands in such a position that he can just see it over the edge of the basin. While he is looking, the water is drawn off. How will this affect his view?

Draw a diagram showing the direction of a ray of light passing from the bead through the water and the air in each case.

2. A thick layer of transparent liquid floats on the surface of water. Trace the course of a ray of light from an object immersed in the water through the floating liquid to the air.

3. Describe an experiment to show the path of a ray of light which passes obliquely through a thick plate of glass. Illustrate your answer by a sketch in which you indicate clearly the path of the ray in the air before it enters the glass, in the glass, and in the air beyond the glass.

4. An upright post is fixed in the bottom of a pond which is three feet deep; the top of the post is three feet above the water. How will the post appear to an eye about the level of the top of the post and four or five feet away from it?

Draw a figure to illustrate your answer.

What will be seen as the eye moves further and further back from the post?

5. Describe in detail, giving a sketch of the apparatus, an experiment to show that when a ray of light passes from water to air it is bent away from the perpendicular or normal.

6 State the laws of the deviation of light when passing from any medium to one of different optical density.

7. A boy wades in a pond which everywhere reaches to about the level of his knees. On account of the water some of the pebbles with which the bottom is covered are invisible and others are not seen in their true places. Explain this, and illustrate your answer by a diagram.

8. A dot is made on a piece of paper, and a prism is laid on the paper, over the dot. An eye in certain positions now seems to see two dots. Draw a diagram to explain this.

9. Explain why a convex lens is also called a converging lens.

10. Give short accounts of (*a*) the photographic camera, (*b*) the telescope.

11. If a thick piece of plane glass be placed with one edge resting on a page of print, the letters seen through the glass are displaced to one side or the other according to the position of the glass. Explain this.

12. Given a small source of light show how you would proceed to obtain parallel rays by using a double convex lens.

13. A person arranges a basin of water and a candle in such a way that the reflection of the candle on the water seems to him to be in the same straight line as a coin at the bottom of the water. Draw a diagram to show the necessary arrangement, and give any explanation that you think necessary.

14. A fish swims in a glass tank; a person whose eye is above the level of the water seems to see two fish. Draw a diagram to illustrate this, and give any explanations you think necessary.

15. Objects seen through hot air rising from a chimney are noticed to flicker. Account for this fact.

16. Common glass in a window pane distorts objects seen through it, while plate glass does not. Why is this? Illustrate your answer by a sketch.

CHAPTER XVI.

ANALYSIS OF LIGHT. COLOUR.

62. DISPERSION.

1. Dispersion by a prism.—In a piece of card cut a slit (P) about 2 cm. long and 1 mm. wide. Place the card, with the slit vertical, in front of a fish-tail gas flame (Fig. 163). Arrange a prism (A) on a stand, so that it is of the same height as the slit, and has its refracting edge vertical. Between the slit and the prism place a lens (L_1). Catch the light emerging from the prism by a second

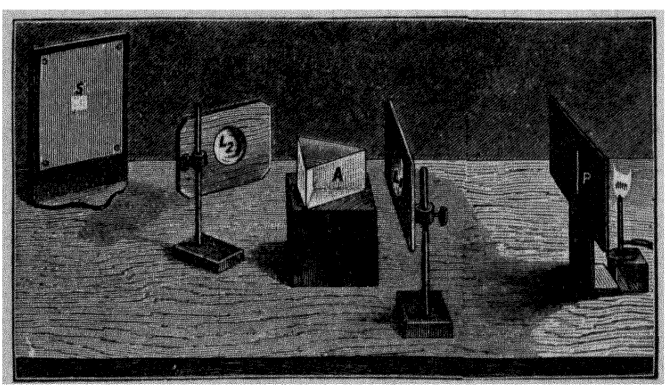


FIG. 163.—Experiment to show the dispersion of light by a prism.

lens (L_2). Move the position of the screen (S) until the coloured band of light is best seen. Observe that the light is refracted towards the base of the prism and that it is decomposed into constituent colours, which are differently bent by the prism. The violet light is refracted most and the red light least. Colours between these limits are bent by intermediate amounts. Name the colours you can see.

The band of colour is called a *spectrum*, and an instrument constructed on the principle of this experiment, a spectroscope; the light is said to be *dispersed* by the prism.

ii. Dispersion a consequence of unequal refraction.—(*a*) Place a red glass against the slit and notice that only a red image of the slit is visible. Still observing the screen, substitute a blue glass for the red one. A blue image is seen, but not in the same position as the red one; it is bent more away from the refracting edge of the prism.

(*b*) Place a second prism with its base in the same direction as that of the first. Notice a longer band of colour than in Expt. 62 i. is obtained, but it is fainter. The amount of dispersion has been increased.

(*c*) Place the second prism so that its base adjoins the apex of the first prism (Fig. 167). The band of colour disappears.

(*d*) Substitute a hollow glass prism (Fig. 164) filled with carbon bi-sulphide for the solid glass prism used in the preceding experiments. Observe the increased length of the spectrum, owing to the greater dispersive power of the carbon bi-sulphide.



FIG. 164.—A hollow prism for holding liquids to show the decomposition of light by unequal refraction.

Analysis of light by a prism.—When sunlight, or, as it is called, white light, is passed through a prism, it is decomposed or dispersed owing to the different **refrangibility** of the various

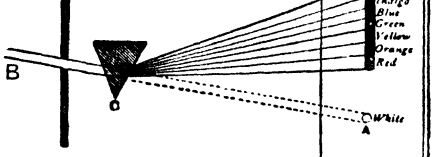


FIG. 165.—To illustrate the decomposition of white light by a prism.

kinds of light contained in it. An examination of the band of colour or **spectrum** will show that one colour shades imperceptibly into the next. There is, then, a large number of waves of different lengths comprised in the white light from the sun, and each ray is bent to an extent depending on its wave-length.

Refraction is accompanied by dispersion.—In all the cases

of refraction which have hitherto been considered, the phenomena have been described as if all the ether-waves contained in white sunlight are bent equally, but this is really not so. What is commonly called "red" light is not so much bent out of its path by a prism as "blue" light. Or, expressing the same fact in another way, red light is less **refrangible** than the blue light. It is found by accurate experiments, which the student will find described in books on light, that the sensation of red upon the retina is due to the absorption of ether-waves which

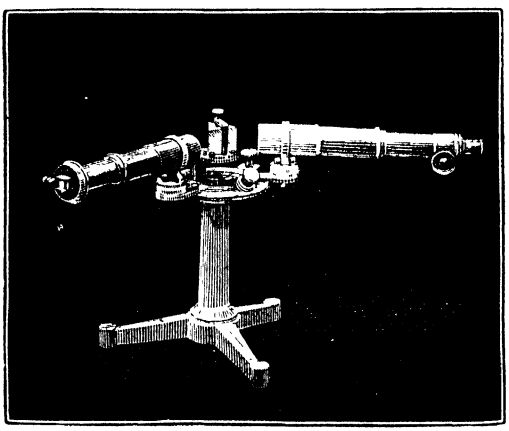


FIG. 166.—A spectroscope for the analysis of light by means of a prism.

are long and the vibration frequency slow compared with those waves which when absorbed by the retina give rise to the sensation of blue. The shortest, most rapid waves are bent most; the slowest, longest waves are bent least. The shortest, most rapid waves which affect the retina give rise to the sensation of violet. Since a prism bends ether waves of different lengths unequally in this way, it provides a means of separating waves of different lengths from one another. Because they are bent differently the ether-waves are separated or **dispersed** by the prism. This may be expressed by saying that a prism can analyse light composed of waves of different lengths.

If the decomposed sunlight, instead of being collected on a screen, be passed through a second prism similarly arranged, it

will be seen that the band is longer, or the dispersion is greater. The amount of dispersion also depends upon the material of which the prism is made. Glass produces a much greater amount of dispersion than water; flint-glass possesses twice the dispersive power of crown glass; carbon bi-sulphide, again, has even more dispersive power than flint-glass.

Although a continuous band of colour is observed when sunlight, or limelight, or a gas- or candle-flame is seen through a prism, this continuous spectrum is not always produced. For when substances such as sodium, strontium, and lithium, or their compounds, are burnt in a non-luminous flame, and observed through a prism, a spectrum consisting of bright lines, which are different for different substances, is seen. A prism may thus be used, and is used, to analyse light. The light of incandescent sodium vapour, produced by burning common salt in a flame, when observed through a prism is characterised by a yellow line, and the lights emitted by other substances when burning are each distinguished by rays of particular colours and positions in the spectrum.

63. RECOMPOSITION OF WHITE LIGHT.

1. Recomposition of light by a second prism.—Place a prism in front of an illuminated slit, as in Expt. 62 1. Observe the band of

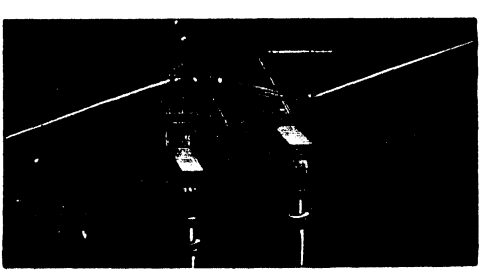


FIG. 167.—Recomposition of light by a second prism.

colour, or spectrum. Place a second prism against the first one, but with the refracting edge adjacent to the base (Fig. 167). This prism undoes the work of the first one, and no spectrum can now be seen, but only the illuminated slit.

ii. Recomposition by colour disc.—Upon a round piece of card paint sectors of the different colours contained in the spectrum, arranging the areas of the coloured sectors as nearly as possible in the proportion in which they occur in the spectrum.

Place the card upon a whirling table (Fig. 138) or upon a top, and rotate it rapidly, when it will be found that light from the card gives rise to the sensation of white or grey.

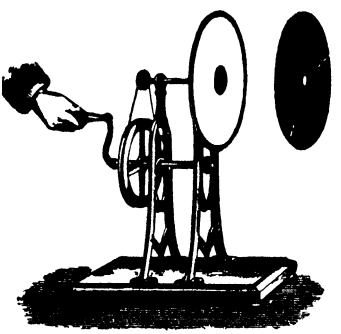


FIG. 168.—Rotating disc, to show the recomposition of the constituent colours of white light.

Formation of white light from its constituents.—Just as it is possible to analyse white light, splitting it up into its constituent colours or wave-lengths, so, by suitable arrangements, these separated or dispersed colours can be made to recombine, forming white light over again. This building up, or synthesis, of white light can be effected in the following ways :

1. By interposing a second prism with its angle reversed The dispersion of the first prism is neutralised, and the beam of light leaves the second prism in a direction parallel to the beam incident upon the first prism.

2. By the colour disc.

The colour disc.—The explanation of the recombination of the separate colours of the spectrum by means of a rapidly revolving disc, as in Expt. 63 ii., is very simple. It is due to what is called the **persistence of images** on the retina of the eye. Each impression the retina receives lasts for a certain length of time—about one-tenth of a second. It is not an instantaneous impression only. Think of the common trick of whirling round a stick with a spark on the end which gives rise to the impression of a continuous circle of light. This is because the second impression of the spark is received by the eye before the first impression has died away. Similarly, the impression of one sector, say, a red one, has not disappeared before the next is received, and while these compounded impressions linger a third

one comes along. The blurred total of all these rapidly occurring impressions produces the greyish white tinge seen when a colour disc is whirled.

Colour.—In every case the colour of a body depends on selective absorption or selective transmission. Of the coloured rays of white light one portion is absorbed at the surface of the body. If the unabsorbed portion traverses the body it is **coloured and transparent**: if, on the contrary, it is reflected the body is **coloured and opaque**. In both cases the colour depends upon the constituents of white light which are left to reach the eye after the other constituents have been absorbed. Bodies which reflect or transmit all colours in the proportion in which they exist in the spectrum are white; those which reflect or transmit none are black. Between these extreme limits infinite tints exist depending on the smaller or greater extent to which bodies reflect or transmit some colours and absorb others.

In the case of a blue glass the red and yellow rays of white light are absorbed, while the blue rays are transmitted almost completely, the green and violet less so; hence the light seen through the glass is blue.

Bodies have no colour of their own; the colour of a body changes with the light which falls upon it. It is interesting to remember that this absorption of certain constituents of light necessitates a using up of energy. But since energy cannot be destroyed it is in these cases converted into heat. Theoretically, a blue glass would get hotter than a red one, because the former absorbs all the red rays, and these have a greater heating effect than blue rays.

CHIEF POINTS OF CHAPTER XVI.

Analysis of light by a prism.—The points to be remembered are:

1. That if a beam of monochromatic light, that is, light of one wave length, traverses a prism, it is bent out of its original direction, but the amount of bending produced by any particular prism depends upon the colour of the beam of light used, being greatest for violet and least for red light.

2. That if light from any source passes through a prism, it is broken up or analysed into its different components, each ray of

elementary colour that enters into the composition of the light being bent by a different amount.

3. That a beam of white light, when it traverses a prism is broken up into

Red
Orange
Yellow
Green
Blue
Indigo
Violet

White light can, therefore, like a chemical compound, be broken up into simpler elements.

Recomposition of light—the colour disc.—The analysis of white light is noted above; the synthesis, or recomposition, can be effected:

1. By making the coloured band or spectrum, produced when light has passed through one prism, traverse a second prism having its refracting angle reversed.

2. By rotating a disc of cardboard painted in segments of violet, blue, green, yellow, orange and red.

Colour of transparent bodies.—The colour of transparent bodies is due to the constituents of white light *transmitted* by them.

Colour of opaque bodies.—The colour of opaque bodies is due to the constituents of white light which they *reflect*.

EXERCISES ON CHAPTER XVI.

1. Describe and explain the effects observed when cards coloured bright red, green, and blue respectively, are passed from the red to the blue end of the spectrum.

2. Some glass houses in which ferns are grown are constructed of green glass. Describe the appearance, to an observer in such a house, of a lady in a red costume carrying a book with a bright blue cover. Give reasons for your answer.

3. How would you explain to a class of children the effect of a stained glass window upon sunlight? What simple experiments would you perform to convince them of the truth of your statements?

4. A ray of white light is passed through a glass prism; make a sketch showing how the direction of the ray is changed by its passage through the prism and the order of the colours seen when the light falls on a screen.

How would you show that when these colours are re-combined white light is produced?

5. Describe an arrangement by means of which a spectrum may be formed upon a screen.

If the light is made to fall upon a piece of red glass before reaching the screen, how and why will the spectrum be affected? What would the effect have been if blue glass had been used?

6. How can it be proved that:
(*a*) White light is a mixture of many colours? (*b*) Different colours have different degrees of refrangibility?

7. What is meant by the *dispersion of light*? On what fact does it depend?

8. Explain the term *refrangibility* as applied to a ray of light. Are rays of all colours equally refrangible?

9. It is sometimes said that "red glass colours the sunlight red," and that "blue glass colours the sunlight blue." Mention facts or experiments which show that this is not accurate. Put the statement in a more accurate form.

10. Bright sunlight falls obliquely upon the surface of the water contained in a white china basin; a penny is held near the surface of the water, and in such a position that its shadow falls upon the bottom of the basin. Parts of the shadow are found to be edged with colour. What colours may be observed? On what part of the shadow is each to be seen? How do you account for the colours?

CHAPTER XVII.

TERRESTRIAL MAGNETISM.

64. NATURAL AND ARTIFICIAL MAGNETS.

i. Attractive property of lodestone.—Examine a piece of lodestone. Dip it into iron filings. Observe that the filings adhere in tufts to certain parts of the lodestone.

ii. Directive property of lodestone.—Take a second piece of lodestone which has been roughly shaped so that the places where the filings adhere are situated near its ends. Support the piece of lodestone in a wire stirrup as shown in Fig. 169, and prove that even if at first arranged in any other way, the piece of lodestone, after swinging about for some time, eventually comes to rest along a certain line, and one end of the lodestone, which you can mark with chalk, always points in the same direction.

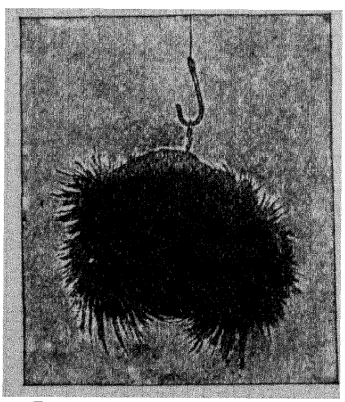


FIG. 169.—A lodestone attracts iron filings, and, if free to turn, sets itself in a definite direction.

iii. Action between two lodestones.—Leave the piece of lodestone of the last exercise suspended in its position of rest. Bring up the first piece of lodestone towards it in such a manner that one of the places where the filings adhered points at the end of the suspended piece. Observe what happens. Now point it at the other end of the suspended lodestone and again observe the result. In one case attraction takes place, while in the other repulsion ensues.

iv. Magnetisation by means of lodestone—Take a good sized sewing needle and fix it on the table with a little soft wax. Using the piece of lodestone from the stirrup in the last experiment, and beginning at the point of the needle, rub the end of the lodestone

along the length of the needle, and when you get to the eye, raise the lodestone and bring it again on to the point of the needle and repeat the stroking process. Do this about ten or twelve times.

v. Properties of a magnet.—Examine the needle which has been rubbed with the lodestone. No change of appearance can be seen, but it will now attract iron filings at its ends. Support it in a tiny stirrup (Fig. 170), and see that it arranges itself as the shaped piece of lodestone did. Also notice that while the point is either attracted or repelled by the end of the shaped lodestone, the eye is repelled or attracted, that is, behaves in an exactly opposite manner.

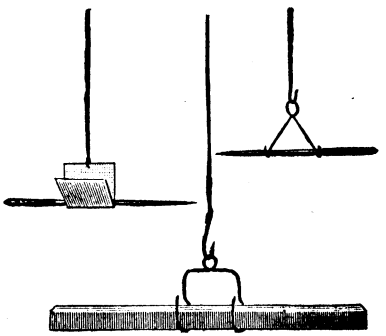


FIG. 170.—Convenient methods of suspending magnets to illustrate their directive properties.

The needle has been made into a magnet, or has become magnetised. Most filings are attracted at its ends, which are in consequence termed the *poles* of the magnet.

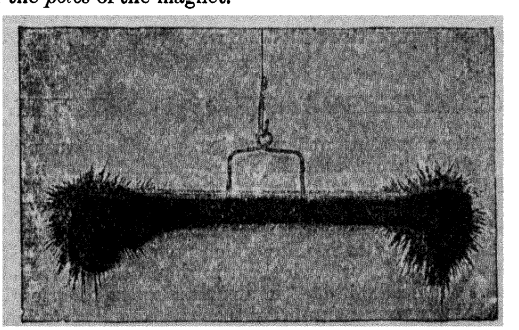


FIG. 171.—Suspended bar magnet with iron filings adhering to its poles.

vi. Artificial magnets.—Examine several forms of artificial magnets. Notice that some are in the shape of bars, others in that of a horseshoe.

Treat the bar magnet in the same way as the shaped lodestone, thus: (*a*) Support in a stirrup, and see that it arranges itself along the same line (Fig. 171).

(*b*) Dip both ends successively into iron filings. Notice and sketch the way in which the filings form a tuft. These ends are the poles. No filings adhere to the centre of the magnet.

(*c*) Support the magnetised needle as in Expt. 64 v. Bring first one end of the bar-magnet, and then the other, up to the point of the needle. Notice and record the result. Repeat, using the eye of the needle.

Lodestone.—Lodestone is a chemical compound of iron and oxygen which occurs in the earth's crust. It is more commonly known as **magnetite**. The name lodestone is reserved for those varieties which have magnetic properties ; it signifies "leading stone" and refers to the early use of this directive property for navigating ships. This mineral is found in considerable quantities in Scandinavia, Asia Minor, United States, and other countries.

Artificial magnets.—The experiments already described teach several important facts. Lodestone is able naturally to attract iron filings. It arranges itself in a particular way when allowed to hang freely. It can impart these properties to pieces of steel, converting them into artificial magnets. These, in their turn, can make other pieces of steel into artificial magnets. All artificial magnets arrange themselves in the same way when freely suspended. **In every respect an artificial magnet has the same magnetic properties as lodestone.**

65. PRIMARY LAWS OF MAGNETISM.

1. **Magnetic attraction and repulsion.**—(*a*) Substituting the bar magnet for the shaped lodestone, magnetise another needle as in Expt. 64 iv.

(*b*) Support the two magnetised needles, which you now have, each in a little stirrup. Allow them to swing freely and come to rest. On the ends of the two needles which point in the same direction stick a piece of paper, or mark them in some other convenient way.

(*c*) Leave one needle in its stirrup and take the other out. Holding the needle in your hand, bring the marked end up to the marked end of the suspended needle. Notice repulsion. Bring the unmarked end of the needle in your hand up to the unmarked end of the suspended needle. Again notice repulsion.

(*d*) Now bring the unmarked end of one against the marked end of the other. Notice attraction.

(*e*) Substitute a soft iron-nail for the needle in your hand in the last experiment. Notice that when either end of the nail is brought up to the marked or unmarked end of the suspended magnetised needle, attraction ensues.

Since unmagnetised iron will attract both poles of a magnetic needle, we are led to an important conclusion, viz. that *repulsion is the only sure test of permanent magnetisation.*

ii. Action between poles of magnet and compass needle.—(*a*) Procure an ordinary compass needle, which is simply a light magnetic needle supported as in Fig. 172, so that it can move easily in a horizontal plane. Notice that one end, which is marked, always points towards the north; it is called the *north-seeking* end. Bring up towards this marked end of the compass needle that end of a bar-magnet which points to the north, when it is freely suspended in a stirrup, and is marked with an *N*. Notice that they *repel* one another. Repeat the experiment using the unmarked, or *south-seeking*, end of the needle and bar-magnet. Observe they similarly *repel* one another.

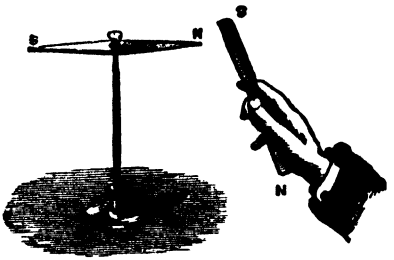


FIG. 172.—Common form of compass needle and stand.

(*b*) Repeat the last experiment, only make the unmarked end of the bar-magnet approach the marked end of the needle. Observe that they rush together or *attract* one another. Similarly, notice that unlike poles in all circumstances attract one another.

iii. Place a bar-magnet upon the table. Arrange a compass needle upon it so that its point of suspension is on the middle line of the magnet, where it is found no filings adhere. Set the compass needle swinging and then allow it to come to rest. Observe that the needle arranges itself, with its marked end pointing towards the unmarked end of the magnet. The reverse is true of the other pole of the needle (Fig. 173).

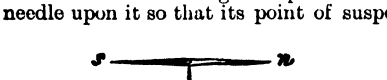


FIG. 173.—Unlike poles attract one another.

We say the magnet exerts a *directive* force upon the needle. Put the needle in other positions on the magnet and notice the same fact.

iv. Result of breaking a magnet.—(*a*) Magnetise a piece of clock-spring. Find which end is repelled by the marked end of a suspended magnetic needle, and stick a piece of paper on this end. Convince yourself that the other end of the piece of hardened spring is attracted by the marked end of the suspended needle. Observe, too, that the middle of the piece of spring has no effect on the needle.

(*b*) Break the piece of spring into halves, and examine each piece by bringing the ends of the pieces in turn up to the suspended needle. What was, before breaking, the middle part of the piece of

spring is now found to affect the needle and to attract the iron filings. Each piece has become a perfect magnet.

(*c*) By means of the suspended magnetised needle, satisfy yourself that the other end of the half which is marked, is attracted by the marked end of the needle, while the other end of the half which is unmarked, repels the marked end of the needle.

Magnetic attraction and repulsion.—Experiments like those now performed teach the rule referred to as the "First Law of Magnetic Attraction and Repulsion," which can be stated thus:

Like magnetic poles repel one another.

Unlike magnetic poles attract one another.

It must, however, be noticed here that, although repulsion is sufficient proof that two like poles are acting upon one another, attraction is not necessarily due to the mutual action of two unlike permanent poles, since, as has been seen in Expt. 65 i. (*e*) magnets can attract soft iron which has no permanent poles at all.

Why a magnetic needle points to the north.—A compass needle always arranges itself with its marked or *N* end towards the magnetic north, because the earth acts like a bar-magnet. A certain part of the earth in the northern hemisphere acts like the unmarked or south-seeking end of the magnet, and attracts the marked end of the needle. The place where this attractive force is greatest is called the **north magnetic pole**, and though at first it seems contradictory, there must evidently be the south-seeking kind of magnetism at the north magnetic pole. The line in which the needle arranges itself is called the **magnetic meridian**. The student must avoid the common error of believing that the south pole of a magnet points towards the north. The **north** pole of a magnet points towards the north; and the question as to the kind of magnetism at the earth's magnetic poles does not affect the names given to the poles of a compass needle or magnet. If it were possible to obtain a magnet with a north pole alone, this magnet would move bodily toward the north pole of the earth; but as every magnet has two poles, the earth's action is only directive.

Lines of force.—When a bar-magnet is covered with a sheet of card or glass, and iron filings are sprinkled on the card, if the card is tapped the filings arrange themselves along definite lines. The filings chiefly collect round the ends, which contain the **poles** of the magnet. The poles are found near the ends of the bar-magnet, where the magnetic power is most strongly

shown. A line joining the poles is called the **magnetic axis.** A line at right angles to the axis and midway between the poles is known as the **neutral line** or **magnetic equator.** It has been seen that there are no filings along this line, where the opposite magnetic properties appear to neutralise one another. Each filing acts like a tiny magnet.

When two magnets are placed near one another, the lines of force due to their mutual action can be shown by means of iron filings. The curves in which the filings arrange themselves indicate the direction of the resultant magnetic force.

66. MAGNETIC DECLINATION.

i. **Magnetic meridian.**—(*a*) Remove all magnets, also any pieces of iron, to a distance. Make circular holes in two pieces of cardboard and fasten fine thread or silk fibres across each as shown in

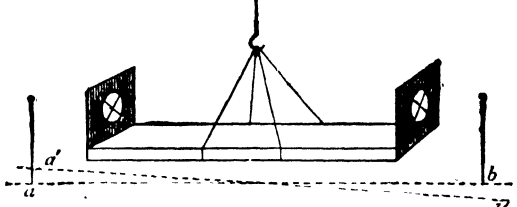


FIG. 174.—Determination of magnetic meridian.

Fig. 174. Fasten these two pieces of cardboard to the ends of a bar-magnet and support the whole as shown in the illustration. Bring the magnet to rest and mark the line of the centres of the cross wires by brass pins stuck into the table. Draw the line *ab*. Reverse the magnet, so that the cross wires are now below it, and repeat. Draw *a'b'*. The line bisecting the angle between *ab* and *a'b'* is the magnetic meridian. Why is it necessary to use pins of brass instead of ordinary pins in this experiment?

(*b*) Freely support above this line each of the magnets you have in order, viz., the shaped lodestone, the magnetised needles, and the horse-shoe magnet. The horse-shoe magnet is best supported vertically by a thread. Notice that on coming to rest they all arrange themselves along the magnetic meridian.

The line along which a freely suspended magnet arranges itself is known as the magnetic meridian, and it can be at once *roughly* traced for any place by the simple experiments you have performed.

ii. **How to find the geographical meridian.**—Allow a freely-suspended compass needle to come to rest. Draw a line on the table

to mark the direction of the needle. Through the point under the spot where the needle is supported draw a line inclined to the

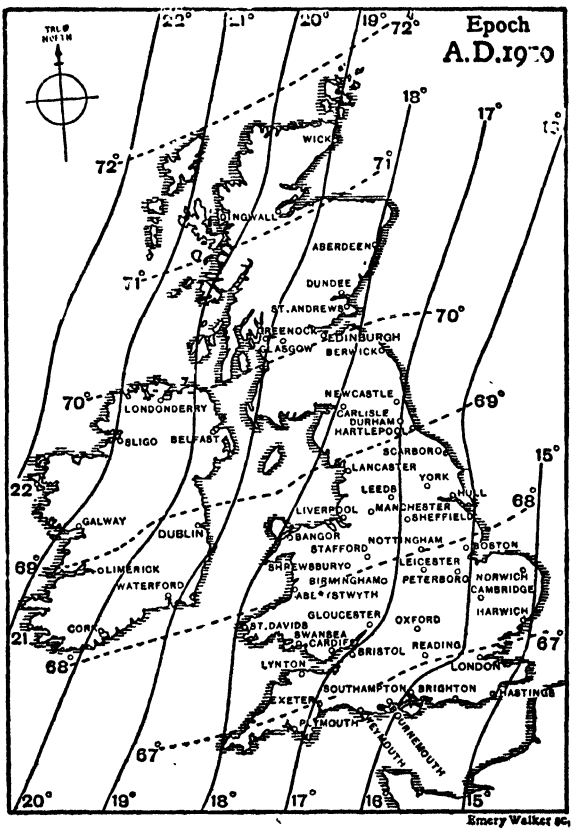


FIG. 175.—The lines running down the figure show where the magnetic declination is the same; the broken lines across the figure connect points of equal magnetic dip. (From Prof. S. P. Thompson's *Lessons in Electricity and Magnetism*. Macmillan.)

magnetic meridian you have just drawn, and at an angle to it equal to the magnetic declination of the place, which can be found from Fig. 175. The lines running from top to bottom are lines of equal

magnetic declination, and the numbers at their ends show the number of degrees by which the north pole of a magnet sets west of true north at the present time in the British Isles.

Declination, or variation.—The magnetic poles of the earth do not coincide with its geographical poles. You will learn later what is known about the position of the former. Great circles round the earth, which pass through the geographical poles, are known as the meridians of longitude. Similarly, curved lines pointing to the magnetic north and south poles of the earth are called **magnetic meridians**; and it is along one of these that a compass needle arranges itself. **The angle between the geographical and magnetic meridians of any place is called the declination or variation of that place** (Fig. 176). Its value for any year at various places is recorded in the *Nautical Almanac*. At Greenwich Observatory in 1915 the declination was 14° 56′ W. Having a compass needle and knowing the angle of declination it is easy to determine the geographical meridian through a place. The direction in which the compass sets itself is found, and then allowance is made east or west of this according to the declination of the place of observation.

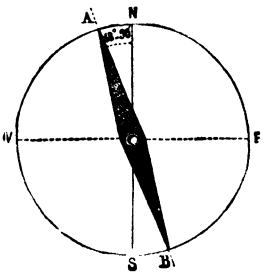


FIG. 176.—Magnetic declination is the angle between the magnetic and geographical meridians of a place.

67. MAGNETIC DIP.

i. Meaning of magnetic dip.—Take a knitting needle and support it by a fibre or two of unspun silk, attached by soft wax, so that it arranges itself horizontally. Now holding it carefully so as not to break the silk, magnetise it by the method of Expt. 65 i. (*a*). Allow it to hang freely again. It is no longer horizontal. One end *dips* down. By means of a compass needle ascertain which pole is dipping, and record your result.

ii. Construction of a dipping needle.—A needle which is free to move in a vertical plane, but fixed as regards movement in a horizontal plane, is called a dipping needle. Either use a bought dipping needle, or make a simple form yourself such as the following:

Select an unmagnetised knitting needle about 6 in. long. Construct an axle for the needle in the following manner: Hold two

short pieces of copper wire on opposite sides of and at right angles to the length of the needle. Twist the ends of the wires together on each side so as to grip the needle tightly, and carefully straighten the twists. Make the wire surfaces as smooth as possible by heating in a gas flame and applying sealing wax, shaking off the excess of wax while still fluid. Apply a spot of sealing-wax so as to connect rigidly the axle to the needle. Make a support for the needle by cutting two rectangular pieces of sheet brass or copper (3 in. × ½ in.), rigidly connect them together at the base with their short edges horizontal and ½ in. apart, and fix them to a suitable base-board. Attach a circular scale of 90° to one of the supports (Fig. 177). See whether the needle is truly balanced by supporting it by its axle on the knife-edges; if necessary, adjust the position of the axle by slightly warming the sealing-wax joint and moving the axle along the needle. Carefully magnetise the needle. Place it on the knife-edges with its axle coinciding with the centre of the circular scale.

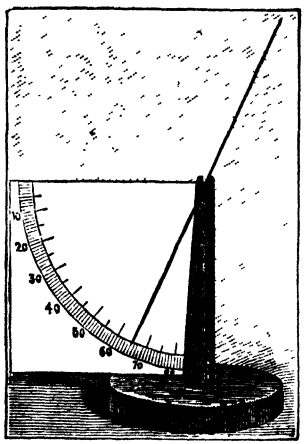
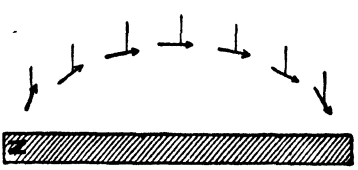


FIG. 177.—A simple form of dipping-needle.

iii. Determination of the angle of dip.—(*a*) To make an accurate measurement of this angle you must be sure of one or two things. *The needle must move in the plane of the magnetic meridian.* One plan to ensure this is as follows: Carefully draw the magnetic meridian by Expt. 66 i., and then arrange the needle so as to lie directly above it. When set free the needle moves in the plane of the meridian.

(*b*) A better way, and the plan generally adopted, is first to rotate the needle until it stands quite vertical, in which case the plane of movement is at right angles to the magnetic meridian. Then turn the plane in which the needle is free to move through exactly 90°, thus making this plane to coincide with the magnetic meridian.


FIG. 178.—Dip of a small compass needle in different positions over a magnet.

iv. Explanation of angle of dip.—Magnetise a sewing needle as in Expt. 65 i. (*a*). Suspend it by a thread so that it hangs quite horizontally. Bring it over the neutral line of a bar-magnet, and notice that it remains horizontal. Gradually move it towards the north-

seeking end of the magnet. Observe that as the pole of the magnet is approached the needle becomes inclined with the south-seeking end downwards at larger and larger angles to the bar-magnet, until when it is over the magnet pole it is vertical. The angle which the needle makes with the bar-magnet corresponds to the angle of dip of a dipping needle.

A dipping needle is simply a magnetic needle suspended in such a way that it is free to move in a vertical plane in a manner similar to that in which the magnetised knitting needle moves in Expt. 67 ii. Fig. 179 will make the principle of its construction quite clear. The value of the magnetic dip at Greenwich in 1915 was 66° 52′.

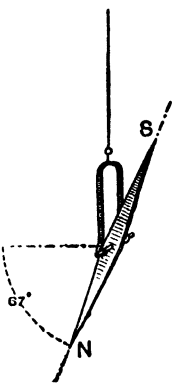


FIG. 179.—Magnetic dip at Greenwich.

Behaviour of the dipping needle at different places on the earth's surface.—It has been seen that a needle assumes a horizontal position when above the neutral line, or magnetic equator, of a magnet. When above the poles of the magnet the needle stands vertical, and in intermediate positions the needle is inclined at a greater and greater angle as the pole is approached. Moreover, over the north-seeking end of the bar magnet the south-seeking pole of the needle is below, whereas over the other end of the bar-magnet the north-seeking pole of the needle is in the lowest position.

Precisely the same thing is observed in the case of the earth; in some places the dipping needle adopts a horizontal position, and a line joining all those stations where this is so marks the **magnetic equator of the earth.** When the needle is moved away from this equator towards one of the magnetic poles of the earth, the dipping needle makes a larger and larger angle with the horizon, or, what is the same thing, the angle of dip increases, until eventually the needle stands vertical, or the angle of dip is a maximum. When this is so the observer knows that one of the magnetic poles of the earth has been reached.

Position of the earth's magnetic poles.—The magnetic poles of the earth, which are located by the vertical position of the dipping needle in their immediate neighbourhoods, do not coincide with the geographical poles. The north magnetic pole,

at which there must be south-seeking magnetism, because the north-seeking pole of the dipping needle is the one which dips, is situated a thousand miles away from the north geographical pole at Boothia Felix in lat. 70° 5′ N., and long. 96° 46′ W. Its position was discovered by Sir James Ross in 1831. The south magnetic pole was found by the Shackleton Antarctic Expedition in 1909 to be a spot in South Victoria Land having the position lat. 72° 25′ S. and long. 154° E.

The earth as a magnet.—The various phenomena, to which

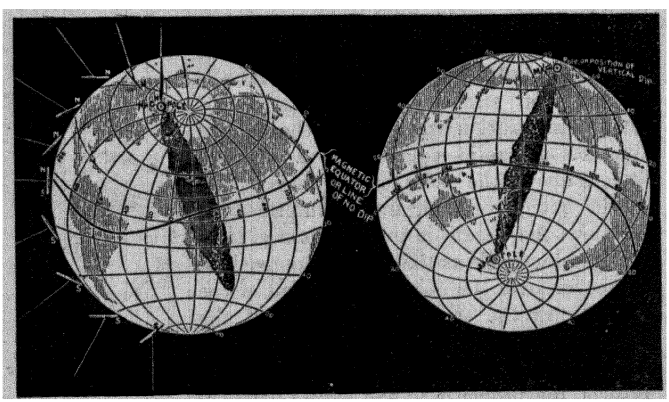


FIG. 180.—Diagrammatic representation of magnetic condition of the earth.

it has been possible only to refer briefly, could be produced in a general way if a huge magnet or lodestone was arranged along a diameter of the earth with its south-seeking end under the north magnetic pole (Fig. 180). The position taken by the dipping needle is that which such a bar-magnet would cause it to adopt. The neutral line of the bar-magnet would be the position of the magnetic equator of the earth, and the magnetic poles of the earth would be over those of the magnet. The earth's magnetic condition can, however, be better represented by considering that there are two magnets within the earth, one much stronger than the other. Of course, it must be understood that no such gigantic magnets are hidden away in

the earth's interior, but the illustration affords a ready means of remembering the chief points which have been enumerated.

The mariner's compass.—Expt. 66 i. (*b*) has shown that every suitably supported magnet arranges itself in the magnetic meridian, and it is on this fact that the mariner's compass depends for its construction. In the actual instrument, a flat needle is suspended by means of an agate cap, placed at its

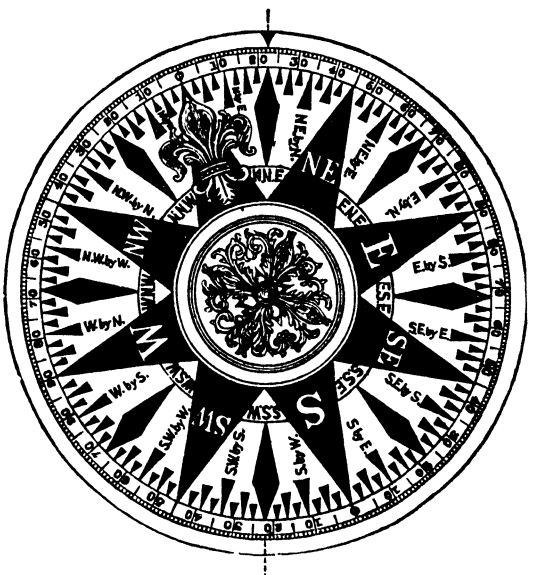


FIG. 181.—A compass card.

centre of gravity. The cap works on a point in such a manner that the needle can move freely in a horizontal plane. On the top of the needle a card, divided as shown in Fig. 181, is fixed, care being taken that the centre of the needle is under the centre of the card, and the north-seeking pole of the magnet under the division marked north on the card. This north point is indicated in the figure by means of the *fleur-de-lis*. With this arrangement the direction of the magnetic north pole is always seen by looking at the *fleur-de-lis*. The dotted line

indicates the direction of the middle line, from bow to stern, of the ship upon which the compass-card in the figure is supposed to be. When the man at the helm wishes the ship to travel in any particular direction, he turns the wheel until the required point of the compass comes under the arrow on the dotted line. In Fig. 181 the compass-card indicates that the ship, the medial line of which is shown, is travelling in a N.N.E. direction.

68. MAGNETIC INDUCTION. METHODS OF MAGNETISATION.

i. Magnetic induction.—(*a*) Place a piece of soft iron in a clamp or on a block on the table, and bring a magnet near one end. It will be found that so long as the iron and magnet are near one another, the iron is magnetic, but directly the magnet is removed, the soft iron loses its magnetism. Test the ends of the iron with a small compass and ascertain what is the nature of the polarity produced in the iron.

(*b*) Repeat this experiment using a very strong magnet and a small piece of steel in place of the soft iron. Similar results will be obtained except that the steel retains its magnetism after the magnet is removed.

ii. Induction by earth.—Hold an iron poker in the line of magnetic dip (p. 251) with its lower end near a compass needle. On tapping the iron a few times with a hammer, it will be found to become magnetic with a north pole at the end near the compass needle.

Magnetic Induction.—The production of magnetism in this way without contact between the *inducing* magnet and the iron or steel is said to be due to Induction.

In Experiment 68 ii. the inducing magnet is represented by the earth, which corresponds to a very weak magnet. Tapping the iron is found to increase the strength of the induced magnetism. The iron poker need not be accurately in the line of magnetic dip; in fact it is a common experience to find workshop tools magnetic after they have been standing vertically for a while in a rack. However, a stronger effect is produced if the iron is actually in the line of dip.

Methods of magnetization.—Artificial magnets may be made from steel bars in several ways.

1. By rubbing with a lodestone as on p. 242.

2. By rubbing with one or more artificial magnets. It is necessary to use the same pole, and to rub the steel in the

same direction the whole of the time. For, supposing the north pole to be used, and the direction of rubbing from left to right, then the new magnet will have a north pole on the left and a south pole on the right. The end last rubbed has a pole opposite in sign to that of the rubbing magnet. Now, a south pole being used in the direction of right to left will produce the same effect as a north pole moved in the opposite direction. Advantage of this fact is sometimes taken, the steel being magnetised by using two magnets at the same time. Commencing in the centre of the steel bar, and using the opposite poles of two magnets, these poles are rubbed on the steel towards the ends, the magnets are lifted, and the strokes repeated from the middle as many times as may be necessary.

3. By passing an electric current round the steel as explained on p. 268. This is the way in which magnets are made at the present time, because not only is it a quicker process, but the steel is magnetised much more strongly than by rubbing with magnets.

CHIEF POINTS OF CHAPTER XVII.

Lodestone is a natural chemical compound of iron and oxygen which possesses the property (1) of attracting iron and steel filings; (2) of coming to rest in a magnetic north and south line when freely suspended or balanced.

Artificial magnets can be made by stroking a piece of steel in one direction with a loadstone or an artificial magnet. In every respect artificial magnets have the same magnetic properties as lodestone.

The primary law of magnetic attraction and repulsion is: like poles repel each other; unlike poles attract each other.

When a magnet is broken, each of the resulting parts is a magnet possessing both a north- and a south-seeking pole.

Magnetic declination or variation is the angle between a true north and south line (shown by a noon-day shadow) and a magnetic north and south line (shown by a compass needle). It differs at different places, and varies from year to year.

Magnetic dip or inclination is the angle which a magnetic needle turning about a horizontal axis makes with the horizon when the vertical plane in which it moves coincides with the magnetic meridian. It varies from place to place and from year to year.

A dipping needle simply consists of a magnet supported on a horizontal axis and free to move in a vertical plane. The needle

must swing exactly in the magnetic meridian before the angle of dip at a place is measured.

The earth's magnetic poles are the points through which magnetic meridians pass, and at which a dipping needle would set vertically. The north magnetic pole is in Boothia Felix, and the south magnetic pole is at about lat. 72° S. and long. 154° E.

Magnetic induction takes place when a magnet is placed near iron or steel. The magnetisation produced in the former substance is temporary, but in the latter it is permanent.

Making magnets.—Steel may be made strongly magnetic by using two magnets and rubbing the steel with opposite poles from its centre towards the ends. The strongest magnets are made by using electric currents.

EXERCISES ON CHAPTER XVII.

1. Suppose you are given a small compass and two knitting needles, of which one is slightly magnetised and the other unmagnetised. How will you find (1) which needle is which, and (2) which is the north-seeking end of the magnetised needle ?

2. A bar-magnet is broken into four pieces. What magnetism do these pieces possess ? How can you prove the truth of your answer ?

3. Two compasses are put near together on a table. In what positions may the needles be expected not to disturb one another ? How will the needles behave when one compass is put to the magnetic north-west of the other ?

4. A magnet is embedded in a piece of wood. How can the position in which it lies be determined without breaking the wood ? How could you determine the magnetic north and south direction by means of such a hidden magnet ?

5. Explain the meaning of the following :

(*a*) The mean declination at Greenwich in 1896 was 16° 56′·5 west.

(*b*) The mean dip in Greenwich in 1896 was 67° 9′. State why it is necessary to specify the place and the year in which the observations were made.

6. Two darning needles are magnetised so that the eye of each is the north-seeking end. They are hung up side by side, each by a separate thread fastened through the eye. What magnetic action between them may be observed ? How do you account for it ?

7. A piece of wood is fastened to a bar magnet in such a way that the latter will float horizontally. The magnet is then placed in a basin of water. What results may be observed ? What do these results teach us about the magnetic force of the earth ?

8. A compass and a magnetised knitting needle are laid upon a table. When one end of the knitting needle is placed close to the compass on the (magnetic) north side, the compass needle settles with its north-seeking end pointing to the (magnetic) north-west. Draw a diagram to show the directions of the various magnetic forces that keep the compass needle in the position in question. You may assume that the further end of the knitting needle is so far off that it does not influence the compass needle.

9. Each of two magnetised knitting needles is suspended by a thread in such a way that it will hang horizontally. In what direction will each needle settle itself when uninfluenced by the other? What magnetic action between the needles may be observed (1) when they are hung side by side, (2) when they are hung with the north-seeking pole of one just under the south-seeking pole of the other?

10. What is a dipping needle? For what purpose would you employ it? How could you satisfy yourself that the dip was not due to the gravitational action of the earth?

11. Describe the various methods of making magnets. Explain some simple experiments to illustrate magnetic induction.

CHAPTER XVIII.

STATIC ELECTRICITY.

69. ELECTRIFICATION.

i. Development of electrification by friction.—(*a*) Arrange a variety of small light fragments, such as pieces of paper, bran, saw-dust, in a heap upon the table. Vigorously rub a rod of glass with a piece of dry silk and hold it over the light fragments. Observe how they are attracted by the rod.

(*b*) Repeat the last experiment, using

(1) Rod of sealing wax and flannel.

(2) Rod of ebonite and a cat's skin.

(3) Sheet of brown paper and a clothes brush.

To obtain satisfactory results the rods and rubbers should be quite dry and warm. This can be ensured by placing them on sand contained in a baker's tray supported on a couple of tripod stands, or in some other suitable manner, and warmed by a Bunsen burner or two. The tray should be covered with a piece of sheet iron bent into the form of an arch.

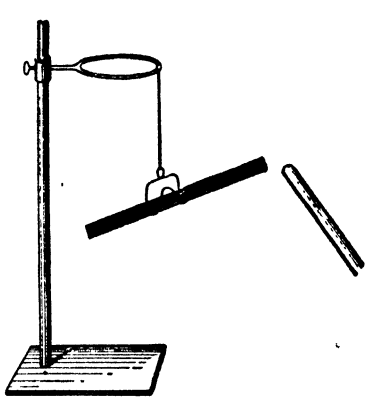


FIG. 182.—To illustrate Experiment 69 ii. (*a*).

ii. Electrical attraction and repulsion.—(*a*) Make a stirrup out of stout copper-wire. Hang it from the ring of a retort stand by means of a thread or ribbon. Balance a round ruler in the stirrup. As in Expts. 69 i. (*a*) and (*b*), electrify a rod by rubbing and bring it near to the balanced ruler. Notice attraction.

(*b*) Substitute other heavy rods for the ruler and observe that all are attracted by the electrified body. Vary the experiment by balancing the rubbed, and consequently electrified, body on the stirrup, and holding the rods, which were previously balanced, in your hand, bring them up to the rod on the stirrup. Again attraction takes place.

(*c*) Repeat Expt. 69 i. (*a*) and observe that after the light particles have been attracted by, and touch, the electrified rod, they are immediately repelled.

(*d*) Or, support two pith-balls by two cotton threads to a wire stuck in a rod of sealing wax supported as in Fig. 183. Bring the electrified rod near the balls. They are attracted by, and touch, the rod. This is followed by repulsion, and afterwards it will be observed that the balls repel one another and stand apart.

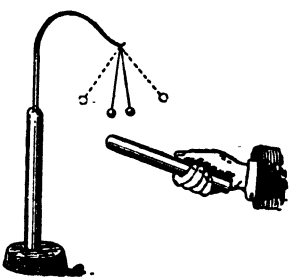


FIG. 183.—Pith-balls supported as described in Experiment 69 ii. (*d*).

iii. Two kinds of electrification.—(*a*) Rub a piece of glass tube with the piece of dry silk, and support it on the hanging stirrup. Then rub a piece of sealing-wax with flannel and bring the rod of sealing wax near the glass tube. Notice *attraction*. Record the fact that glass rubbed with silk is attracted by sealing wax rubbed with flannel.

Repeat the experiment, first rubbing and supporting the sealing wax, and then bringing near it the rubbed glass tube. Notice the same result.

(*b*) Support one piece of glass tube which has been rubbed with the silk rubber (in this case it is advantageous to hang the stirrup by a silk thread) and bring up a second glass tube which has been similarly treated. Notice *repulsion*. Record the fact that glass rubbed with silk is repelled by other glass similarly treated.

Repeat the last experiment, using two sticks of sealing wax and a flannel rubber. Record the result.

(*c*) Support a single pith-ball by a silk thread to a support with a varnished glass leg. Bring in contact with the ball a glass rod which has been rubbed with silk.

(*d*) Using two *pith-ball pendulums*, as the arrangement in the last exercise is called, one charged as in the last experiment, and the other similarly, but by means of a piece of sealing wax rubbed with flannel, test the nature of the electrification generated on the rubbed body in each of the following cases: (i.) sulphur rubbed with fur; (ii.) sulphur with flannel; (iii.) ebonite with silk; (iv.) ebonite with fur; (v.) glass with flannel; (vi.) amber with flannel.

Electrification.—That some substances when rubbed acquire the peculiar power of attracting light bodies has been known from very early times. It was recorded by Thales (B.C. 600)

that **amber** when rubbed became different from all other substances by possessing this power of attraction. The supposed cause of this property is called **electricity**, from the Greek word for amber (electron). Up till the end of the sixteenth century amber was considered to be unique. When, however, men began to observe and experiment, it was soon found that many substances behaved in a similar way. Substances which, like amber, became electrified by rubbing were called **electrics**.

It is now known that, under suitable conditions, most, if not all, substances can be **electrified** by rubbing with a suitable substance. A rod of sealing-wax rubbed with flannel will cause quite a flutter among fragments of tissue paper, and a sheet of warm brown paper, after stroking with a clothes brush, also will attract light bodies.

In order that these electrical effects may be well seen, it is essential that the materials used be quite dry. The best way to ensure this is to place them in front of a clear fire or in an oven such as that sufficiently described in Expt. 69 i. (*b*).

Electrical attraction and repulsion.—To obtain a fuller knowledge of electrification, it is necessary to observe more closely and to experiment with precision—that is, scientifically—on electrified bodies. Are all light bodies attracted, or only a selected few? It is easy to detect even a small attraction when a body is suspended in an apparatus similar to Fig. 183, and some such form may be used conveniently. Small balls of various substances are easily attached, by means of thread, to a support having a varnished glass leg. Whatever the composition of the balls, they are attracted without distinction by an electrified rod. If the electrified body itself is suspended, any substance brought near it will attract it.

The attraction between a body that is electrified and one that is not is mutual; each attracts the other. When one of the suspended balls is allowed to touch the attracting rod, after a momentary hesitation, it flies away and afterwards refuses to approach the rod again. If two balls hanging together have been attracted and touched by the rod, not only are they repelled by the rod, but they also repel one another (Fig. 183).

Two kinds of electrification.—If a suspended ball is **repelled** by a glass rod which has been rubbed with silk, a

piece of sealing wax rubbed with flannel will **attract** it. Thus, the sealing wax and glass, though both electrified, behave differently. It is found that all electrified bodies behave either like the rubbed glass or like the rubbed sealing wax. Electrification is, then, of two kinds. It has been seen that when an electrified body shares its electrification with another body, the former immediately repels the latter; also that a glass rod rubbed with silk will repel another similarly treated. From such facts as these it may be concluded that **bodies similarly electrified repel one another.** Sealing wax rubbed with flannel will, however, attract a glass rod previously rubbed with silk. It may be stated, therefore, that **bodies oppositely electrified attract one another**; but attraction must not be considered as a criterion of opposite electrification, for an electrified glass rod will not only attract an electrified rod of sealing wax or ebonite, but any electrified rod will attract an unelectrified body.

The names **vitreous** and **resinous** were given originally to the states of electrification of the glass and sealing wax respectively. When, however, it was found that glass rubbed with fur possessed the same kind of electrification as sealing wax rubbed with flannel these names fell into disuse. Instead of vitreous and resinous, the words **positive** and **negative** are now used. Glass rubbed with silk becomes **positively electrified** (+ly), while sealing wax rubbed with flannel is **negatively electrified** (− ly). When a body is electrified it is also said to be **charged.**

70. ELECTRIC CHARGES.

1. Equal and opposite charges.—(*a*) Make a flannel cap just to fit the end of a stout rod of sealing wax or other convenient substance. Attach a silk thread to the flannel cap. See that both rod and cap are dry and warm. Have a positively charged pith-ball pendulum near. Rotate the flannel cap once or twice by pulling the piece of silk which has been previously wound round it.

(*b*) Holding the cap by the silk thread, bring it near the +ly charged ball. Notice repulsion. The cap is therefore +ly charged.

(*c*) Touch the pith ball with a finger, and then charge it negatively by means of a piece of sealing wax rubbed with flannel. Bring near it the end of the rod on which the flannel cap has been rubbed. Again notice repulsion. The rod was therefore −ly charged by the flannel cap.

(*d*) Put on the cap again and repeat the rubbing. Do not take off the cap, but bring both up to an uncharged pith ball. There is neither attraction nor repulsion.

ii. **Conductors and insulators.**—(*a*) Rub with a piece of dry silk a brass tube held in the hand. Bring the tube near to the cap of an electroscope. There is no divergence of the leaves.

Procure a piece of brass to which a varnished glass handle has been attached, and holding the rod by the glass handle, flick the brass with the silk, or with a cat's skin. Now, quickly bring the brass in contact with the cap of the electroscope. Notice the divergence of the leaves.

Try to think what causes the difference between this result and the former one.

(*b*) Cause the leaves of an electroscope to diverge widely by means of small positive charges from a proof plane, such as a metal disc attached to a glass rod. Touch the disc of the electroscope in succession with pieces of glass, sealing wax, solid paraffin, ebonite, and metal rod.

Charge the electroscope again and touch the disc with the finger. Record your results.

During electrification equal and opposite charges are produced.—So far only the thing rubbed has been considered; the glass has been observed and the silk neglected; the sealing wax attended to and the flannel disregarded. But careful experiment shows that not only is the rod electrified but the rubber also. The electrification of one is of the opposite kind to that of the other. While the sealing wax becomes negatively electrified the flannel is positively charged. If a body is rubbed and the rubber left in position, no trace of electrification can be observed. The charges are produced in quantities which balance one another. When the bodies are separated, however, each shows its appropriate kind of electrification.

Electroscopes.—An electroscope is an instrument for detecting the presence of small quantities of electricity. The instrument can also be used to determine the kind of electrification. A suitably suspended ball of pith on a varnished glass leg, or on a stick of sealing wax, serves this purpose. Electrified bodies attract the ball. When, however, the ball is itself charged, by touching an electrified body, it enables the observer to discriminate further.

Bodies electrified in a similar way to the ball repel it; all other bodies attract it. It must be noticed carefully that attraction does not tell whether a body is electrified or

not. The attracting body may either be charged with the opposite kind of electricity to the pith ball, or it may be uncharged. With another pith ball charged oppositely to the body brought near to it, **repulsion** would be **a sure test of electrification.**

The **gold leaf electroscope** (Fig. 184) is much more convenient in use than the pith ball pattern. Two forms of gold leaf electroscope are shown in Figs. 184 and 185. In Fig. 184 the metal wire, carrying the gold leaves at one end and a metal disc at the other, is supported inside a glass tumbler in the manner the illustration makes clear. In Fig. 185 a bottle takes the place of the tumbler and the metal wire passes through an india-rubber stopper in the neck of the bottle. The approach of an electrified body is indicated by the divergence of the leaves of the electroscope. This form of electroscope can be charged by carrying to it small quantities of electricity on a **proof plane,** that is, on a small metal disc fixed to a handle of varnished glass or sealing wax.

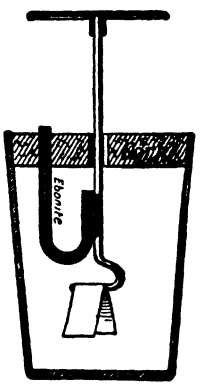


FIG. 184.

Conductors and insulators.—The reasons for many precautions, which it has been necessary to take, have not as yet been given. The varnished glass stand of the pith-ball electroscope, the ebonite support of the metal carrying the gold leaves of the gold leaf electroscope, and the varnished glass handle of the proof plane, are all provided for a definite reason. When the cap of a charged electroscope is touched by the hand or a metal rod, its charge disappears. Recharged and touched by a rod of glass, ebonite, or sealing wax, it is undisturbed; its charge remains.

The metal rod and the hand **conduct** the electric charge away. Along the glass, ebonite, or sealing wax, the electric charge cannot escape. Substances which allow electricity to pass along them are called **conductors**; ebonite and similar substances which prevent its flow, are called **insulators.** Consequently, **to preserve an electric charge it must be separated from the earth by an insulator.**

71. ELECTRIC INDUCTION AND STORAGE.

1. **Induction.**—Cause a charged rod to approach towards the disc of an electroscope (Fig. 185). Notice the divergence of the leaves. Still holding the rod in the same position, touch the cap of the electroscope. The leaves completely collapse. Remove, first the finger from the disc of the electroscope, and second, the rod from the neighbourhood of the electroscope. The leaves again diverge and remain divergent. Test the electrification of the leaves and convince yourself that the charge is of the opposite kind to that on the rod.

Draw a series of diagrams showing the condition of electrification of the rod and parts of the electroscope during each of the stages of this experiment.

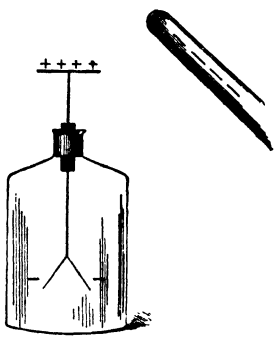


FIG. 185.—Charging a gold leaf electroscope by induction.

Electric Induction.—When a body, charged electrically, is brought near an insulated cylinder, which is connected by a wire to an electroscope, the leaves diverge. On removing the wire with an insulator the leaves remain divergent, that is, they have received an electric charge. If the charged body is now taken away the insulated cylinder, as well as the electroscope, is found to be charged. The charge on the cylinder is of the opposite kind to that originally used, while that of the electroscope is of the same kind. The mere nearness of the rod has caused a separation in the cylinder of positive and negative electricity. This influence is called **electric induction.**

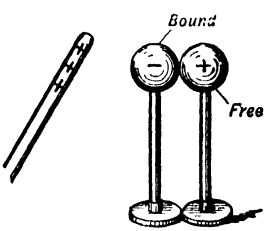


FIG. 186.—Bound and free electric charges.

If two insulated metal balls be placed in contact and a positive charge brought near them (Fig. 186), each will, if the balls are separated while under the influence of the positive charge, be found to be electrified, the nearer ball with negative

electricity, and the more remote positively. Brought into contact again all electrification of the balls disappears.

The charges are not only opposite in kind but equal in amount. The action of the original charge, called the **inducing charge,** may be thought of as being the separation of the electricities, attracting near to it and holding fast the opposite kind while repelling the like kind of electricity.

The charging of a gold leaf electroscope by induction affords a good example of what happens. When a negatively charged rod is brought near the disc (Fig. 185) induction takes place, positive electricity is drawn to the disc, while negative electricity is repelled to the leaves. The leaves, thus similarly charged, repel one another (Fig. 185). On now touching the disc with the hand all signs of electrification disappear and the leaves collapse. If now the hand is taken away and the charged rod removed, then the leaves once more diverge, this time with a positive charge. The repelled negative electricity was alone **free** to escape, the positive charge being held tightly **bound** by the attraction of the negative charge of the rod. On removing the rod the positive charge became free also, spread to the leaves and caused them to repel each other. The electroscope is now charged by induction. The two kinds of electricity separated by an inducing charge are spoken of as **bound** and **free,** from their behaviour while under the influence of the charge.

CHIEF POINTS OF CHAPTER XVIII.

Production of electrification.—Many substances when rubbed with suitable material are able to attract light bodies—they become electrified.

Electrification is of two kinds, vitreous and resinous, or better, positive and negative. These two kinds are always produced together in equal quantity.

Attraction and repulsion.—Bodies charged with the same kind of electricity *repel* one another, those charged with opposite kinds, *attract* one another.

Induction.—When a charged body is brought near an insulated conductor, the conductor is found to be charged; the near side of the conductor having an opposite charge to that of the body, while the far side is similarly charged. The similar repelled charge is said to be *free,* while the opposite attracted charge is *bound.*

EXERCISES ON CHAPTER XVIII.

1. How would you prove that an electrified body is attracted by an unelectrified body ?

2. How would you prove to a class that there are two kinds of electricity ?

3. How would you prove that glass and silk when rubbed together are equally and oppositely charged ?

4. Having given a gold leaf electroscope and a piece of ebonite and cat's skin, what experiments would you make to determine whether the electrical charge on a given charged insulated body was positive or negative ?

5. How would you show that a brass rod is capable of being electrified ? Explain why on rubbing a brass rod and a glass rod the latter only ordinarily appears to be electrified by the friction.

6. The caps of two gold-leaf electroscopes *A* and *B* are connected by a long wire, and a positively charged sphere is brought near *A*. What will be the indications of the electroscopes, and how will they alter if either *A* or *B* is touched ?

7. Explain what is meant by electrostatic induction.

Two small light pith balls are in contact, and are supported by separate threads. A charged glass rod is brought in the neighbourhood of the balls. What will happen (*a*) when the threads are wet and conducting, (*b*) when they are dry and insulating ?

CHAPTER XIX.

VOLTAIC ELECTRICITY.

72. THE ELECTRIC CURRENT.

i. Fundamental experiments.—(*a*) Prepare some dilute **sulphuric acid** containing one part of strong acid to eight parts of water. First measure out the water into a large beaker, and then gently pour the measured quantity of strong acid into the water, keeping the latter briskly stirred with a glass stirring rod. Notice the large amount of heat generated and set the mixture on one side to cool.

(*b*) Plunge a strip of commercial zinc into a beaker of cold dilute sulphuric acid prepared as in the last experiment. Notice the brisk evolution of gas which takes place.

(*c*) Repeat the last exercise substituting first a rod of pure zinc and then a strip of copper. Observe that there is no chemical action in either case.

(*d*) Place the rod of pure zinc and the strip of copper into the dilute acid, taking care that the two metals do not touch one another. No gas is given off from either metal.

Now tilt the pieces of metal towards one another until they touch outside the liquid. Observe that bubbles of gas appear on the copper plate.

ii. Amalgamated zinc.—Prepare a plate of amalgamated zinc by dipping a plate of ordinary commercial zinc into dilute sulphuric acid, and, after it has been acted upon for a minute or two, rubbing some mercury completely over its surface with a piece of cloth. Repeat Expt. 72 i. (*c*), and observe that amalgamated zinc behaves just like pure zinc.

iii. Magnetic action of electric current.—(*a*) Into some dilute sulphuric acid contained in a beaker plunge a plate of amalgamated zinc and one of copper, to each of which a copper wire is attached by a suitable binding screw. Procure an ordinary compass needle, and bring this up and try to arrange matters so that the wire in connection with the copper and zinc plates is parallel to the magnet and in the same vertical plane. Notice the deflection of the magnet.

(*b*) Wind a cotton-covered wire which is connected with the copper and zinc plates round a piece of galvanized iron, as in Fig. 187. Notice that the piece of iron will attract iron filings.

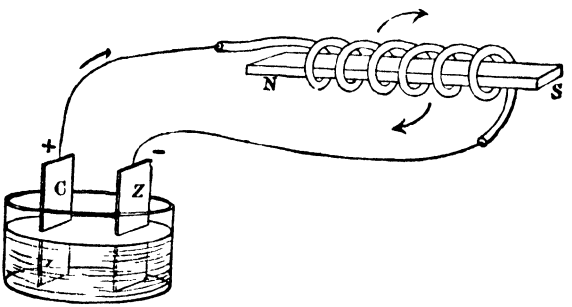


FIG. 187.—To illustrate Exercise 72 iii. (*b*).

iv. Polarisation.—Repeat Expt. 72 iii. (*a*), and notice that after a time the force which the wire exerts over the magnetic needle becomes feebler. Rub the copper plate with a piece of wood until all the bubbles of gas which had collected have disappeared, and notice that the power of the wire to deflect the magnetic needle is regained.

The simple cell.—When a piece of commercial zinc is placed in dilute sulphuric acid, bubbles of gas escape from the liquid. This is an example of **chemical action.** The zinc ceases to exist as zinc, while a new substance, the gas hydrogen, makes its appearance (Chapter XXV.). A rod of copper, or rods of pure, or amalgamated, zinc are unaffected by weak sulphuric acid ; consequently, on placing copper and zinc rods in the acid, without bringing them into contact, no effect is observed ; when the metals are made to touch—either inside or outside the liquid—bubbles of the gas hydrogen are rapidly given off from the copper plate. This connection of the metals, therefore, seems to be an essential condition of activity in the cell. It is not necessary actually to make the metal plates touch. If they are connected by means of a wire outside the liquid the same effect takes place.

If now a small magnetic needle is brought near this wire, the wire is found to have acquired a new power. The needle is disturbed as though in the presence of another magnet. Similarly,

if the wire is wound round a piece of soft iron, such as a strip of galvanised iron, and its ends kept in contact with the metals, the soft iron becomes a magnet under the influence of the wire connecting the copper and zinc plates. The wire is now said to have a current of electricity, or an **electric current** passing along it. The electric current flows, outside the liquid, from the copper to the zinc. The part of the copper plate outside the liquid, and with which the zinc plate is connected, is referred to as the **positive pole**, and the part of the zinc plate, with which connection is made, is termed the **negative pole**. This apparatus for producing an electric current is known as a **simple voltaic cell.**

Polarisation.—The electric current in the wire is neither permanent nor constant; it wanes and finally ceases. Simultaneously, the action in the liquid is found to have ceased, and on examination the copper is found to be covered with adhering bubbles of gas. On brushing off these bubbles, the chemical action in the cell recommences, and an electric current again flows through the wire, which regains its power of influencing a neighbouring magnet. The accumulation of gas on the copper plate appears to choke the cell. This effect is called **polarisation.** A cell so impeded by the products of its own activity is said to be **polarised.**

The simple voltaic cell, on account of the defect of polarisation, is not of practical importance. The cells in common use get rid of the obstructing gas either by mechanical or chemical means. In cells depending on the first method the negative plate is roughened, thus facilitating the escape of the gas. Those in which chemical means are used are of many types.

73. TYPES OF VOLTAIC CELLS.

i. **Daniell's cell.**—Examine the parts of a Daniell cell. Connect covered copper wires to the binding screws—one connected with the outer copper vessel, the other with the rod of zinc. To charge the cell proceed as follows: Fill the inner vessel with dilute sulphuric acid, and then three parts fill the outer vessel with copper sulphate solution.

ii. **Bunsen's cell.**—Examine a Bunsen cell. Fit up and charge it, and, in the same way as before, satisfy yourself that an electric current is circulating. Also notice that a small spark occurs when

the ends of the two wires from the carbon and zinc poles are brought together and separated suddenly.

Daniell's cell.—In most of the cells in which polarisation is prevented by chemical means, there are two vessels one placed inside the other. The inner one is made of porous earthenware which permits a slow passage through it of the liquids on either side of it. In Daniell's cell the outer vessel is of copper and serves as the copper plate. This outer vessel contains a solution of copper sulphate (blue vitriol), the strength of which is maintained by placing some crystals of the same substance on a tray, which extends round the top of the inside of the copper vessel (Fig. 188). The inner porous pot contains

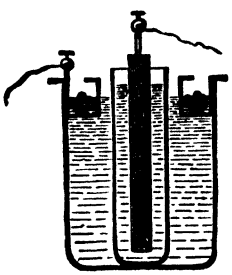


FIG. 188.—A Daniell cell.

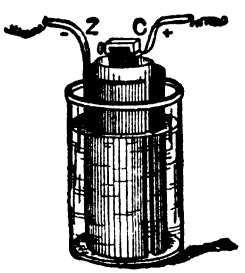


FIG. 189.—A Bunsen cell.

dilute sulphuric acid into which dips a rod of amalgamated zinc. In this cell the hydrogen, which would otherwise be deposited on the copper vessel, as in Expt. 72 i. (*d*), acts on the copper sulphate and converts it into sulphuric acid, while the copper from the blue vitriol is harmlessly deposited on the copper vessel.

Bunsen's and Grove's cells.—The only difference between these two kinds of voltaic cells is that, whereas, in the former a piece of hard carbon replaces the copper plate, in the latter there is a plate of platinum. Owing to the cheapness of the carbon, Bunsen's cell is the more commonly used.

In Bunsen's cell there are two separate vessels; the inner smaller one alone is porous and is filled with strong nitric acid, into which the piece of carbon dips. The outer vessel contains dilute sulphuric acid, and in it is placed a zinc

plate, which is usually made cylindrical in shape. The arrangement of the parts is easily understood by a reference to Fig. 189.

The hydrogen-destroying agent in both these cells is nitric acid. As soon as hydrogen is formed, instead of adhering to the platinum or carbon plate, it reacts with the nitric acid, giving rise to poisonous red fumes which escape into the air.

74. MAGNETIC ACTION OF ELECTRIC CURRENTS.

i. Magnetic field due to a current.—(*a*) Arrange the wire, which in Expt. 72 iii. (*a*) connects the plates of the voltaic cell, in a vertical position. Hold a compass needle at the side of the wire and then proceed to move it slowly round the wire. Change the connections of the wire with the battery and try again. Record your observations. The needle always tends to set itself at right angles to the line joining its centre to the nearest part of the wire.

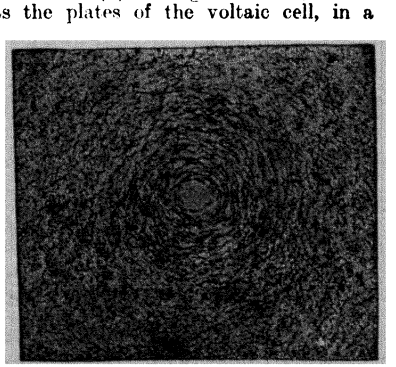


FIG. 190.—Map of the magnetic field perpendicular to a wire conveying a current.

(*b*) If a very strong current—say that from six quart Bunsen cells—is available, perform the following experiment: Before connecting the vertical wire in the last experiment with the poles of the battery, thread on it a fairly large piece of cardboard through which a hole has been punched. Connect up the battery. Fix the card horizontally by means of a wooden universal joint, and sprinkle iron filings upon it. Gently tap the card and observe and sketch the arrangement of the filings in the neighbourhood of the wire (Fig. 190).

ii. An electro-magnet.—Bind covered copper wire around a horse-shoe, or a ∩-shaped piece of soft iron, so as to make an electro-magnet having opposite poles at the two ends (Fig. 192).

Magnetic field due to a current.—Magnets placed in the neighbourhood of an electric current are influenced by it. Electric currents are surrounded by magnetic fields. The strength of such a magnetic field is found to depend on the strength of the electric current, and the direction of the lines of force to depend on the direction of the electric current. If a

small pivoted magnet is moved in the vicinity of a vertical wire along which an electric current is flowing, it will be noticed that the needle always tends to set itself at right angles to the line joining its centre to the nearest part of the wire. It has been shown already that a small magnet in a magnetic field always takes up a definite position along a line of force.

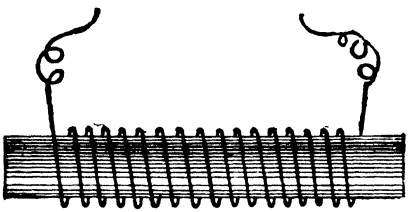


FIG. 191.—An electro-magnet.

Expt. 74 i (*b*) shows that around a wire carrying a current the lines of force are concentric circles with the wires as the centre of the circles. By passing the vertical wire through a horizontal card on which are sprinkled iron filings, the structure of the field of force is clearly seen when the card is tapped (Fig. 190). A north pole introduced into such a magnetic field would circle continually round the wire carrying the electric current.

Electro-magnets.—A coil of wire conveying a current can act like a magnet. On placing an iron core inside the coil, its magnetic strength is greatly increased. For the time being the iron becomes a magnet, and the magnetic power of the combination exceeds very much that of the current alone. Such a combination is called an **electro-magnet** (Fig. 191). When the iron core is bent into the form of a horse-shoe, we have the horse-shoe electro-magnet (Fig. 192). In winding this form of magnet, care must be

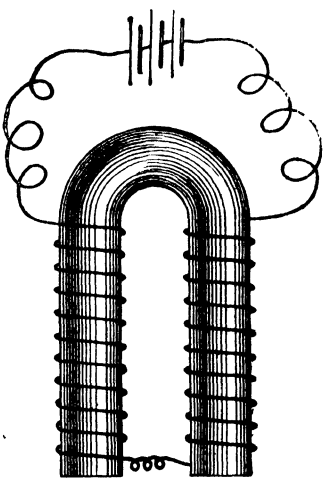


FIG. 192.—A horse-shoe electro-magnet.

taken that the ends of the horse-shoe are of opposite polarity. For lifting purposes the horse-shoe form of magnet is very effective.

By employing very soft iron, wrapping many turns of wire round the iron, and passing a strong current through the coil of wire, an electro-magnet of enormous power may be constructed.

75. GALVANOMETERS.

1. Direction in which a magnetic needle is deflected by the electric current.—(*a*) Using whichever kind of cell is most convenient, proceed to study the action of an electric current upon a compass needle placed in the magnetic meridian. Stretch a piece of covered copper wire, about a yard long, between two universal joints, and arrange it in the magnetic meridian. Name one end *A* the other *B*. To each end of this wire attach binding screws of the pattern shown in Fig. 193. Put an ordinary compass needle *under* the wire and let the needle come to rest; it will of course, in the circumstances, be parallel to the wire. Connect the wires from a single voltaic cell with the binding screws attached to the wire *A B*. Notice and record the direction in which the marked end of the magnetic needle is deflected. Disconnect the wires from the battery, and reverse the connection with the wire *A B*. Observe that the marked end of the needle is now deflected in the opposite direction.



FIG. 193.—A binding screw.

(*b*) Repeat the exercise, this time arranging the magnetic needle *above* the support. Notice and record the way in which the needle is deflected. Connect the wires in the second way described in the last experiment, and again observe and record the direction in which the needle is deflected.

Make a table of the results.

DIRECTION IN WHICH CURRENT FLOWS ALONG WIRE *AB*.	POSITION OF NEEDLE.	DIRECTION OF DEFLECTION OF MARKED END OF NEEDLE VIEWED FROM ABOVE.
Example— *From A to B*	*below*	*to left*

(*c*) Place the compass needle in the magnetic meridian as before, and connect the terminals of the voltaic cell with the copper wire *A B*. Observe and record the direction in which the north end of the

needle moves for the four following cases with the wire *A B* **vertical**: (i.) near north end and current running down, (ii.) near north end and current running up, (iii.) near south end and current running down, (iv.) near south end and current running up. Remember that the current flows from the carbon or copper outside the cell.

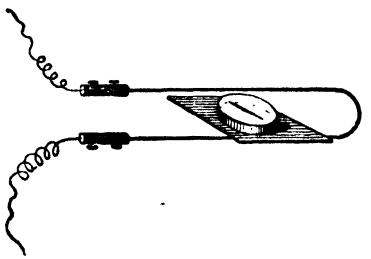


FIG. 194.—Principle of the galvanometer.

ii. Principle of the galvanometer.—Support the compass needle on a piece of cardboard, held horizontally in a clamp; now bend the wire *A B* so that the needle can be arranged in the loop of wire formed (Fig. 194). Arrange the loop of wire and the needle in the magnetic meridian, pass the electric current, and notice the amount of deflection of the needle.

(*e*) Now coil the wire so that there are two lengths above and below the needle, and repeat the previous experiment. The deflection of the needle is greater than before. This experiment illustrates the principle of construction of the *galvanometer*.

Ampère's rule.—It is often desirable to know in what direction to expect the deflection of a magnet which is being influenced by an electric current passing along a wire; or, from the magnet's movement to determine the direction of flow of the electric current. By placing a wire carrying an electric current in various positions near a suspended compass needle, a simple relation is discovered between the direction of movement of the marked pole and the direction of the current. The statement of it is known as **Ampère's rule**. One of the most convenient ways of expressing the rule is to say: If the fingers of the right hand are made to point along the direction of the current when the palm of the hand faces the magnet pole, the outstretched thumb will indicate the direction in which the marked or north pole tends to move (Fig. 195).

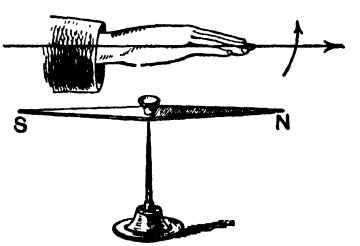


FIG. 195.—Ampere's rule.

Galvanometers.—To detect an electric current, use is made of the fact that it can deflect a magnetic needle in its immediate neighbourhood. When the direction of the deflection can be observed easily, and the amount of such deflection bears a definite relation to the strength of the electric current causing it, the instrument is called a **galvanometer.** Fig. 196 represents a simple type of such instrument. On applying Ampère's rule to each part of each turn of the wire it is found that in a galvanometer the tendency is everywhere to move the magnet pole in the same direction. The effects, too, are added together in the instrument so that even a weak current may make itself apparent. With the same current, the deflection shown by the galvanometer is greater as the number of turns of wire is

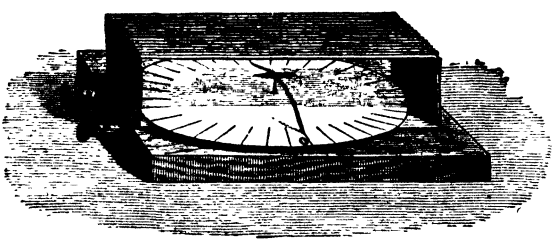


FIG. 196.—A simple galvanometer. The short compass needle has a light pointer fixed at right angles to it.

increased. To use the instrument it is so arranged that the electric current, the direction and strength of which it is desired to measure, passes through the coil of wire round the needle. The current enters by one terminal, flows round the coil, and leaves by the other terminal. The coil itself is always kept in the magnetic meridian parallel to the direction of the undisturbed needle. On the passage of the current the needle is turned out of its normal position. From the magnitude of the angle of deflection of the needle the strength of the current can be calculated, if certain facts about the instrument are known.

If the magnetic force tending to hold the needle in the magnetic meridian is increased, a stronger electric current will be required to produce the same deflection of the needle as before. The galvanometer, in other words, will have become less sensitive. This state of affairs is sometimes required. The earth's

magnetic field can be aided, in such circumstances, by placing a bar-magnet in the magnetic meridian, with its marked or North pole pointing North, so that the South end attracts the North pole of the galvanometer needle. The galvanometer needle under the influence of this combined field is less easily deflected, and is suited now to respond to currents of increased strength.

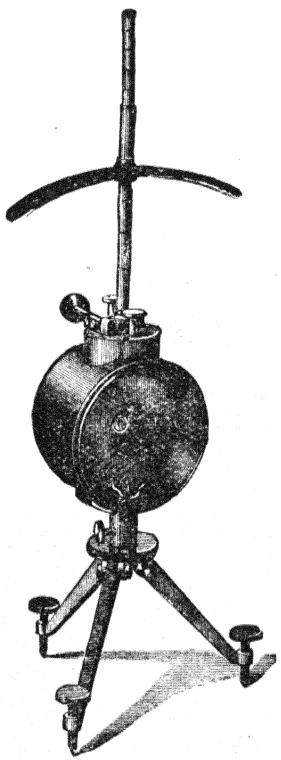


FIG. 197.—A mirror galvanometer.

If the galvanometer coil is used in an East and West position the passage of a current will probably leave the magnetic needle undisturbed. The newly created magnetic field due to the current will, in this case, either be in the same direction as that of the earth, or exactly opposed to it. If the first state of affairs holds good, the magnet is simply confirmed in its normal position, while if the second set of conditions obtain, the needle may swing round completely, so that its marked end now points south.

The mirror galvanometer.—When an instrument is required to detect very weak currents or currents of short duration, a mirror galvanometer is employed. The principle of this instrument is the same as that of the simple galvanometer. It owes its great sensitiveness to a number of different causes. The small magnet or magnets, are attached to a tiny mirror, and are delicately hung by a silk fibre in the centre of a large coil of many turns of wire. The mirror is made to reflect the image of a wire on to a horizontal scale, and any disturbance of the needle causes a magnified movement of the image of the wire on a scale (p. 216). The image of the wire

can be adjusted, and the sensitiveness of the instrument altered, by means of a directing magnet sliding in a vertical rod (Fig. 197).

76. ELECTRICAL RESISTANCE.

i. Electrical resistance.—Connect one pole of a Bunsen's cell with one of the binding screws of a galvanometer. Join the other binding screw of the galvanometer and the other pole of the battery with a yard of fine German-silver wire. Notice the amount of deflection of the pointer. Substitute a yard of thinner German-silver wire for the first piece, and again notice the deflection. It will be much less. Similarly compare the electrical resistances of pieces of thick and thin copper wire.

ii. Heat produced by the current.—Connect the poles of a strong battery to the ends of a short fine platinum wire. The platinum rapidly becomes red hot. If silver wire of the same diameter be substituted for the platinum the heat noticed is practically nothing.

Difference of potential or electro-motive force.—A wire connecting the poles of a Bunsen or other cell is said to have an electric current flowing along it. The use of such words as "flow" and "current" will probably suggest previous facts which the student has learnt. It has been seen that a flow of heat takes place from a body of high temperature to one of low temperature when the bodies are placed in contact, and that such flow continues until both bodies are of the same temperature. Similarly, there is a flow of water from one vessel to another, when they are connected, if the level of the water in one vessel is higher than in the other. Hence we are face to face with the question, What difference is there between the poles of a cell which causes the condition of things called the electric current? The name given to the difference of condition in the plates of the battery which corresponds to temperature and water-level is **difference of potential.** The electric current continues to flow along the copper wire until the potential of the two plates becomes the same; then the current ceases.

There is another way of regarding the flow of electricity. It has been seen that matter requires a force to move it or to change its motion. It is customary also to speak of the force which causes the flow of electricity. Regarding the cause of the electric current as a force, it is generally spoken of as

electro-motive force—that is, the force which tends to cause a movement of electricity. The cause of the electric current is then spoken of either as a **difference of potential** or as an **electro-motive force**, often written **E.M.F.**

Forces of two kinds have now been mentioned, and these should be carefully distinguished; one is a matter-moving or matter-motive force, the other is an electrically-moving or electro-motive force.

If several different voltaic cells are made in succession to send a current through the same galvanometer, the amount of current sent will vary in each case. Each cell urges the current in a different degree. The cells differ in electro-motive force.

Cause of the electric current.—The difference of potential between the poles of a battery which causes the flow of electricity is maintained by the solution of the zinc in the acid. Or, expressed in terms of energy, the work of maintaining the current is performed by the solution of the zinc. This is similar to the maintenance of the work done by a steam-engine by the burning of the coal in the furnace. It is for this reason that the zinc decreases in weight after the electric current has flowed for some time along the wires joining the poles of a voltaic cell.

Electrical resistance.—Not only does electricity, like matter, require a force to move it, but its motion, like that of matter, may be impeded and stopped. Bodies which allow electricity to flow along them are called conductors (p. 263). Substances differ greatly in their capacity to conduct electricity. All conductors, however, offer some **resistance** to the passage of an electric current. If two perfectly similar cells are used to send electric currents through wires of the same substance and length, but of different diameters, the current through the thicker wire will be the greater. The thinner wire offers a greater resistance. With the same E.M.F. a weaker current flows through a long wire than through a short one. There is a good analogy between the flow of electricity along conductors and the flow of liquids through tubes. With the same pressure urging it, water flows in greater quantity through a wide tube than through a narrow, and the resistance to its flow is less in a short than in a long one.

Different substances resist the passage of an electric current in different degrees. Through some, the electro-motive force can barely force the current at all, while through others it flows with considerable strength.

Referring again to the analogy of the passage of an electric current through a wire and the flow of liquids through tubes, the substance of high resistance would be represented by a tube corroded and choked with dirt, while the good conductor corresponds to a clean tube permitting an uninterrupted flow of liquid.

Heating of wire by the electric current.—The quantity of heat produced in the wire depends upon the resistance of the wire, the strength of the current and upon the time during which the current passes. The quantity of heat produced in a wire may be measured by twisting it into a spiral and placing it in a test tube containing a known weight of water and a thermometer.

A very familiar instance of the heat produced by a current is the glow lamp, the carbon of which offers very considerable resistance and consequently becomes red hot.

CHIEF POINTS OF CHAPTER XIX.

The simple cell.—When plates of copper and zinc, immersed in dilute sulphuric acid, are connected outside the liquid by a wire, bubbles of hydrogen are given off at the copper plate and the wire acquires the property of influencing a magnetic needle.

After use, the accumulation of gas on the copper plate causes *polarisation* and stops the current.

In Daniell's, Grove's, and Bunsen's cells this defect is remedied.

A coil of wire conveying a current is found to act in every essential like a magnet.

The electro-magnet.—When a coil of wire conveying a current is supplied with a soft iron core its magnetic power is greatly strengthened. Such a combination is called an *electro-magnet.* Electro-magnets are of varied shapes: bar, horse-shoe, or closed ring.

A galvanometer is an instrument for detecting and measuring an electric current.

Difference of potential causes the flow of the electric current through conductors. It is produced in voltaic cells by **electro-motive force** (often abbreviated E.M.F.).

The electrical resistance of a substance is its property of hindering the flow of electricity through it.

EXERCISES ON CHAPTER XIX.

1. Explain the cause of *Polarisation*, and describe the chief methods of preventing it.

2. Two compass needles are arranged near each other so that both point along the same straight line. A wire connecting the platinum and zinc ends of the battery is stretched vertically half-way between the needles. How will the current in the wire affect the needles, and how will the result depend upon whether the platinum terminal is connected with the upper or lower end of the wire respectively?

3. What are the materials used in the construction of a Daniell cell? and what chemical changes occur in the cell when in action?

4. A plate of pure zinc, and a plate of copper, are dipped into dilute sulphuric acid, and then connected by copper wire. What changes take place in the plates, wire, and acid, when the circuit is complete?

5. When a galvanic cell, consisting of zinc and copper plates immersed in dilute sulphuric acid, has its terminals joined by a wire, the E.M.F. rapidly decreases. How do you account for this? Describe a cell designed to prevent this decrease in E.M.F., and explain how it acts.

6. A long straight wire is stretched on a table in the direction of the magnetic meridian, and a dip circle, with its plane parallel to the magnetic meridian, is placed on the table near to the wire and on the west side of it. Will the dip of the needle be altered when an electric current is passed along the wire from south to north, and, if so, how? Give reasons.

7. A straight horizontal wire is placed near and parallel to a compass needle, and in the same horizontal plane with it. If a current is then passed through the wire, what effect is produced on the needle, and what occurs if the wire is (i.) slightly raised, (ii.) slightly lowered?

8. Describe a simple experiment by which it may be proved that a long wire has electrical resistance.

CHAPTER XX.

CHEMICAL CHANGE INDUCED BY THE ELECTRIC CURRENT.

77. ELECTROLYSIS.

i. The passage of the electric current through liquids.—Fit up a Bunsen's cell for the generation of an electric current. Attach pieces of platinum foil, by means of suitable binding screws, to the ends of two copper wires. Attach one of these wires to one pole of the battery. Connect the other pole to one of the binding screws of a simple galvanoscope, and to the other screw of the galvanoscope attach the remaining wire with the platinum plate on the end (*see* Fig. 199). Dip the platinum plates—1st, into some mercury, and notice there is a great deflection of the needle of the galvanoscope and no alteration of the mercury; 2nd, into some turpentine, and notice there is no deflection of the needle; 3rd, into some acidulated water, and notice there is a smaller deflection than in the first case, and at the same time there are bubbles of gas given off from both platinum plates.

ii. Electrolysis of copper sulphate.—(*a*) Make a strong solution of copper sulphate (blue vitriol) in water. Pour some into a beaker. Dip two platinum plates connected by copper wires, as previously described, into the solution, and notice that after a few minutes a deposit of copper is found on the *kathode*, or plate connected with the negative pole, and that bubbles of gas (which if collected and tested are found to be oxygen) are seen to rise from the *anode*, or plate connected with the positive pole.

(*b*) Arrange the apparatus as described in Experiment 77 i., substituting copper plates for the platinum ones there used, and weigh the copper electrodes before passing the current. After the current has passed for, say, ten minutes, break the circuit and weigh the electrodes again. Notice the anode has lost in weight by a certain amount, and that the kathode has increased in weight *by the same amount*. There is no evolution of oxygen.

There is a continual withdrawal of metallic copper from the solution by the passage of the electric current, and the liquid is gradually converted in[illegible]uric acid. This can be demonstrated by its effect on a piece [illegible]mus paper.

The passage of the electric current through liquids.—1st Case. Passage of the current through mercury.—The student will learn shortly that mercury is one of the chemical elements, and that it is so regarded because it can be decomposed neither by the passage of an electric current nor by any other known methods. The great deflection of the needle of the galvanoscope reveals the fact that a considerable current passes through its coil of wire, hence we say that mercury is a good conductor of the electric current, or expressing the same truth in other words, that it offers very little **resistance** to the flow of the current.

Similarly, it would be found that other metals, when in the liquid condition, which they can assume if the temperature is raised sufficiently, are also good conductors of the electric current.

2nd Case. Passage of the current through turpentine.—There is in this case no deflection of the needle of the galvanoscope; it is therefore evident that no current passes through the coil of wire round the needle, and since the battery is arranged precisely as in the experiment with the mercury, the explanation must be that the turpentine prevents the flow of the electric current round the circuit. Turpentine is consequently known as a **non-conductor**, a class of bodies which also includes such liquids as petroleum and other oils.

3rd Case. Passage of the current through acidulated water.—Here we have the current conducted and the liquid decomposed by the passage of the current. This is the condition of things in all liquid compounds which conduct the electric current. Such a decomposition as this is known as **electrolysis**, and we shall have to study this case more fully.

Electrolysis of water.—Pure water is a very bad conductor of the electric current, and hence it is necessary to add a drop or two of acid to make it conduct. To understand exactly the result of the passage of the electric current, some means must be devised by which the gases which appear at the platinum plates can be collected. Such an arrangement constitutes what is known as a **voltameter**. A convenient pattern to use for the decomposition of water consists (Fig. 198) of a glass vessel in the bottom of which are fixed two slips of platinum connected, by means of copper wires, with binding screws. Before

connecting the binding screws with the poles of a battery, acidulated water is poured into the vessel, and two glass tubes of exactly equal size and carefully divided into equal volumes by divisions etched on the glass, are completely filled with acidulated water and inverted over the platinum plates, as shown in the illustration. When the binding screws are connected with the poles of a battery, consisting of two or three Bunsen's cells, bubbles of gas are seen immediately to appear on the platinum plates, and in a few minutes it will be observed that a quantity of gas has collected in each tube. If we

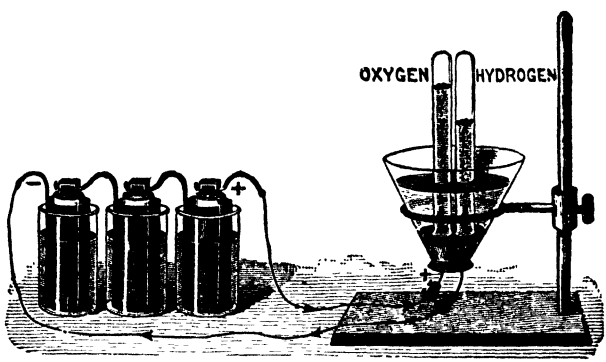


FIG. 198.—Electrolysis of water.

allow the electric current to pass round the circuit for 15 or 20 minutes and then measure the volume of the gas which has collected in each tube, it will be found that the tube over the platinum plate which is connected with the zinc pole of the battery contains twice as much gas as that over the plate connected with the carbon pole. Moreover, if the bottom of the tube containing the larger amount of gas be covered with the thumb and the tube be lifted out of the liquid, inverted, and a light applied to the gas, it will be found to burn, showing it to be hydrogen (Chap. XXV.). Similarly, if a glowing splinter of wood be plunged into the other gas it will be rekindled, showing it to be oxygen. Hence we see that the passage of a sufficiently strong electric current through water causes it to be decomposed into hydrogen and

oxygen, and that twice as much hydrogen by volume as oxygen is formed. Or,

$$\text{WATER}\left\{\begin{matrix}\text{is decomposed by the}\\ \text{electric current into}\end{matrix}\right\}\ \begin{matrix}\text{HYDROGEN}\\ \text{[2 vols.]}\end{matrix}\ \text{and}\ \begin{matrix}\text{OXYGEN}\\ \text{[1 vol.]}\end{matrix}$$

Terms used in describing electrolysis.—It is customary in speaking of the decompositions effected by the electric current to use certain terms originally adopted by Faraday. The liquid which conducts the electric current, and is itself decomposed, is known as the **electrolyte**; the platinum plates in the voltameter described, or, generally speaking, the ends of the wires coming from the poles of the battery, are called **electrodes.** Names are

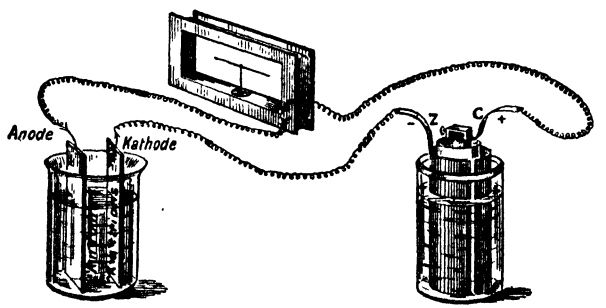


FIG. 199.—Current passing through galvanoscope and an electrolyte.

given to each of these to distinguish them. That by means of which the current *enters* the electrolyte, or what is the same thing, the electrode in connection with the carbon pole of a Bunsen's battery, or the positive (+) pole of any battery, is called the **anode.** The electrode by means of which the electric current *leaves* the electrolyte or that in connection with the zinc or negative (−) pole of the battery, is called the **kathode.** The atoms into which the molecules of the electrolyte are decomposed are referred to as **ions.** Evidently the ions are invisible during their passage through the electrolyte. Those ions which collect at the anode are spoken of as **anions,** those collecting at the kathode are called **kathions.** It will be clear to the reader that the direction in which the kathions move through the liquid is that which we have spoken of as the direction of the current, namely, from the anode to the kathode (Fig. 199).

CHIEF POINTS OF CHAPTER XX.

The **galvanoscope** is a convenient apparatus for recognizing the passage of an electric current.

The passage of the electric current through liquids.—

(*a*) Liquid metals conduct the current without being decomposed.

(*b*) Certain liquids, like turpentine, will not conduct the electric current and are consequently not decomposed by it.

(*c*) Compound liquids, which, like acidulated water, conduct the electric current, are decomposed by its passage.

Electrolysis is the term used to refer to condition (*c*) above. It means the process by which electric currents pass through compound liquids and so cause them to be decomposed. The liquid which conducts the electric current and is itself decomposed is known as the **electrolyte.**

The ends of the wires coming from the poles of the battery are called **electrodes,** that by which the current enters the electrolyte is known as the **anode,** that by which it leaves the **kathode.**

The atoms into which the electrolyte is decomposed are called **ions.**

The ions which collect at the anode are the **anions**; those which collect at the kathode the **kathions.**

Electrolysis of water.—

WATER { is decomposed by the electric current into } HYDROGEN [2 vols.] and OXYGEN [1 vol.]

QUESTIONS ON CHAPTER XX.

1. Describe fully what takes place when each of the following liquids is included in a circuit round which a strong electric current is circulating:—

(*a*) Liquid mercury.
(*b*) Petroleum.
(*c*) Acidulated water.

2. What means would you adopt to recognise the passage of an electric current through a wire?

3. What do you understand by *electrolysis*? Give an account of the electrolysis of water.

4. Explain the terms:—electrolyte, anode, kathode, ion.

5. What happens precisely when an electric current is passed through a solution of copper sulphate in water?

CHAPTER XXI.

GRAPHIC REPRESENTATION.

78. CO-ORDINATES. LOCI.

i. Plotting points.—(*a*) Transfer the points *a*, *b*, *c*, *d*, *e* from Fig. 200 to squared paper. Take two dark lines as the *axes ox*, *oy*, and use the sides of the small squares as the unit of length. The lines *ox* and *oy* are called the axes of *x* and of *y* respectively. The distance of a point from the line *oy* is called its *abscissa* and the distance from *ox* is called its *ordinate*. Read off the abscissae and ordinates of each point thus:

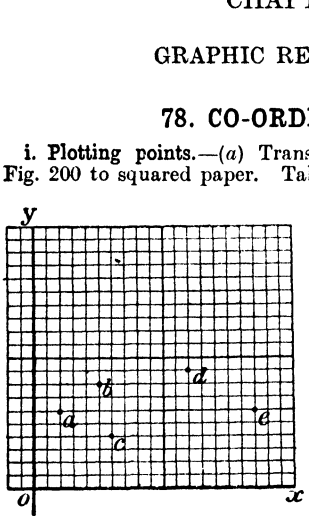


FIG. 200.—For Expt. 78. i. *a*.

	Abscissa.	Ordinate.
a	2	6
b	5	8
c	6	4
d	12	9
e	17	6

(*b*) Use two of the thick lines on a sheet of squared paper as axes; take the side of a small square as the unit of length. Mark the position of the following points.

	Abscissa.	Ordinate.		Abscissa.	Ordinate.
(1)	8	4	(5)	16	8
(2)	10	5	(6)	18	9
(3)	12	6	(7)	20	10
(4)	14	7	(8)	22	11

ii. Plotting loci.—(*a*) As before, take two dark lines on the squared paper at right angles to one another to represent the axes of *x* and *y*. Mark in succession the points having both their abscissae and ordinates equal to 1, 2, 3, 4, 5, etc., and join these

points by a line OP (Fig. 201). The line OP is called the *locus* of all points which have their abscissae and ordinates equal.

(*b*) Find the locus of all points the abscissae and ordinates of which, added together, always equal 12 units of length. Take a succession of values for the abscissae and calculate the corresponding value of the ordinate, thus:

	Abscissa.	Ordinate.
(1)	1	12 − 1 = 11
(2)	2	12 − 2 = 10
(3)	3	12 − 3 = 9
(4)	4	12 − 4 = 8
(5)	5	12 − 5 = 7
(6)	6	12 − 6 = 6
(7)	7	12 − 7 = 5
(8)	8	12 − 8 = 4
(9)	9	12 − 9 = 3
(10)	10	12 − 10 = 2

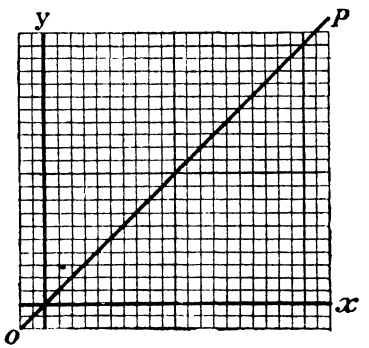


FIG. 201.—Locus of points with equal abscissae and ordinates.

Plot these points on squared paper as before (Fig. 202).

(*c*) Find the locus of the points the abscissae and ordinates of which when multiplied together equal 24. Take, as before, a

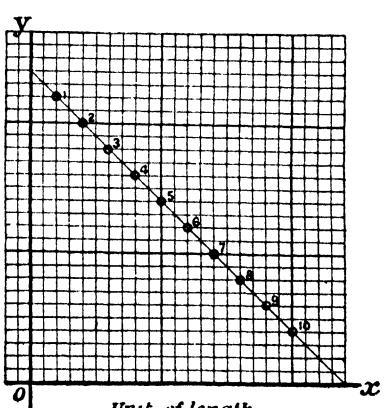


FIG. 202.—Locus determined in Expt. 78. ii. *b*.

succession of values for the abscissae, and calculate the values of the ordinates, thus, 1-24, 2-12, 3-8, etc. Find by plotting on squared paper the locus of these points.

Co-ordinates.—Suppose it was necessary for some reason to define exactly the position of any letter on this page of print; the first letter *s*, for example, in the ninth line from the bottom. How could we do it?

One way would be to count the number of lines from the bottom line of print, and the number of letters along the line itself from the outside edge of the print. This would really amount to measuring two lengths at right angles to one another. The two lines at right angles from which the measurements are made being the bottom line of print and the outside margin of the lines of print.

Any two such lines at right angles, from which measurements of this kind are made, are known as **the axes,** the horizontal line being generally called the **axis of x,** and the vertical line the **axis of y.** The point where the axes intersect is called the **origin.** Distances along the axis of x are **abscissae** (singular, *abscissa*), distances along the axis of y are called **ordinates.** The abscissa and ordinate of a point are together spoken of as the **co-ordinates** of the point.

Meaning of loci.—Notice in Experiment 78 i. (*b*) that every abscissa is twice the corresponding ordinate. Observe that if you join the points (1) and (2) with a straight line, this will, if continued, pass through each of the other six points. Similarly, the line continued would pass through all points the abscissa of which was twice the corresponding ordinate. This is expressed by saying that the straight line you have drawn is the **locus** of all points the abscissae of which are twice the ordinates. Similarly, the locus of points having their abscissae and ordinates equal is also a straight line.

The locus of points which all satisfy some given condition is **not necessarily a straight line.**

79. SYMBOLIC REPRESENTATION OF A LOCUS. GRAPHIC DIAGRAMS.

i. The equation of a curve.—(*a*) In the equation $xy = 36$ find the values of x which correspond to $y = 1$, $y = 2$, etc. Take the values of x as abscissae and those of y as ordinates and plot the locus of the equation, that is, find the curve which joins the points thus obtained.

(*b*) Find equations which will represent the lines AB, CD, in Fig. 203.

(*c*) Obtain the locus of the equation $x + y = 30$.

(*d*) Find where the curve $x - y = 2$ cuts the axes.

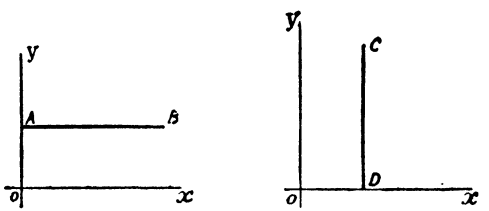


FIG. 203.

ii. Representation of varying quantities.—Construct graphic diagrams for the following cases, the number of passengers or of the population being taken as ordinates placed at convenient and equal distances apart.

(*a*) The number of 3rd class passengers by a certain popular train throughout a week :

	Passengers.		Passengers.
Monday,	250	Thursday,	220
Tuesday,	215	Friday,	185
Wednesday,	190	Saturday,	235

(*b*) The approximate population of Blackburn at the census of the following years is given in the table. As in previous exercises plot these numbers on squared paper. Join the points so obtained, and from the resulting locus read off what you would expect the population to be in 1846, 1856, 1866, 1876, and 1896.

YEAR.	POPULATION.	YEAR.	POPULATION.
1841	36,600	1871	76,300
1851	46,500	1881	104,000
1861	63,100	1891	120,100

(*c*) Repeat the preceding exercise, using the populations of the parish of St. Mary, Newington, London, S.E.

YEAR.	POPULATION.	YEAR.	POPULATION.
1841	54,700	1871	88,900
1851	64,800	1881	107,800
1861	82,200	1891	115,700

Symbolic representation of a locus.—Instead of giving different values for the abscissae of points as we have done in the previous exercises, a general expression may be used for them all, such as x. That is, let x stand for the distance from the axis of y measured along or parallel to the axis of x. How shall we represent each of the loci we have plotted ? Let y stand for the distance of any point on the locus from the axis of x measured along or parallel to the axis of y. We proceed as follows :

To find the locus of points the abscissae of which are twice the ordinates. The abscissae are represented by x, the ordinates by y. By the terms of the problem x is always twice as great as y, or, $x=2y$, is an equation which stands for the line plotted in Experiment 78 i. (*b*).

To find the locus of all points the abscissae and ordinates of which are equal. Here, as the student will at once see from Fig. 201, the equation is $x=y$.

The equation of the curve obtained in Experiment 78 ii. (*c*) is $xy=24$. Notice carefully the kind of curve which corresponds to this equation.

Graphic diagrams.—Newspaper records of thermometer and barometer readings are often made on ruled paper, the actual mode depending upon the particular newspaper. In every case the changes of reading from day to day are shown by a wavy line joining the tops of ordinates placed at equal distances apart. This mode of exhibiting a variable quantity has certain advantages over a list of figures. The steepness of the line joining two ordinates indicates very clearly the rate at which the values change. The graphic method of exhibiting the results of experiments is of frequent use in the laboratory.

Instances in which it is employed with advantage are the results obtained by stretching a piece of india-rubber or an

elastic spring, bending a lath, in experiments with levers, or on fluid pressure, etc.

Graphic representation of a series of results will enable us sometimes to discover a law, or will frequently show whether a supposed law accurately represents the facts of an experiment. Thus, if the results obtained in Experiments i. and ii. on Boyle's Law (p. 114) be plotted, a smooth curve as shown in Fig. 204 will be obtained. The points in Fig. 204 correspond to the following results :

PRESSURES (in cm. of Mercur .	VOLUMES (in c.c.).	PRESSURES (in cm. of Mercury).	VOLUMES (in c.c.).
105·0	7·7	67·3	12·0
90·8	8·8	57·3	15·9
77·9	10 2		

Plot your own values in this way.

Join in the points so located on your squared paper with a fine pencil mark. If this continuous line does not form a smooth curve, if, for instance, one point is a long way out, it will probably mean you have made a mistake in your reading of pressure or volume. If the fine pencil mark is only slightly

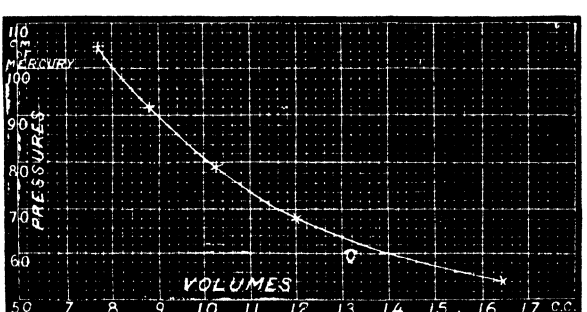


FIG. 204.—Graphic representation of Boyle's law.

irregular, then the irregularities show you the departure from strict accuracy of which you have been guilty in your experiments. In this case sketch in a smooth curve which shall have as many of the points on one side of it as on the other. Such a

curve represents the way in which pressure and volume vary together when the temperature remains constant.

Moreover, since volume × pressure always gives the same product, we can say that the equation

$$p \times v = \text{some constant number}$$

represents the smooth curve we have obtained.

Solubility Curves —An interesting and important application of graphic representation is to show easily how the solubility of a solid in a liquid varies with the temperature. Thus, Fig. 205 shows the number of grams of the three solids, nitre, common salt, and chlorate of potash, which will dissolve in 100 grams of water at different temperatures. The degrees on a Centigrade thermometer are marked along the bottom horizontal line, and the length of the side of one square represents five degrees.

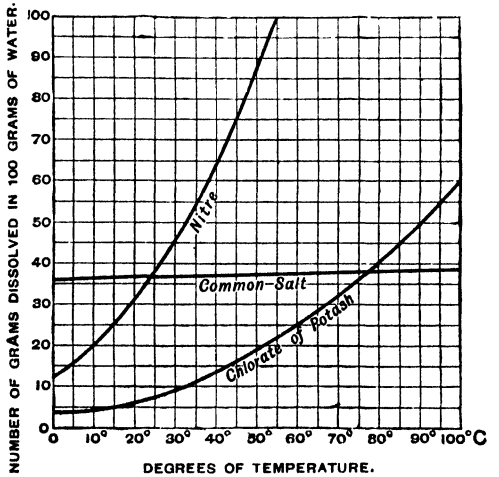


FIG. 205.—The number of grams of nitre, common salt, and chlorate of potash which can be dissolved in 100 grams of water at any temperature from 0° to 100° C. is shown in this diagram.

The number of grams of solid which the 100 grams of water contain is read off from the scale on the left-hand of the diagram. The length of the side of one square represents five grams of dissolved solid. Thus, a reference to Fig. 205 shows that 100 grams of water dissolve at 0° C. 12½ grams of nitre.

At 5° C.	100 grams of water	dissolve	15	grams of nitre.
10° C.	,,	,,	20	,,
15° C.	,,	,,	25	,,
20° C.	,,	,,	32	,,
25° C.	,,	,,	$37\frac{1}{2}$	,,
30° C.	,,	,,	45	,,
35° C.	,,	,,	55	,,
40° C.	,,	,,	64	,,
45° C.	,,	,,	75	,,
50° C.	,,	,,	$87\frac{1}{2}$	,,
55° C.	,,	,,	100	,,

We could read off the amounts of common salt and chlorate of potash dissolved in 100 grams of water at different temperatures in just the same manner.

But when we have several solubility curves together in this way, we can very easily compare the solubility of the different solids together. We see, for instance, that the amount of nitre which will dissolve in 100 grams of water increases very rapidly as the temperature rises, as the steepness of this particular curve shows. The amount of common salt which 100 grams of water will dissolve increases very little as the temperature rises. The curve is almost a horizontal line, for while at 0° C. about 36 grams are dissolved by 100 grams of water, at 100° C. the amount in solution is only 38 grams.

CHIEF POINTS OF CHAPTER XXI.

Co-ordinates. The position of a point on squared paper is fixed by its distances from the co-ordinate axes. Distances measured along *ox* are represented by x and are called *abscissae*; those measured along *oy* are represented by y and are called *ordinates*.

The locus of a point is the curve which indicates the various positions of a point of which the abscissae and ordinates are related in some definite manner.

The equation of a curve represents the fixed relationship between the abscissae and the ordinates of all points in the curve.

Graphic representation of varying quantities is employed to make clear the rate of rise or fall in the value of the quantities.

Graphic diagrams are used to exhibit the results of experiments in which two quantities are being compared. The form of the curve obtained indicates the relationship between the two quantities.

The curve for results in experiments on pressures and volumes of a certain quantity of air at constant temperature indicates by its shape that $pv = \text{constant}$ is the equation expressing the relationship stated in Boyle's law.

EXERCISES ON CHAPTER XXI.

1. Plot the curve of which the equation is $x + y = 3$.

2. What lines are represented by the following equations ?

$x = 0.$	$xy = 1200.$
$x + y = 5.$	$x - y = 0.$

3. Suggest a use for the graphic representation of the experimental results from stretching a piece of india-rubber.

4. What do you understand by the terms axes of x, co ordinates, abscissa, locus, equation of a curve ?

5. In comparing the readings of two thermometers what use could be made of squared paper ?

6. With inches as ordinates and centimetres as abscissae, construct a diagram to show the relation between inches and centimetres up to a length of 20 centimetres.

7. Use the following numbers to construct a diagram showing the relation between lbs. and kilograms :

lbs.	kilos.	lbs.	kilos.
2	0·9	6	2·7
4	1·8	8	3·6

8. Taking the value of a rupee to be 15 pence, construct diagram to show the relationship between shillings and rupees.

CHAPTER XXII.

INTRODUCTORY. SIMPLE CHEMICAL OPERATIONS.

80. INTRODUCTORY.

Physical and chemical changes.—Matter is subject to two kinds of change. Hitherto we have been chiefly concerned with those which influence the properties of matter, leaving its composition unaltered. It has been seen that a body, such as a piece of iron, may gradually increase in temperature, changing from cold iron to hot, and, becoming hotter and hotter, may change in colour, passing from a dull gray to red, and from red to almost white, becoming incandescent and emitting light rays. But if left to itself the iron will begin to cool, passing through the same changes in the reverse order until it reassumes precisely the former condition; and in all these changes the weight of the iron remains unaltered. Or, again, we might take a piece of soft iron, and, having wound silk-covered copper wire round it several times, pass an electric current through the wire. It would be found, on examining the iron, that new properties had been imparted to it, that it was now able to pick up other pieces of iron, or had become magnetised. If the electric current be discontinued, the new power, too, disappears. Such changes as these, where the substance or composition of the body remains unchanged, are known as **physical changes.** On the other hand, if a piece of iron be left exposed to damp air for some hours it becomes covered with a reddish-brown powder, which the most superficial examination will show is a different substance from the iron with which we started. There is a very large number of changes of the same kind as this continually taking place around us.

When gunpowder explodes, we have an abundance of smoke formed and a black residue left behind, and it is easy to see that the smoke and deposit are quite unlike the gunpowder before the flash. Such changes as these are called **chemical changes.** It is with changes of the second kind that Chemistry is concerned, and it may be defined thus: **Chemistry is that branch of knowledge which deals with chemical changes; those, namely, which result in the formation of new substances with new properties.**

Chemical elements.—The result of a large number of experiments made from time to time by different chemists has been to show that there are upwards of seventy different forms of matter which can by no known methods be broken up into anything simpler. By this is meant that if any one of these be selected and treated in any way with which chemists are familiar—for example, if it were subjected to a very high temperature—it would be found impossible to obtain anything having properties different from those of the substance itself; bodies of this simple kind are called **elements.**

But it must be borne in mind carefully that, as the methods which chemists adopt become more and more refined, it is quite likely that some of these may be found wrongly to be regarded as elements. Up to the time of Davy (1807) the substances soda, potash, and lime were regarded as elements. He found, however, that they could be split up into simpler constituents. From soda he obtained a soft metal, sodium, and two gases, oxygen and hydrogen, and from that time, of course, soda could not be regarded as an element. Similarly, if at any future time it should be found that any of the forms of matter which are called elements can be split up into simpler bodies with different properties, the element which is thus decomposed will have to be struck off the list.

Metals and non-metals.—A good many of the elementary substances are possessed of certain distinctive characters in which they resemble one another. They have a bright lustre, a high specific gravity (p. 62), are good conductors of heat and electricity, and are known to chemists as **metals.** There is no difficulty in deciding in a large number of instances that the substances possess the characters of a metal, and the student

will immediately think of gold, silver, copper, iron, etc. Other bodies, however, are quite as plainly not of this class; they have no lustre, they are not heavy, nor do they conduct heat and electricity well. These are spoken of as **non-metals**, and phosphorus, sulphur, and carbon will serve as good instances. But there is no hard and fast line between the two classes, for one or two substances possess some of the properties which distinguish a metal, and yet for other reasons, which the student will understand better later, are not classed with the metals, but with the non-metals. Arsenic may be cited as an instance of a substance which possesses properties common to both classes.

81. SOLUTION. EVAPORATION. DISTILLATION.

i. **Solution.**—Place a piece of sugar in water, and note that it soon disappears, and gives a sweet taste to the whole of the water, so that the particles of the sugar must be spread through the entire mass of the water.

ii. **Saturated solution.**—Weigh out 50 grams each of the following substances:—finely powdered nitre, sugar, salt; to each add water, in small quantities, with vigorous shaking after each addition. Determine thus the quantity of water necessary to form a saturated solution of each.

iii. **Solution is a physical change. Evaporation.**—Weigh out a quantity of salt in an evaporating basin and dissolve it in water. Heat gently over a Bunsen burner so that the water boils and evaporates away completely. Note that a white solid remains in the basin, and again weigh. Satisfy yourself that the weight is equal to the weight of the basin and salt before solution, and that the solid left is still salt.

iv. **Distillation.**—An arrangement for condensing steam or vapour is shown in Fig. 206. The steam that is driven off from the water in the retort passes through a long tube kept cool by being surrounded with water, and is thus condensed.

v. **Evaporation of distilled and tap water.**—Evaporate a little distilled water in a platinum or porcelain crucible. Notice the absence of any residue. Repeat the experiment with tap-water, and note the residue.

Solution.—When sugar is placed in water it disappears, but it is not lost, for we can taste it in the water. It is said to have **dissolved** and formed a **solution** of sugar. Similarly, a great number of substances can dissolve, or are soluble, in

water, but not all to the same extent, and when, at a given temperature, no more of the substance can be dissolved the solution is said to be **saturated.**

By boiling away the water the substance dissolved in it may be recovered without loss.

Hence, during solution we have in general no chemical change, but merely a change of physical state, although we shall find later that in certain cases solution may be accompanied by chemical changes.

Evaporation.—By very gently warming water or other liquid, or even by allowing it to remain exposed to the air for some time, the liquid passes off as vapour without actually boiling. Rain drops on window panes or pavement stones disappear by reason of this process which is called **evaporation.** Any dissolved substance remains behind, forming a **residue.**

Distillation.—By the evaporation of the solvent, *i.e.* the liquid in which the solution takes place, we can separate it from the dissolved substance, and this process is frequently used not only for obtaining the dissolved substance, but also

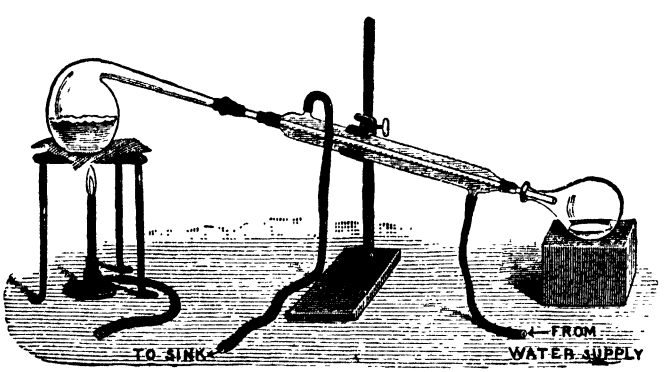


FIG. 206.—The distillation of water.

for the purification of liquid from dissolved solid material, since if the steam which is formed by boiling water containing any of these dissolved substances be condensed, the water formed is quite pure. To obtain pure water from any kinds of water, then, whether fresh or salt, all that has to be done

is to boil it and condense the steam which is given off. The dissolved materials are all left behind in the vessel in which the boiling takes place.

82. FILTRATION, SUBLIMATION, DECANTATION.

i. **Filtration.**—Mix well together some fine sand and salt: place the mixture in a glass flask with sufficient water to dissolve the salt, and shake vigorously. Take a circular piece of filter-paper and fold it in two, in the form of a semicircle, then again to the form of a quadrant; open out into the form of a hollow cone and fit it into a glass funnel (Fig. 207). Pour the liquid through this filter and note that the solution of salt runs through perfectly clear, leaving the insoluble sand in the filter paper. Evaporate off the water from the salt solution and obtain the salt.



FIG. 207.—Filtration.

ii. **Sublimation.**—Mix some sand and a little sal-ammoniac (ammonium chloride) and heat the mixture strongly in a dish. White fumes will soon be noticed escaping, and if a dry beaker be held mouth downwards over the dish it will be found that the fumes condense on the glass as a white powder which consists of the ammonium chloride driven from the mixture by the heat.

Filtration.—When a muddy liquid, that is, a liquid containing small solid particles floating in it, is poured on filter paper, a thin kind of blotting paper, the solid particles remain on the paper while the clear liquid passes through. This process is called **filtration** and is that which is usually adopted in chemical operations when it is wished to separate insoluble suspended

material from a liquid, either for the isolation of the solid or for the purification of the liquid.

Decantation.—If a powder such as chalk be stirred up with water it may be separated from the water partially by allowing the mixture to stand and settle for a while. With care the clear liquid may be poured or siphoned off from the sediment, and this process is known as **decantation.** It is impossible to remove the whole of the liquid in this way, so that filtration is the method usually employed.

Sublimation.—Sometimes heat may be used to separate a mixture of solids one of which is driven into vapour by the rise of temperature while the other remains unchanged.

A substance which on being heated passes straight from the solid to the vaporous state, and *vice versa* on cooling, is said to **sublime.** Examples of such substances are ammonia compounds, mercury, chloride, etc.

83. CRYSTALLISATION, PRECIPITATION.

i. **Solvent powers of hot and cold water.**—Place a quantity of powdered nitre in water and allow it to stand for some time with frequent vigorous shaking, so that a cold saturated solution is formed. Now heat the solution and see whether more solid dissolves or not.

Do the same experiment with other solids and satisfy yourself of the truth of the following statement, that in general hot water dissolves more of a solid than cold water, or in other words, that the solvent power of a liquid increases with rise in temperature.

ii. **Making crystals.**—Allow a hot saturated, or nearly saturated, solution of nitre obtained as above to cool slowly, and observe that the nitre separates out from the solution as clear glassy solids, which may be observed to increase in size as the solution cools. Examine a few of these and see that they are bounded by plane surfaces.

Repeat this experiment, using alum, potassium chlorate, blue vitriol, etc. Notice that each substance forms crystals of a special shape.

iii. **Precipitation.**—Dissolve some lead acetate in distilled water; add to it a solution of salt. Observe the formation of a white powder in the previously clear liquid.

Solvent Power of Hot and Cold Water.—Common experience shows that as a rule a substance dissolves more readily in hot than in cold water. Not only does solution occur more

quickly but more of the substance dissolves; that is, the solvent power of water rises with the temperature.

Crystallisation.—A hot saturated solution contains more dissolved substance than a cold one: it follows, therefore, that when the hot solution is cooled, some of the dissolved substance separates from the liquid. This it usually does in the form of **crystals**, and their formation is known as **crystallisation.**

Crystallisation usually occurs when a dissolved solid separates from its solution, either by cooling of the solution, or by the evaporation of the solvent; and the slower the evaporation, or the more gradual the cooling, the larger and more perfect are the crystals obtained. We shall also find later that crystals may be formed during the solidification of a melted solid, especially when it is allowed to cool and solidify slowly and undisturbed.

Every substance capable of crystallising forms crystals of a definite shape. Thus, common salt crystals have the shape of *cubes*, alum that of *octahedra*, each of which appears to be made from two equal square pyramids placed base to base. Other shapes are much more difficult to describe accurately, but we may say that crystals, in general, have shining faces and sharp edges.

Precipitation.—Precipitation always happens if to a solution we add some material capable of converting the dissolved substance into a new product which is insoluble in water, and therefore separates. If, for example, some common salt is added to solutions of lead acetate or of silver nitrate, a white powder forms in each case and falls to the bottom of the liquid. In the former case lead chloride, and in the latter silver chloride is formed, and both these chlorides are insoluble.

84. MIXTURES AND COMPOUNDS.

i. Separation of a mixture by solution.—(*a*) Mix some sand and salt together. Stir the mixture well with cold water and allow to stand. Pour off the clear solution which contains most of the salt. Repeat the process till the solution over the sand no longer tastes of salt. The salt may be obtained from the water by evaporation.

(*b*) Mix together some copper filings and powdered sulphur, and examine the result, which is a *mixture* of copper and sulphur. Note that its colour lies between the yellow of sulphur and the red of

copper, and that the particles are quite distinct. Wash a little in a gentle stream of water, and observe that the sulphur is washed away more readily than the copper, leaving the latter as a residue. Shake up the mixture with a little of the liquid known as carbon disulphide, and see that the sulphur dissolves but the copper is left; pour off the solution and allow it to evaporate, and observe that the sulphur is left.

ii. **Separation of solids by magnetism.**—Mix some iron filings and some sulphur in a mortar, stirring with a pestle until the mixture is as complete as possible. Now hold a magnet close to the mixture and the filings will be drawn towards the surface of the mixture. By stirring the powder well with one end of the magnet it will be found possible to remove the whole of the iron and the mixture will lose its grey colour and become yellow.

iii. **Formation of a compound.**—Now again mix some copper filings and powdered sulphur in a porcelain crucible, using a weighed quantity of copper, and heat either in the open air or in a draught cupboard. Observe that the sulphur melts and some burns away. Add more sulphur, and again heat until no more sulphur burns away, and do this three or four times. Weigh the product so obtained, and then examine it carefully. Try to separate the sulphur and copper by washing, and try to dissolve the sulphur out by carbon disulphide. You will now find that you cannot separate the two components, neither can you distinguish the individual particles. The substance formed is a *compound* of copper and sulphur, so that during the heating a change has taken place, with the formation of a new substance with properties of its own. This is an example of a chemical change—the result of chemical action. The new substance is a chemical compound. If you have taken the weight carefully you will have found also that 100 parts of copper form, roughly, 125 parts of the compound, that is, unite with 25 of sulphur, and the result does not depend on the actual weight of copper and sulphur used, all excess of sulphur over this being always burnt away.

Mixtures and Compounds.—When two substances are simply placed together so that we can distinguish each component or separate them again by some simple process in which one is removed and the other left, the substances are said to form a **mixture.** Sand and salt may be thus mixed, in any proportion, and they may be easily separated by dissolving the salt in water. Iron and sulphur when mixed can be separated by using a magnet to draw out the particles of iron. Mixtures of liquids, for example spirits and water, may be made and again separated by distillation. Mixtures of gases may be separated by solution.

When two substances are mixed and so treated, by heating or otherwise, that we can no longer easily distinguish or separate

them, and the new substance has properties unlike those of either, a **chemical compound** is formed and the substances are said to have **combined.** Very many experiments have shown that compounds are formed in fixed proportions only. One substance will combine with another always in the same proportion no matter how much of either is used in the experiment. We thus find the following differences between a mixture and a compound:

Mixture.—In a mixture the components exist side by side and can be separated by simple mechanical methods. The ingredients may be present in any proportions, and the properties of the mixture are intermediate between those of the constituents.

Compound.—In a compound the components cannot be separated by the simple means available in the case of mixtures. The properties of the compound are quite different from those of the constituents, and these constituents are always present in certain definite proportions which for each compound are invariable.

In all cases of chemical action it is most important to remember that the total weight remains absolutely unchanged, that is, the total weight of all the products is exactly equal to the total weight of all the components forming these products.

85. CHEMICAL DECOMPOSITION AND COMBINATION.

i. Chemical Decomposition.—(*a*) Heat some *red oxide of mercury* in a test tube. The powder will be found to darken in colour and a thin gray film consisting of mercury drops of minute size will be noticed to form on the cooler part of the tube. At the same time a gas (oxygen) escapes from the oxide and will re-kindle a smouldering cedar wood splint held in the mouth of the tube. The powder has decomposed into mercury and oxygen.

(*b*) Sodium (see caution on p. 334) dropped into water causes the production of bubbles of hydrogen gas due to the decomposition of the water.

(*c*) Heat some crystals of potassium chlorate in a test tube and notice the escape of oxygen by using a smouldering splint as in Expt. 85 i.

(*d*) Heat some lead nitrate crystals in a test tube. Deep red coloured fumes escape and yellow oxide of lead remains.

ii. **Chemical combination.**—(*a*) A candle burning in air disappears. The constituents of the candle wick and wax combine with a portion of the air to form new compounds of gaseous nature.

(*b*) Sulphur burning in air forms a gas having a very strong smell. This gas is produced by the sulphur combining with a portion of the air.

(*c*) Magnesium ribbon when heated bursts into flame. The metal disappears and in its stead there is found a white powder, formed partly of the metal and partly of the portion of air referred to in Expts. (i) and (ii).

(*d*) A few drops of water falling on some fresh quicklime give rise to much heat. This is explained as being due to the lime combining with the water.

Chemical Decomposition, or the separation of a compound into elements or simpler substances, is illustrated in Experiments 85 i. (*a*), (*b*), (*c*), (*d*), which, however, cannot be fully understood at this stage of the subject.

Chemical Combination is illustrated in the four experiments 85 ii. (*a*), (*b*), (*c*), (*d*). We see permanent changes in appearance and in the general properties of the substances experimented upon. In nearly all cases of combination heat is produced though not always sufficient in amount to cause flame.

CHIEF POINTS OF CHAPTER XXII.

Physical changes are those in which the composition of the body experiencing the change remains unaltered. The science concerned with these changes is called **Physics.**

Chemical changes are those which result in the formation of new substances with new properties. The study of such changes is called **Chemistry.**

Chemical elements are kinds of matter which can, by no known means, be broken up into anything simpler. They can be subdivided into **metals** and **non-metals.**

Chemical operations.—Solution is the process by which some substances, when placed in water or other liquids, disappear and their particles spread through the entire mass of the water or other liquid.

When no more of the substance will *dissolve* the liquid is said to be *saturated.* Solution is generally a physical change and is unaccompanied by any change of weight.

Distillation is the process by means of which we are able to separate the dissolved substance from the solution containing it. The liquid itself is recovered by condensing the vapour.

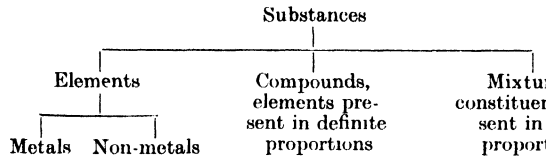
Filtration consists in utilising the porosity of unglazed paper or other material for separating insoluble substances from solutions.

with which they are mixed; or, it is a means of purifying liquids and isolating solids.

Crystallisation denotes the separation of a dissolved solid in masses having regular shapes from a solution as the liquid evaporates or cools. The regular solids formed are called crystals.

Precipitation means the formation and separation of an insoluble substance from a solution.

Compounds and mixtures.—



In a **mixture** the components exist side by side and can be separated by simple mechanical methods. The ingredients may be present in any proportion and the properties of the mixture are intermediate between those of the constituents.

In a **compound** the components cannot be separated by the simple means available in the case of mixtures because chemical combination has taken place. The properties of the compound are quite different from those of the constituents, and those constituents are always present in certain definite proportions which for each compound are invariable.

QUESTIONS ON CHAPTER XXII.

1. How may muddy water be (*a*) made clear, (*b*) purified from dissolved material?

2. Explain the differences between a mixture and a compound.

3. What do you mean by a "physical" and what by a "chemical" change? If a substance changes under the influence of heat, how would you endeavour to find out whether the change is chemical or merely physical?

4. A mixture of salt and powdered glass is given to you. How could you (*a*) obtain separately the two constituents, (*b*) find the quantity of each in the mixture? Could you separate sugar from salt in the same way?

5. What do you mean by a saturated solution? How would you prepare a saturated solution at a given temperature? What is generally the effect of cooling a saturated solution?

6. A white powder is shaken up with water. How would you ascertain whether any of it dissolves?

7. What is meant by the statement that two liquids mix? Give examples. Can a liquid be soluble in another liquid without being capable of mixing? If so, give an example.

8. How may fresh water suitable for drinking be obtained from sea water ?

9. By what means would you endeavour to find the quantity of a given solid which a given quantity, say 1 gallon, of water is capable of dissolving ?

10. What is meant by "precipitation" ? In what circumstances may precipitates be formed in a liquid ?

11. How would you endeavour to obtain large crystals of alum ?

12. Describe an experiment to show the solubility of ether in water.

13. Sand and salt are stirred up in a bottle containing water in which some gas is dissolved. What methods would you use to recover the sand, salt, and gas from the water ?

14 Give examples of the solvent power of water. How could you determine whether a sample of water contains dissolved solid matter ?

15. Describe the apparatus you would employ to obtain water free from dissolved substances.

16 State in general terms the differences observable between a mixture and a compound of any two substances.

17. Explain the terms Mixture and Compound. How could you show that copper and sulphur (or iron and sulphur) can exist together, either as a mixture or as a chemical compound ?

CHAPTER XXIII.

BURNING AND RUSTING.

86. RUSTING OF IRON.

i. Increase of weight when magnesium is burnt.—Weigh a crucible and its lid with a piece of magnesium, which, folded lightly, is placed in the crucible. Heat it strongly in a burner, taking care to let no fumes escape (Fig. 208) To do this, keep on the lid, and only raise it a little when the flame is removed. The magnesium is seen to burn brightly in places, but if care is taken, no fumes are lost. When finished, the mass should be in the form of a white powder. Allow to cool, and weigh the crucible with the lid and powder. Subtract the weight of the crucible and lid to find the weight of the powder. The powder will almost certainly be found to weigh more than the original piece of metal.

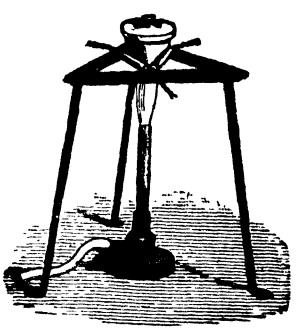


FIG. 208.—How to heat magnesium powder or ribbon in a crucible.

ii. Increase of weight when iron rusts.—Carefully weigh a watch-glass with some iron filings or tacks in it. Because iron rusts best when damp, add a few drops of water to the iron in the watch-glass, and allow it to stand for a day or two. At the end of this time warm the watch-glass gently, so as to evaporate any water left. When the rusty iron is *dry*, weigh the watch-glass and its contents.

Its weight will be found to be more after the rusting has taken place; the iron in getting rusty has gained in weight.

iii. Air absorbed during rusting.—Place some iron filings in a muslin bag, and tie the bag to a piece of glass rod. Moisten well, and place it in a bottle of air inverted over water. If necessary, put something on the bottle to keep it upright (Fig. 209) Examine after a few days. It will be seen that the water has risen in the

bottle, showing that part of the atmosphere has been abstracted by the iron in rusting.

iv. Alteration of air when iron rusts in it.—Place a ground glass plate tightly under the mouth of the jar so as to allow no water to escape; set the jar upright, and plunge a burning taper into it. Notice that the flame is extinguished; do not throw away the water.

Effects of heating metals in air.—Some of the effects of heat upon substances have already been observed and described. It has been seen that, by heat, solids may be melted into liquids, and liquids converted into vapours. If a piece of platinum wire or platinum foil is held in the smokeless flame of a laboratory burner, it becomes red-hot; but when it is taken out of the flame it quickly resumes its ordinary colour, and no change can be seen to have taken place. Substances like paper and wood when heated strongly in air take fire and burn. Smoke is given off, and an ash remains. Metals like copper become covered with a film, or tarnish, when they are heated in air. Lead when strongly heated melts and becomes covered with a scum on the surface. If the scum is removed the bright metal is seen shining below, but it also quickly dulls.

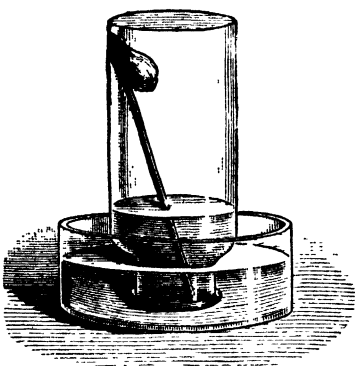


FIG. 209.—Experiment to show the action of iron filings on air.

If metals are heated in closed tubes without air they do not change in this way. It thus seems that the tarnish is due to the absorption of something from the air. But if something is taken from the air when a metal tarnishes, or when a metal like magnesium burns, the tarnished metal, or the ash of the magnesium, should weigh more than the original substance. Observations show that this is actually the case. As air seems responsible for the effects described, it is evidently well worth investigation.

Chemical properties of air.—It is necessary carefully to consider the changes which different substances undergo when

exposed to the atmosphere. It is best to begin with those cases which appear to be simplest. Every one has noticed that when iron is exposed to damp air it becomes rusty. Does the iron lose or give up something when it rusts? Or, does it, on the contrary, take up or gain something? These questions can be best answered by properly arranged experiments.

Iron gains in weight during rusting.—If a known weight of iron is allowed to rust by contact with damp air, it can be shown easily by weighing it after the rusting has taken place that it has increased in weight. The result of this experiment is very important. If the weighings are carefully made, the iron is always found to gain in weight when it rusts. The substance causing the increase of weight, when damp iron filings rust, could come from the water, or moisture, or from the air. If the iron is allowed to rust in a closed space, and the experiment is so arranged that if anything is taken from the air the loss can be detected, it can be decided whether the air causes the rusting. Fig. 209 shows a convenient way of doing this. Some iron filings are placed loosely in a muslin bag, and the bag is tied to a piece of glass rod. The bag of filings is well moistened, and arranged in a bottle of air inverted over water in a basin, in the manner shown by the illustration. The apparatus is then left undisturbed for a day or two. When it is examined after this time, the water is seen to have risen in the bottle. Why is this? It is quite clear that there is less air in the bottle now than there was before the iron became rusty. Some part of the air has, therefore, been used by the iron as it rusted, and this part of the air has joined with the iron to help to make the rust.

Air as well as iron undergoes change.—When iron rusts, the change which it has undergone is visible. No difference can, however, be seen between the character of the air left in a bottle in which iron has rusted and ordinary air. But there is a great difference. A lighted taper is extinguished by the gas left in a bottle in which iron has rusted, hence the gas cannot be air, for in air a taper will burn quite easily. But before the rusting of the iron took place in it, the air *was* ordinary air. Hence, it is clear that the rusting of the iron is accompanied by a change in the character of the air in the bottle. It is reasonable to suppose that the gas which disappears is concerned in

the formation of iron rust, and this will be shown later to be actually the case. When iron rusts, it takes out of the air that part of it which helps burning, and moreover, the iron and the part of the air concerned in burning combine together to form iron rust. The part of the air left in the bottle will not let things burn in it. It may therefore be stated that: **Iron in rusting gains in weight, taking some material from the air, and this material is the part of the air which causes substances to burn in it.**

87. CHANGE CAUSED IN AIR BY RUSTING.

i. **Volume of air used up in rusting.**—Measure in a graduated vessel the quantity of water in the bottle from Expt. 86 iv. This is equal to the quantity of gas which has been used up and has joined with the iron. Also measure the quantity of water the bottle holds. This gives us the volume of air the bottle originally held. What fraction of the total volume of air enclosed has been used up?

ii. **Change produced in air by iron rusting in it.**—Repeat the experiment of allowing iron to rust in an enclosed volume of air. After the iron has been left for a day or two, and there is no further rise of the water, mark the level of the water in the jar by a narrow strip of gummed paper on the outside. Carefully introduce another muslin bag of iron which is not rusty. This can be done by using a large enough basin of water, and pushing the bag through the water, being careful to allow no more air to get into the bottle. Examine the bottle after another day or two. There is no further rise in the level of the water, and the iron is not rusty. Evidently the gas which is left will not allow more iron to rust in it, though it is colourless and transparent, like ordinary air.

What proportion of the air is taken by iron in rusting?—Only a certain proportion of the enclosed air is taken up when iron rusts in it. Suppose some damp iron filings are allowed to rust in an enclosed amount of air, contained in a bottle inverted over a basin of water. The amount of water which rises into the bottle can be measured by means of a graduated vessel. A moment's thought will tell you that, as this water gradually takes the place of the part of the air which the iron uses, its volume must be the same as the volume of the gas so taken out of the air. The amount of water the bottle holds when full can be found easily, and the result shows the volume of air in the bottle to begin with. If observations of this kind are made, it will be found that, when the bottle is full, it has five times more water in it than it has after the iron

has rusted. Even if the experiment is repeated several times with bottles of different sizes, the result is always the same. It is always found that **one-fifth of the volume of the enclosed air is used up by iron in rusting.**

Chemical composition of the air.—That part of the air which helps substances to burn, and is taken out of the air by iron in rusting, may be called the **active part** of the air. That part which is left by the iron, and will not allow a taper or candle to burn in it, may be called the **inactive part.** The observations just described show that air is made up, or composed, of 1 volume of the active part to 4 volumes of the inactive part, in every 5 volumes. In other words, in 100 pints of air there are 20 pints which will unite with iron to make iron rust, or will assist a candle to burn, and 80 pints of the inactive part which will not assist burning.

Other metals combine with the active part of air.—When copper is heated in air, it gradually blackens and increases in weight. When hot, copper has the power of combining with

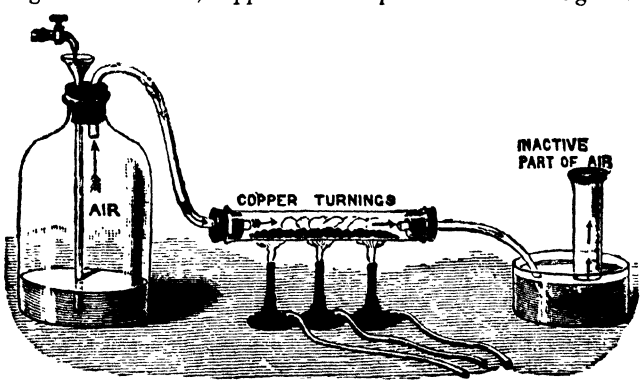


FIG. 210.—When air passes over hot copper, it is deprived of its active part, and the inactive part may be collected as shown.

the active part of the air in just the same way as the iron does gradually when cold, and it is reasonable to conclude that the black substance formed is copper rust, though it is not generally known by that name.

That copper only combines with the active part, and leaves the inactive part, can be shown by a suitable experiment, though

not so easily as in the case of iron. Some copper turnings are placed in a hard glass tube like that shown in Fig. 210, one end of which is connected with an aspirator full of air, and the other by means of a well-fitting cork with a tube which dips under water in a trough. Then a bottleful of water is inverted in the trough of water exactly over the end of the small tube, which dips into it, and is connected with the hard glass tube containing the copper. The copper is heated strongly, and air is forced over it by making water take the place of the air in the aspirator. As the air passes over the heated copper, the active part of the air joins with the copper to form the black copper rust, and the inactive part passes on alone into the inverted bottle in the trough. That the inactive part collects in the bottle is indicated, though not proved, by the fact that it puts out the flame of a taper. If this gas is, in the same way, passed over some more heated copper, it has no effect on it; the copper does not blacken. Moreover, if the amount of air which has come out of the aspirator is measured, and also the amount collected in the bottle, it is found that in passing over the copper the air loses one-fifth of its volume.

88. BURNING OF PHOSPHORUS.

(*Read the caution on p. 389.*)

i. **The burning of phosphorus.**—Place a little phosphorus upon a slate, tile, or an old saucer. Apply a light to it. It catches fire and burns brightly. As it burns, dense white clouds are formed.

ii. **Volume of air used up when phosphorus burns.**—Place a little red (or yellow) phosphorus in a test-tube fitted with a good cork. Fix the cork firmly in the test-tube. Hold the test-tube slantingly, by means of a test-tube holder, over a flame for a second or two, so as to heat the phosphorus and make it burn. When it will burn no longer, take away the test-tube and let it cool for five or ten minutes.

Then hold the mouth of the test-tube well under water, and carefully take out the cork. Water rises inside the tube to take the place of the air used up. Notice that the rise is practically one-fifth the volume of air enclosed.

Phosphorus readily burns in the air.—It is only necessary to touch a piece of dry phosphorus with a hot wire to make it catch fire and burn. It burns with a dazzling bright flame, and

at the same time dense clouds of white fumes are formed, which spread throughout the room. These facts are noticed until all the phosphorus has disappeared.

What happens when phosphorus burns in this way? Is the change anything like that when iron rusts? Does the phosphorus gain or lose in weight? These and several other questions present themselves, and they can be answered now.

Changes produced in air by burning phosphorus.—To decide whether phosphorus in burning causes the same change in air as iron does when it rusts, it is best to burn some phosphorus in an enclosed amount of air in a way similar to that which has already been described for an experiment with damp iron. One way to do this is to place a little phosphorus on a cork or basin which floats on the surface of water, under a bell jar, or a stoppered bottle having no bottom. After the experiment is over, and the fumes have disappeared, the water is seen to have risen in the jar, indicating that there is less gas in the jar than before the phosphorus was burnt in it (Fig. 212)

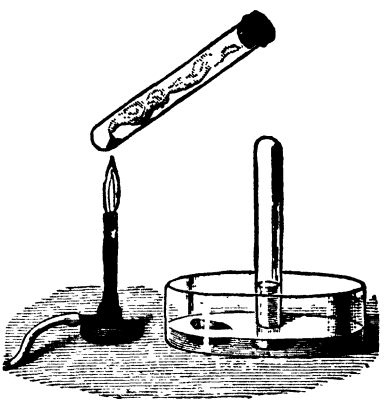


FIG. 211.—After the phosphorus has burned in the corked test-tube, the test-tube is uncorked under water, and water enters to take the place of the air used up.

From what has been previously said, it can be understood at once that phosphorus in burning takes out the active part of the air, and leaves the inactive part behind. So far, then, the changes which occur when phosphorus burns are like those when iron rusts. Some differences will be studied a little later.

The fraction of the air which disappears as a result of the burning of the phosphorus in a stoppered jar, can be measured easily enough after the jar has been raised a little, so that its mouth is still under water, but no longer rests on the bottom of

the basin. As in the case of the rusting of iron, one-fifth of the air is taken out of it by the phosphorus in burning.

That the gas left behind is really the inactive part of air can be proved by quickly pulling out the stopper of a jar in which phosphorus has been burnt, and introducing a lighted taper. The flame is at once extinguished.

Phosphorus slowly takes out the active part of the air without being lighted.—It has been seen that iron slowly takes the active part of the air and combines with it to form rust. This happens without heating the iron. Will ordinary phosphorus do the same when it is not alight? This question, too, is easily answered by a simple experiment. When a piece of clean phosphorus is exposed to an enclosed quantity of air over water, the rapid changes just described take place slowly. The only difference in the two cases is the rate at which the active part of the air is taken out. Burning phosphorus combines with the active part very quickly; cold phosphorus but slowly. Still, given time enough, ordinary phosphorus will remove all the active part of air, and at the end of the experiment it will be found that again one-fifth of the air has disappeared.

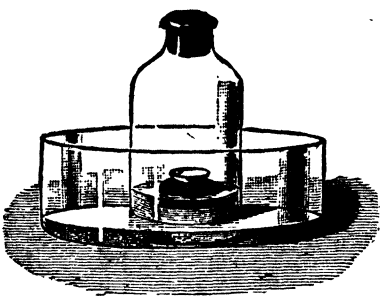


FIG. 212.—Phosphorus uses up one-fifth of the air in the bottle when it burns, and water rises to take the place of the air used.

89. BURNING OF A CANDLE.

i. Moisture is formed when a candle burns.—Over a burning candle hold a clear cold tumbler, which has been carefully dried inside and out. Notice that the inside of the tumbler becomes covered with mist, and, after a short time, drops of water are formed, which run down the sides of the tumbler (Fig. 213).

ii. Properties of the gas left after a candle has burnt in air.—Wind a piece of copper wire round a small candle and light the candle. Push the top of the wire through a small hole in a disc of cardboard, and then lower the candle into a dry, clear glass bottle in such a manner that the top of the jar is covered by

the cardboard disc (Fig. 214). Observe that the flame of the candle becomes dimmer and dimmer, and soon goes out altogether. Water collects on the inside of the jar, as in the last experiment. Take out the candle, and cover the jar with a greased glass plate. Quickly insert a burning taper, or the re-lighted candle; it is at once put out. Pour in a little fresh clear lime-water, and shake it up in the jar; notice that it is turned milky.

iii. **Volume of air used by a candle in burning.**—Fix two or three small candles of different lengths upon the inside of the top of a tin canister (Fig. 215). Float the lid upon the surface of water in a basin, or sink it to the bottom of the basin, if the candles project well above the surface while it is in that position. Light the candles, and



FIG. 213.—To show the formation of moisture by the burning of a candle.

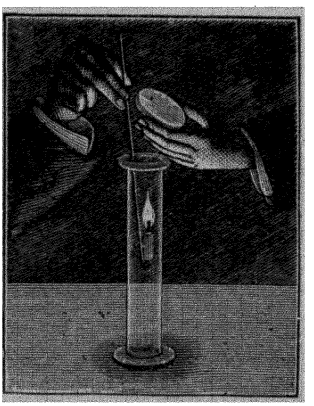


FIG. 214.—Burning a candle in air in a glass jar.

while they are burning, hold over them a wide-mouthed bottle, so that the mouth of the bottle is beneath the surface of the water. When the candles have gone out, and the air in the bottle has become cool again, mark the place to which the water has risen, by means of a strip of gummed paper. Take out the bottle and find the volume of water which just fills it. Find also the volume required to fill it to the edge of the gummed paper. Subtract this volume from the preceding one, and thus obtain the volume of air used. Notice that it is *about* one-fifth of the whole volume of air that was in the bottle.

The burning of a candle.—Several facts have been learnt about the burning of phosphorus in air, and it will be desirable before proceeding farther to study the burning of some more

common combustible substance, such, for instance, as a candle.

In what respects is the burning of a candle similar to the burning of phosphorus, and does it differ from it in any way? It has been shown already (p. 315) that a candle will not continue to burn very long in an enclosed quantity of air. Unless the air is renewed in some way, the candle goes out. This gives us a convenient starting-point for the inquiry, Why does the candle go out, and what changes take place when the candle is burning?

Water is formed when a candle burns.—When a clear glass bottle, which has been carefully dried inside and out, is held over a burning candle, it is soon noticed that drops of

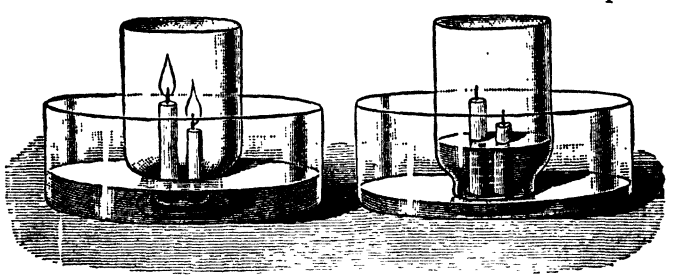


FIG. 215.—When the candles burn, a certain fraction of air is used up, and water rises into the inverted bottle to take its place.

liquid begin to collect on the inside of the bottle, and after a time they run down the sides. In some way or other, then, the burning of a candle causes a liquid to be produced. If a sufficient quantity of this liquid is collected, it can be proved to be water by tasting it, or by determining its density, or its boiling and freezing points. Water is the only liquid which boils at 100° C. and freezes at 0° C., and the density of which is 1.

Another substance besides water is formed when a candle burns.—If a candle is burnt in a clear glass bottle in the way shown in Fig. 214, the gas which is left behind can easily be examined. Experiments with this gas show that, like the inactive part of the air, it will not allow things to burn in it. But besides this it is found that *the gas turns lime-water milky*. The gas left after phosphorus has been burned in a similar jar, has not this property of making lime-water milky; it thus

appears that there is something else in the cylinder, besides the inactive part of the air already described. Hence, when a candle burns, it not only forms water, but also a colourless gas which turns clear lime-water milky. Just as it is possible to show, by burning phosphorus in a jar inverted over water, that phosphorus takes out the active part of the air as it burns, so the same fact can be made clear in the case of a candle by a similar experiment. When this experiment is performed, it is always found that at a certain stage the candle goes out. When this occurs, the water has risen and filled about one-fifth of the jar. This result has been noticed so often that there is good reason to believe that the candle is extinguished because it has used up the active part of the air, which makes up one-fifth of the whole. In fact this is true in all cases of burning in the air. It does not matter what the substance which burns is like; if it burns in air it does so because it takes out the active part to unite with it to form new substances, and the inactive part is always left behind.

The structure of a candle flame.—With care it is possible to distinguish four parts in a candle flame. (*a*) A dark inner cone surrounding the wick consists of the gases distilled from the wax or wick, and we may examine the nature of these gases by holding one end of a glass tube in the flame at this place and applying a light to the other end of the tube (Fig. 216). (*b*) Next outside this dark inner cone is found the brightest part of the flame, consisting of another cone, in which the gases from the wax and wick are partly burnt. This is the part of the flame from which soot is deposited when a cold object is placed in the flame. (*c*) Outside the bright cone is an almost colourless envelope, in which the burning is completed. It can be seen by holding a card, cut in the



FIG. 216.—Structure of candle flame.

shape of the flame, before the lighted candle. (*d*) At the base of the flame is seen a small blue space in which there is complete combustion, and the colour is somewhat similar to that of a Bunsen burner, in which there is also complete combustion.

Other familiar cases of burning.—Candles are not now commonly used for lighting our houses. Sometimes lamps are employed, and in large rooms ordinary coal-gas is often the substance which is burnt. Is the burning of oil and gas like that which occurs when a candle is lighted, and if not, what are the differences ?

By holding a clear dry bottle over the flame of a burning candle it is seen that water is formed as the burning is continued, because the vapour becomes condensed on the inside of the cold glass. In exactly the same way, it is found that water is formed when the flame is due to the burning of either oil or gas. But when a candle is burnt, a gas which turns lime-water milky is formed as well as water. Is this gas also produced when oil and gas are burnt ? If an oil lamp, or a gas jet, or a splinter of wood, is allowed to burn for a few minutes in a glass jar, and is then removed and the jar covered with a glass plate, lime-water poured into the jar is turned milky. We may consequently say that when a candle, a lamp, coal-gas and wood are burnt, two substances are formed, namely, water and an invisible gas which turns lime-water milky.

To prove that nothing is lost when oil burns or a jet of coal-gas is lighted, an experiment must be made similar to that already described with a candle (p. 17). The mass of the oil or gas used up, would have to be determined and also the masses of the water and the other substances produced by the burning. When this is carefully done, and none of the products of combustion are allowed to escape, the total mass of the substances formed is found to be greater than that of the oil or gas burnt. In every case which chemists have examined, the same thing is proved to hold true. In no kind of chemical change is there any loss of matter.

CHIEF POINTS OF CHAPTER XXIII.

Action of metals on air.—Metals increase in weight when they are tarnished by being heated in the air. Magnesium increases in weight when it burns in air. When iron rusts it increases in weight; the increase is due to a gas taken out of the air by the iron as it rusts. The part of the air left behind as iron rusts will not allow a taper to burn in it.

Composition of the air.—Air contains two gases; one is used up by iron in rusting, the other is left behind, and will not allow things to burn in it.

Proportion of the two gases in air.—When an excess of iron rusts in an enclosed amount of air, it always takes up one-fifth of the volume of the air. This part, which is also concerned in burning, may be called the *active part* of air. The remaining four-fifths of the air, which will not allow things to burn in it, in which iron will not rust, may be called the *inactive part* of the air.

The burning of phosphorus in air.—Phosphorus readily burns in the air; in doing so it takes out the active part and combines with it to form a snow-like powder. It can also slowly take out the active part of air without being lighted. When phosphorus is burnt in an enclosed volume of air, one-fifth of this air is used up, and four-fifths remain.

The burning of a candle. When a candle burns in air, water is formed, and also a gas which turns lime-water milky. As in other cases of burning, one-fifth of the air is used up by a candle in burning.

EXERCISES ON CHAPTER XXIII.

1. If the inside of a bottle had iron filings spread over it, and the bottle were corked up and left in a warm room for a few weeks, what would you expect to see if the cork were then withdrawn while the mouth of the bottle were held under water?

2. Explain the changes that would take place in the bottle mentioned in the previous question.

3. Half an ounce of magnesium is burnt in the air, and the ashes formed are carefully collected and weighed. Compare the weight of the ashes with that of the magnesium burnt, and explain the reason for the difference in the weights.

4. Describe an experiment (making a drawing of the apparatus you would use) to show that air contains a gas in which a taper will not burn.

5. A piece of phosphorus is burnt in a stoppered bottle containing 100 c.c. of air. The stopper of the bottle is then taken out under water. How many cubic cms. of water enter the bottle? What facts do you learn from this observation?

6. Compare and contrast the rusting of iron and the slow burning of phosphorus.

7. A piece of phosphorus is placed in a bottle, and the bottle is corked, suspended from one arm of a balance, and counterpoised. The phosphorus is then caused to burn by heating the bottle. Will the balance remain counterpoised after cooling? Give reasons for your answer.

8. A lighted candle is stood upright in water, and a bottle is placed over it, with the neck under the surface of the water. Describe and account for what happens.

9. If two bottles were given you, in one of which phosphorus had been burnt, and in the other a candle, how could you decide which bottle was used for the phosphorus?

10. Write a short essay on "the burning of a candle."

CHAPTER XXIV.

NITROGEN AND OXYGEN AS CONSTITUENTS OF AIR.

90. SEARCH FOR THE ACTIVE PART OF AIR.

i. Changes produced by heating lead in air.—Heat a few pieces of clean lead in an open crucible (Fig. 217). When the lead has melted, stir the liquid metal with a stout iron wire. Notice the formation of a powdery scum upon the lead. Observe that the colour of the powder is darker when hot. Let the crucible cool. Notice that it now contains a yellow powdery substance in addition to the unchanged lead. By strongly heating this powder its colour changes again, and it becomes red lead.

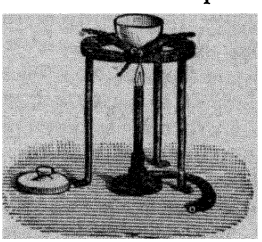


FIG. 217.—Heating lead in contact with air.



FIG. 218.—When red lead is heated, a gas is given off which will re-kindle a glowing splinter.

ii. Gas produced by heating red lead.—Place a little red lead in a hard glass tube, and strongly heat the tube as in Fig. 218. Notice that the red lead undergoes a change of colour. Into the tube insert a glowing splinter. Observe that the splinter is re-kindled. Why is this?

iii. Change produced by heating mercury rust.—Repeat the preceding experiment with some red oxide of mercury, and notice the formation of the silvery, mirror-like

deposit of mercury, or quicksilver, round the cold upper part of the tube. Insert a glowing splinter of wood, and watch it re-kindle.

Where to look for the active part of air.—Since, in suitable circumstances, iron, copper, and lead take the active part out of the air, and unite with it to form fresh substances, it should not be difficult to make these, or similar substances, give up the part of the air which they take up, and so procure the active constituent of air by itself in a pure form. But a little more thought suggests that probably some of these substances would do better than others. It is quite certain that some are formed more easily than others. Will those which are most easily formed be the best from which to get the active part? No. The reason is this. When a chemical change takes place easily, it generally means that the substances taking part in the change have, as it were, a great liking for one another, and when they combine together they form a compound which it is difficult to separate into its parts again. The easiest way to set to work is, therefore, first to find some substance which only combines with the active part slowly and with difficulty, for the compound such a substance forms with the active part will most likely be a weak one, and easily broken up again.

The compounds which lead forms with the active part of the air.—When lead is heated in contact with the air, a yellow powder which is much darker in colour when hot is formed. If the heating is continued long enough, all the metal is changed into powder. The change takes place fairly easily, so that from previous reasoning it may be concluded that it is probably difficult to get the active part of the air again from this powder. This is so. But it is found that, when some of the yellow powder is heated for a long time at the temperature at which lead melts, it slowly takes up still more of the active part of the air, and changes in colour, becoming red. The first powder obtained, which is yellow, is in some states called **litharge**; the second red powder is known as **red lead**. It is easy to get the second lot of the active part of the air again from the red lead. A third powder of a black colour, and having less of the active part of air than either litharge or red lead, is also formed when lead is heated in air.

How the active part of air is obtained from red lead.—When red lead is heated, it changes in colour, and if the heat

has not been great, it regains its original red colour when allowed to cool. But if strongly heated, the red lead gives up some of the active part of the air which it contains, and is reconverted into litharge. The amount of the active part of the air which it thus gives up on being heated is the second quantity referred to in the last paragraph, which is taken up slowly when the heating of lead is continued for a long time. If red lead is strongly heated in a tube, as in Fig. 218, and a glowing splinter of wood is pushed down the tube, the splinter bursts into flame and continues to burn brightly. The active part of air has been obtained alone, and supports burning strongly.

Other ways of obtaining the active part of air.—Quicksilver, or mercury, when strongly heated in the air, slowly combines with the active part, and gradually becomes converted into a bright red powder, which is the rust of mercury. If some of this rust of mercury is heated in a hard glass tube, as in Fig. 218, it soon changes in colour; and as the heating is continued it is noticed that a mirror-like deposit is formed round the top, cold part of the tube. When this deposit is rubbed with a penholder or pencil, it runs together and forms little drops of quicksilver. Moreover, if a glowing splinter of wood is introduced into the tube, it bursts into flame, showing that the active part of the air is being driven out of the red mercury rust. This change is just the reverse of what takes place when mercury itself is heated. The active part of the air, with which hot mercury slowly combines, is driven out of the red mercury rust when that is strongly heated. But red mercury rust is expensive, and it is too costly a plan to heat it to obtain a quantity of the active part of air sufficient for the study of its properties. Several other and cheaper substances easily give up the active part of air when heated.

The active part of the air is called oxygen.—As it will be more convenient in the future to speak of the active part of the air by the name chemists use for it, we may state here that it is always called **oxygen**, but the meaning of this name will be better understood later.

91. PREPARATION OF OXYGEN.

1. Oxygen from potassium chlorate.—Place a little potassium chlorate or chlorate of potash (which is the same thing) in a test-

tube, and heat it as in Fig. 219. Observe that the powder **crackles,** melts, and gives off a gas. Test by a glowing splinter of wood, and see that the gas behaves like oxygen, the active part of the air.

ii. Preparation of a small quantity of oxygen.—Powder some crystals of potassium chlorate, and mix the powder with a little manganese dioxide (sometimes called pyrolusite). Heat some of the mixture in a test-tube, as in the last experiment. Observe by putting in a glowing splinter that oxygen is given off. Notice that in this case there is no melting, and the gas comes off more readily.

iii. Preparation and collection of oxygen.—Into a hard glass tube, closed at one end, fit an india-rubber stopper, with one hole in it, through which a tube, bent as in Fig. 219, is passed. The other end of this tube, called the delivery tube, dips under water in a trough. Mix together some potassium chlorate and manganese dioxide, as in the previous experiment, and place the mixture in the tube. Support the tube and delivery tube as shown in the illustration. Fill several bottles with water, and invert them in the trough. Gently warm the tube, and place one of the bottles of water over the end of the delivery tube. As the oxygen is driven off, it displaces the water and gradually fills the bottle. When the bottle is full of oxygen, cover its mouth with a greased glass plate, and lift it out of the trough. In this way fill five or six jars with oxygen.

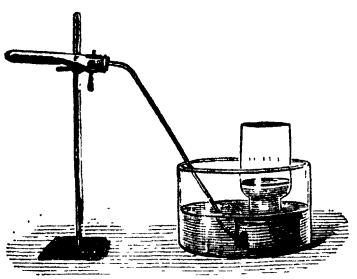


FIG. 219.—Preparation of oxygen by heating a mixture of potassium chlorate and manganese dioxide. The gas is being collected over water.

Caution.—*Be careful not to take away the burner from under the hard glass tube before removing the delivery tube from the trough.*

Preparation of oxygen from potassium chlorate.—As the quantity of oxygen obtained by heating red oxide of mercury is comparatively small, and the oxide of mercury is expensive, a more convenient source of the gas is the white crystalline powder called potassium chlorate.

If this white crystalline compound is heated, in the same way as the red oxide of mercury, it melts and gives off bubbles of oxygen, and after all the oxygen has been given off a white substance like table salt is left behind.

By heating, the potassium chlorate is broken up into two

things, a gas and a white substance like common salt, which is called potassium chloride.

POTASSIUM CHLORATE	gives when heated	POTASSIUM CHLORIDE	and OXYGEN.

Use of oxygen mixture.—By adopting a slightly different method, oxygen can be obtained more readily and easily, for it has been found that by mixing the potassium chlorate with certain other substances, as, for instance, a black compound, manganese dioxide, the oxygen from the chlorate comes off more easily and at a lower temperature. This mixture may be called *oxygen mixture.* If after all the oxygen has been driven out of the mixture the residue left behind is boiled up with water in a flask and the turbid liquid filtered, the manganese dioxide remains unchanged on the filter paper.

92. PROPERTIES OF OXYGEN.

The jars of oxygen prepared as described in Expt. 91 iii. are required.

i. Physical properties of oxygen.—Take one of the bottles of oxygen (one of those collected last should be chosen). Notice everything you can about the contents of the bottle. The gas in the bottle has *no colour.* Remove the plate from the mouth and test its smell; it has *no smell.* Try the taste by breathing some of the gas; it has *no taste.* See if the gas has any effect on moistened litmus papers, one blue and the other red. There should be no effect; we say oxygen is a *neutral* substance.

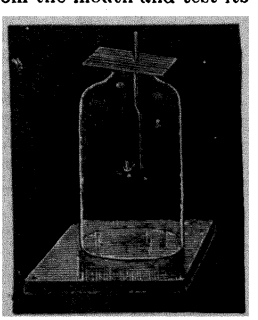


FIG. 220.—Carbon burns in oxygen with a much brighter flame than in air.

ii. The burning of a candle in oxygen.—Attach a piece of stout wire to a wax taper and having lighted the taper plunge it into another of the jars of oxygen. Notice that it is *not extinguished,* but continues to burn, and with a *larger* and *brighter* flame.

iii. The burning of charcoal in oxygen.—Into another jar of oxygen thrust a splint of wood red-hot at the end, or a piece of red-hot charcoal placed in a deflagrating spoon (a small upturned iron spoon with a long handle) (Fig. 220). Note the brilliancy of the combustion. Now pour into the jar some clear lime water (*i.e.* some of the clear liquid which is formed if lime and water be shaken together and allowed to stand. It is really a solution of lime in

water). Notice that it turns milky owing to a white powder being diffused through the liquid.

iv. The burning of phosphorus in oxygen.—In another jar burn a piece of phosphorus about half the size of a pea, contained in the deflagrating spoon. Note the great brilliancy of the combustion and the dense white fumes. Add water and shake—the fumes dissolve. Into the solution put a blue litmus paper. Observe that it is turned *red*.

v. The burning of sulphur in oxygen.—Perform, with another jar, a similar experiment with sulphur. There are few fumes, but a strongly smelling gas is obtained, also soluble in water, turning blue litmus red.

vi. The burning of magnesium ribbon in oxygen.—Ignite a small piece of magnesium ribbon and hold it by means of crucible tongs in a jar of oxygen. Notice the white solid formed. Test its solubility in water, and show that unlike the previous products, it will not turn blue litmus solution red, but will turn red litmus solution blue.

vii. The burning of sodium in oxygen.—Put a small piece of sodium[1] in a dry deflagrating spoon, light the sodium, put it into another jar of oxygen. Observe the fumes formed. Dissolve these in water and try the effect of the solution on litmus paper. It does not turn the blue colour to red, but has exactly the reverse effect, it changes red litmus to blue. Feel the water; it has a soapy feel.

viii. The burning of iron in oxygen.—Obtain a piece of iron wire (a thin steel watch-spring will do), and dip one end into a little melted sulphur, and when the sulphur is burning place the wire in another jar of oxygen. Observe that the sulphur burns and also starts the combustion of the iron, which continues to burn with a brilliant shower of sparks. After the burning has ended, observe that a quantity of an insoluble solid (iron rust) has been formed.

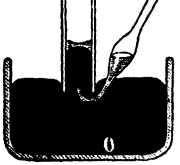


FIG. 221.—How an alkaline solution of pyrogallol can be introduced into a tube containing oxygen.

ix. Oxygen is soluble in an alkaline solution of pyrogallol.—(*a*) Collect some oxygen in a test-tube over mercury. Dissolve a little caustic potash in pyrogallol, so as to make a strong alkaline solution. Force some of the solution into the oxygen by partly filling a bent pipette with it, and blowing down the pipette when the bent end is under the bottom of the test-tube, as shown in Fig. 221.

Note that after a little while the oxygen is wholly absorbed.

(*b*) Repeat the last experiment, substituting air for oxygen, and notice that the air is only partially absorbed.

Properties of oxygen.—Oxygen is a gas which has no colour, no smell, and no taste. It has no action upon litmus paper, and is

[1] Important! Carefully read the note on p. 334.

for this reason said to be neutral. Ordinary combustible substances burn more brightly in oxygen than in the air.

Oxygen has no effect on substances like sulphur and carbon when they are at the same temperature as the room, but if these elements are heated to the point of ignition the oxygen *combines* with them very readily, causing them to burn vigorously.

Some substances which will not burn under ordinary conditions can be made to burn in oxygen, and the case of iron affords a good example of this. Consider what the result would be if there were only oxygen in the air. As soon as iron got red-hot it would start burning. We could not in these circumstances use iron for our grates, furnaces, and similar things.

Oxygen is not very soluble in water—one hundred parts of water dissolve three parts of this gas. That the amount of oxygen dissolved by water is very small is seen by the fact that oxygen prepared for experiment is usually collected over water. But though the amount is small it is of great importance in the economy of nature, for it is due to this dissolved oxygen that water animals are able to breathe. Oxygen can, however, be readily dissolved by some liquids, such as a solution of pyrogallol in caustic potash.

Though oxygen exists in a gaseous condition under ordinary conditions of temperature and pressure, yet it can, by lowering the temperature and very much increasing the pressure, be made to assume the liquid, and even the solid, state.

Oxygen is indispensable to life. It is the constituent of the atmosphere which is used up in the processes of combustion, decay, and fermentation.

Formation of oxides explained.—Whenever oxygen combines with another element an oxide is formed. Indeed, oxygen is so active and powerful that it forms oxides with every simple substance except fluorine.

In all the cases of burning studied experimentally when considering the properties of oxygen, new substances with new properties have been formed; they are therefore chemical compounds, and the experiments afford instances of chemical action. Taking some of the experiments performed as examples, when sulphur burns in oxygen a compound which has a

distressing smell and reddens a blue litmus-paper is formed; it is called Sulphur Dioxide.

SULPHUR burning in OXYGEN forms SULPHUR DIOXIDE.

Similarly, when carbon burns in oxygen, a gas which extinguishes a burning taper and turns lime-water milky is formed. This compound is known as Carbon Dioxide.

CARBON burning in OXYGEN forms CARBON DIOXIDE.

Again, when iron burns in oxygen, a brown powder, which is really ordinary iron rust, is formed, as well as a brittle, black solid quite unlike the original iron. These compounds are both of them oxides of iron.

IRON burning in OXYGEN forms IRON OXIDE.

An examination of these oxides shows that they are identical with the products formed when the substances burn in air.

93. THE INACTIVE PART OF AIR.

i. **The inactive part of air.**—Repeat the experiment of allowing iron to rust in an enclosed quantity of air over water (p. 307), and satisfy yourself that the gas left behind (*a*) extinguishes a flame; (*b*) has no action on a litmus paper; (*c*) does not turn lime-water milky.

ii. **Air can be obtained again by mixing oxygen with the inactive part of air.**—Again allow iron to rust in an enclosed amount of air. When the volume of gas in the bottle has ceased to diminish, remove the muslin bag containing the iron. Place the delivery tube from an oxygen apparatus, similar to that used in the experiment on p. 324, and by heating the hard glass tube bubble oxygen into the bottle until it is again full of gas. Cover the mouth of the bottle with a piece of cardboard, lift out the bottle, and test the gas it contains with a burning taper: it behaves just like air.

Nitrogen.—The gas which is left in a bottle of air after iron has rusted in it, or phosphorus has burnt in it, or in which any one of the instances of burning which have now been studied has occurred, will no longer allow a candle or taper to burn in it. This is one reason why it is called the inactive part of air. The inactive part of air does not affect damp iron at all; that is, the iron does not rust when put into it. The name by which this gas is known to chemists is **nitrogen.**

Properties of nitrogen.—The gas nitrogen is very inert, since it can be made only with difficulty to combine with any other substance. It does not burn, nor will it allow other

substances to burn in it; or, as usually expressed, **it does not support combustion.** A mouse dies if put into this gas.

If the negative nature of the properties of nitrogen is borne in mind, and is considered in connection with the very active powers of oxygen, it will be understood that the presence of the gas in the air serves the purpose of diluting the oxygen, weakening its powers, and making combustion much less intense than it would otherwise be.

If nitrogen obtained from air be heated with either magnesium or lithium, it is found that nearly 1 per cent. of it remains unabsorbed. This residue is another substance present in the atmosphere to the extent indicated, and called **argon.** Argon is also very inert, more so even than nitrogen. Owing to this, even until the year 1894, its presence in the air had been completely overlooked, although, nearly a century before, the eminent chemist Cavendish had unknowingly obtained some, regarding it as an impurity which he had overlooked—an example of the importance of giving attention to the minutest details in scientific investigations. After the discovery of argon, four other new gases—helium, neon, krypton, and xenon—were found to be permanent constituents of air, but they occur in such minute quantities that their presence need only be mentioned here.

Chemical composition of air.—What may be termed the fundamental gases in air are oxygen and nitrogen. Argon and the four other new gases may for our purposes be regarded as part of the nitrogen. Carbon dioxide and water vapour are practically always present; and various other gases, or vapours frequently occur in small quantities, but these may be regarded as impurities, and ought not to be considered as constituents of pure air. The following table shows the percentage composition of air as regards volume; that is, for instance, the number of cubic feet of the various gases present in 100 cubic feet of the atmosphere:

Oxygen, a gas which supports combustion,	21·00
Nitrogen, an inert gas,	78·03
Argon, an inert gas,	0·94
Carbon dioxide, a suffocating gas,	0·03
Water vapour,	Variable
Nitric acid, Ammonia, Ozone.	traces

These proportions are remarkably constant in ordinary air, and it is only in localities or under conditions of an exceptional character that they vary to any noteworthy degree. In the air of mines, the oxygen has been found as low as 18·6 per cent., but this represents almost the lowest percentage of oxygen ever obtained from a place where human beings could live. In the midst of vegetation, or open ground, especially in the daytime, oxygen is present in the proportion of about 21 per cent., but never more.

Nitrogen not only serves to tone down the activity of oxygen as a supporter of combustion, but it is also useful to the life of plants. A few lowly plants appear to absorb nitrogen direct from the air, but the majority of them obtain it indirectly as the result of the action of bacteria existing in the soil or in their roots.

The proportion of carbon dioxide rarely exceeds 3 parts in 10,000 in pure air, and is not often less than 2·7 parts per 10,000. During the night the proportion is slightly greater than in the day. In the streets of a town the amount of carbon dioxide only exceeds the average amount of the open country by about 1 part in 10,000. In rooms, however, and badly-ventilated places, carbon dioxide is often greatly in excess, and oxygen is present in a much smaller proportion than it ought to be. Carbon dioxide is not essentially a poisonous gas, but it is often found in bad company, and when it occurs in excess the air of which it forms a part is unfit to breathe.

Air always contains a certain proportion of invisible water vapour, and when the air is cooled to a sufficient degree this vapour becomes visible in the form of mist, fog, cloud, rain, or other familiar forms of water (p. 193). **Ozone** is a peculiar form of oxygen, and is usually present in the air of the open country or over the sea, but not in that of towns.

In addition to these and other gases, numerous minute solid particles are suspended in air, some of them being living germs. They are more abundant in the town than in the country.

The air is not a chemical compound, but a mixture of gases.—In proof of this statement, the following facts and experiments may be cited:

1. The composition of a chemical compound never varies, while the composition of air does vary slightly.

2. Whenever a chemical compound is formed, a certain amount of heat is developed, and there is generally a change of volume. When, however, oxygen and nitrogen are mixed in the proportion in which they occur in the air, there is no evolution of heat or change of bulk, though the mixture cannot be distinguished from air.

3. The proportion in which oxygen and nitrogen are mixed in the air does not bear any simple relation to their combining weights (p. 383), whereas in the case of every true chemical compound, the amounts of the constituents always bear some simple ratio to these weights.

4. When air is shaken up with water, some of it is dissolved. If air were a chemical compound it would be dissolved *as a whole*, and therefore the dissolved part would have the same composition as the undissolved part. But this is not found to be the case. The air dissolved in water can be expelled by heat, and if it is collected the oxygen in it is found to be more in proportion to the nitrogen than it is in ordinary air, thus showing that water dissolves more oxygen than nitrogen.

Ordinary air consists of about one-fifth oxygen and four-fifths nitrogen by volume; but air expelled from water contains about one-third oxygen and two-thirds nitrogen.

5. When air is liquefied by intense cold and great pressure, and the liquid air is then permitted to evaporate, the nitrogen is first given off, so that the liquid becomes richer and richer in oxygen. If air were a compound, no one part of it would be more volatile than the other.

CHIEF POINTS OF CHAPTER XXIV.

Oxygen.—Oxygen is the active part of the air. When red oxide of mercury is heated it decomposes into mercury and oxygen. Oxygen can also be obtained by heating potassium chlorate, when potassium chloride, a substance very like common salt, is left behind.

Oxides.—When some simple substances are heated in oxygen they unite with it, forming *oxides*. Thus:

Iron and oxygen form an oxide of iron.
Phosphorus and oxygen form an oxide of phosphorus
Carbon and oxygen form an oxide of carbon.
Sodium and oxygen form an oxide of sodium.

Some oxides unite with water to form *acids*, which turn blue litmus red; the oxide of phosphorus is an example.

Other oxides unite with water to form *alkaline solutions*, which have a soapy feel and the power of turning reddened litmus blue again; the oxide of sodium is an example.

Nitrogen.—Nitrogen is the inactive part of the air. It makes up 80 per cent. (roughly) of the volume of the atmosphere. It is very inert. It neither burns nor allows things to burn in it. It is left behind when substances burn in air.

Air.—The air is a *mixture* of gases and not a chemical compound. Its chief constituents are nitrogen, oxygen, argon, carbon dioxide, and water vapour.

EXERCISES ON CHAPTER XXIV.

1. What are the chief properties of oxygen?
Describe the experiments you would make to illustrate these properties.

2. Oxygen and nitrogen are the two chief ingredients in common air. State reasons for the usual belief that they are not combined together chemically.

3. Describe experiments which prove that air is composed of at least two different gases.

4. How would you show by experiments that only one part of the air is concerned in (*a*) burning, (*b*) rusting?

5. State approximately what the difference in composition is between ordinary air and "dissolved air" which has been expelled by boiling from solution in water.

6. Describe and explain the various methods with which you are acquainted by means of which air can be deprived of its oxygen.

7. What is left behind as a gas when phosphorus burns in a closed vessel? What properties has it? Has it been made to combine with oxygen; and, if so, how?

8. Lead forms three well-known oxides. Describe the appearance of each substance. How does the proportion of oxygen differ in the three kinds?

CHAPTER XXV.

WATER AND HYDROGEN.

94. ACTION OF METALS ON WATER.

i. Iron becomes rusty in water containing dissolved air.—Leave a piece of clean iron in water for a few days. Notice that rust forms upon it.

ii. Iron in boiled water does not rust.—Boil some water in a flask or clean saucepan for some time so as to drive out the dissolved air. Obtain a wire-nail, or a steel pen-nib, and after washing it in soap and water to remove oil, dry it, place it in a bottle, and quickly fill up the bottle with the boiled water. Insert a tightly-fitting india-

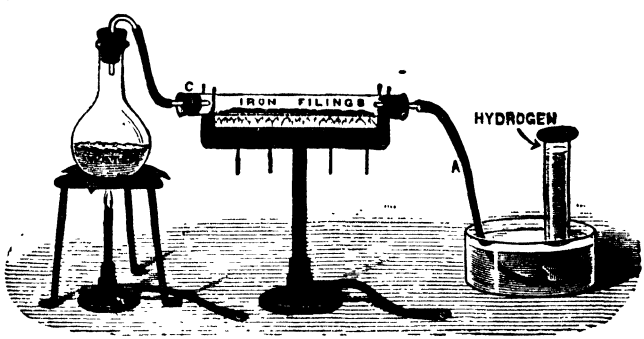


FIG. 222.—A means of obtaining hydrogen gas from water by depriving steam of oxygen.

rubber stopper, being careful to enclose no air between the stopper and the top of the water. Place on one side, and examine the iron after a few days. No rust will be seen.

iii. Action of heated iron on water.—Place some iron filings in a hard glass tube, having at each end a cork through which a short glass tube passes (Fig. 222). To the end *C* fit a delivery tube from a flask containing water previously boiled to drive off the dissolved air, and to the other end fit a short piece of rubber tubing. Heat

the iron filings well, and boil the water in the flask so that steam passes over the heated iron, and then into the water, where it condenses. Now place over the end of the tube *A* an inverted jar of water, and note that the steam is not completely condensed, but that minute bubbles ascend to the top of the test-tube. When you have so obtained a sufficient quantity of gas (a few cubic centimetres), first disconnect *C* from the flask of water, then stop the boiling. Close the end of the jar with a glass plate. Hold a lighted match to the mouth of the jar and take away the glass plate. Observe that the gas *burns*. Examine the iron filings in the hard glass tube, and see that a quantity of rust has been formed.

iv. Action of sodium on water.—Place a small piece of sodium [1] in water in an evaporating basin, and quickly put a large glass shade over the latter; observe the action. Feel the water left after the sodium has all disappeared, and test it with red litmus. Evaporate away the water. Note the residue.

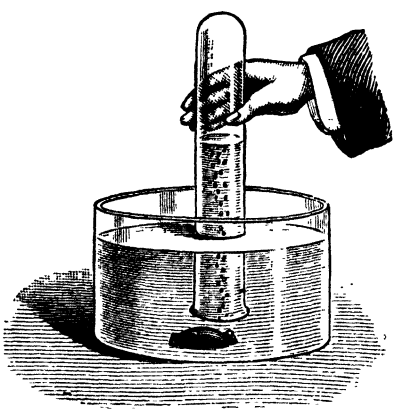


FIG. 223.—Sodium combines with the oxygen of water and turns out the hydrogen.

v. Collection of the gas which sodium turns out of water.—Place a small piece of sodium in a small piece of lead tubing, the ends of which are nearly closed, and gently drop the lead into a pan of water. A gas is seen to come off (Fig. 223). Collect this in an inverted tube full of water, and by this means obtain three test-tubes of the gas. Observe that the gas is colourless and odourless.

vi. Examination of the gas produced by the action of sodium upon water.—Take out two of the tubes and hold them for the same time, say 30 seconds, (1) with mouth up, (2) with mouth down. Then apply a light to the mouth of both.

Now try the 3rd tube, holding it mouth down, and place a lighted match up into the tube. Note that the gas does *not* explode, but burns quietly, while the match is extinguished.

Chemical examination of water.—Before proceeding to study the chemical behaviour of water it will be well to

[1] Great care must be taken when using sodium, which must never be allowed to touch damp materials. It is kept under naphtha until used, and should never be handled with the fingers. It should be dried by blotting-paper when taken from the bottle and cut with a clean knife, the pieces not used being immediately replaced in the bottle.

recapitulate what has been already learnt about it. Water is a clear liquid with a blue-green colour, best seen by causing light to pass through a considerable length of it. It boils at 100° C., and is then converted into steam. It freezes at 0° C., becoming ice. Its density is 1, *i.e.* the weight of 1 c.c. is 1 gram at 4° C. and slightly less at any other temperature, owing to the fact that water expands when either cooled below or heated above 4° C. It has the power of dissolving many substances—*e.g.* salt, sugar, etc.—forming solutions from which the water may be evaporated away, leaving the solid behind. But these facts tell nothing of the chemical nature of water. By suitable experiments, however, the mutual action between water and some other substances may result in entirely different products. It has been seen that when iron is left in water it forms a considerable quantity of *rust*. But this may be due to the air which may be dissolved in water, for if the iron be placed in a tube containing water, which has first been well boiled to drive off the air, and the tube is then sealed in a blow-pipe, the iron either does not rust or does so to only a very slight extent.

If, however, the iron is heated and water is passed over it in the form of steam, in the manner described in Experiment 94 iii., a chemical action begins which teaches several important facts about the composition of water. Not only do the iron filings become rusty, just as they do when exposed to damp air, but a gas, insoluble in water, which can be collected over a trough, as shown in Fig. 222, is obtained. This gas burns when a lighted taper is brought near it.

Action of sodium upon water.—It has been seen that from steam and iron it is possible to obtain iron rust, that is, iron oxide, and an inflammable gas. This fact suggests that water contains this inflammable gas as well as oxygen, and a way to test this is to find something which has a powerful chemical attraction for oxygen, for, if this substance is placed in contact with water, it will take up the oxygen and leave the other constituent or constituents of water. Such a substance is the metal sodium.

When a small piece of sodium is thrown upon water it swims about on the surface with a hissing noise; and the solution, after the sodium has all disappeared, has a soapy feel and turns red litmus blue. In this case it can be seen only that the action

is energetic and a new product is formed, while the soapy feel of the water and its action on litmus appear to indicate that this product is the same as that obtained when sodium is burnt in oxygen and the fumes formed are dissolved in water. If the pellet of sodium is enclosed in a piece of lead piping, the ends of which are flattened, the gas comes off in such a way that it can be collected with ease, and is then observed to be colourless and to have no smell. If two tubes, which have been filled with the gas (Fig. 223) and held for thirty seconds, one with its mouth upwards and the other with its mouth downwards, have a lighted match applied to them, it is found that there is no effect with the first tube and a slight explosion with the second.

The slight explosion of (2) shows it to contain an explosive gas, while the absence of any effect with (1) shows that the gas has all disappeared. Hence, it is seen that the gas escapes from a tube held mouth upwards, but not so quickly from one held mouth downwards. It is therefore lighter than air, being in fact the lightest gas known.

If a third tube of the gas is examined in the same way, immediately on taking it out of the water, the gas does not explode, but burns quietly.

95. PREPARATION AND PROPERTIES OF HYDROGEN.

1. **Preparation of hydrogen.**—Select a flask and fit it up as is shown in Fig. 224. Be very careful that the stopper and the tubes respectively fit very closely. Into the flask put enough granulated zinc to cover the bottom. Pour some water upon the zinc. Arrange the delivery tube in the trough as when making oxygen. Pour a little sulphuric acid down the thistle-headed acid funnel, and be quite sure that the end of the funnel dips beneath the liquid in the flask. Do not collect bottles of the gas until you are sure pure hydrogen is being given off, which you can find out in this way. Fill a test-tube with water and invert it over the end of the delivery tube. When it is full of gas, still holding it upside down, take it to a flame (which should not be near the flask in use); notice that there is a slight explosion. Continue this until the hydrogen burns quietly down the test-tube. When this happens proceed to fill one or two bottles. When the bottles have been filled, it is better not to remove them from the water until you want to use them. Collect also a soda-water bottle half full of the gas.

Caution.—*Be careful not to bring a light near the thistle funnel or tube delivering the gas, even when the action in the flask seems to have ceased, or a dangerous explosion may occur.*

Be careful also that none of the acid used gets upon your fingers or clothing.

ii. The liquid left in the flask.—Filter off the liquid in the flask from the undissolved zinc (sufficient zinc should be used to leave a

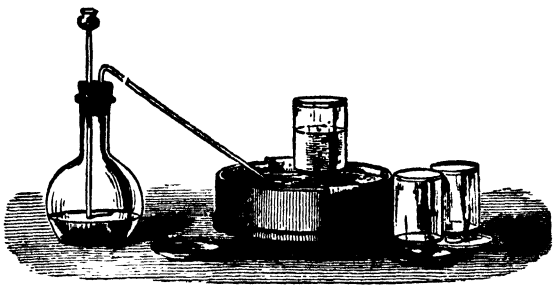


FIG. 224.—Zinc turns hydrogen out of dilute sulphuric acid.

quantity still undissolved: if all has disappeared add more and wait till the action ceases). Partially evaporate the liquid and allow it to crystallise. A quantity of clear colourless crystals are formed. Examine them and sketch the most perfect. Heat some of the crystals in a tube and observe that they melt, give off water (which can be collected and proved to be water), and leave a white powder.

iii. Hydrogen burns but extinguishes a flame.—Test one jar of the gas by means of a lighted match or taper. Observe that the gas burns at the mouth of the jar and that the taper is extinguished when thrust into the tube; on being taken out, the taper again becomes alight on passing through the flame of the burning hydrogen (Fig. 225).

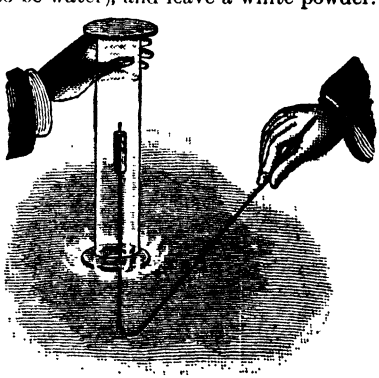


FIG. 225.—The hydrogen burns at the mouth of the jar, but the candle is extinguished when inside the jar of hydrogen.

iv. Hydrogen is lighter than air.—Take a full jar of the gas and hold it mouth upwards below a second smaller jar held mouth downwards, as shown in Fig. 226. On testing with a lighted taper observe that the gas has left the lower jar and filled the upper. Many experiments, as the filling of balloons or soap bubbles, may also be performed to demonstrate the extremely low density of hydrogen.

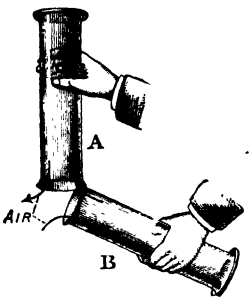


FIG. 226.—Hydrogen is lighter than air and can be poured upwards.

v. Hydrogen forms an explosive mixture with air.—Wrap your hand in a duster and with it hold the soda-water bottle (Expt. 95 i.). Take it out of the water so that the water runs out, and the bottle is now filled with a mixture of hydrogen and air. Apply a light and you will not fail to observe that an explosion results.

vi. The flame of burning hydrogen.—Fit a right-angled tube drawn out to a point to a hydrogen generator, as shown in Fig. 227. After pouring a little sulphuric acid down the thistle funnel, collect a test-tube of the gas issuing from the straight tube, and hold the mouth of the test-tube near a flame, *which must be a few feet away from the generator.* The gas will at first go off with a pop, or burn with a squeaking noise, but after two or three trials it will burn quietly, with a blue flame. *When you can carry this flame of burning hydrogen to the apparatus from which the gas is being produced, do so, and use it to ignite the gas escaping from the pointed tube.* If you remember always to do this, there can be no danger, for when you are able to carry a flame of hydrogen in a test-tube of the gas, for a distance of two or three feet to the generating apparatus, you may be sure that the hydrogen issuing from the apparatus is not mixed with air.

Light the hydrogen in this way.

Observe that it burns with a pale blue flame, which after a time becomes yellow. This coloration is due to the glass becoming hot, and some of the substances in it being burnt in the flame of hydrogen.

vii. Hydrogen can be collected by upward displacement.—Substitute for the delivery tube in Expt. 95 i. a tube bent twice at right angles, as shown in Fig. 228, and arrange a jar on a retort-stand in an inverted position. Place a test-tube over the upright tube, and allow it to stay there for a minute, and test the gas as described in Expt. 95 i. When it is pure, substitute the inverted jar, and after a few minutes lift it off the stand and apply a light, first taking the precaution to wrap a duster round the jar and to hold it away from your face. The jar will be found to contain hydrogen.

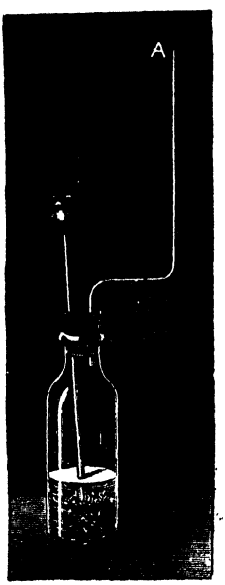


FIG. 227.—Arrangement for obtaining a flame of hydrogen.

Preparation of hydrogen in large quantities.—It has been seen that from water and sodium it is possible to obtain an inflammable gas, lighter than air, which does not support combustion; and a solution which behaves like the solution of sodium oxide is also formed. The most natural inference is that the water contains this inflammable gas, which is called **hydrogen**, combined with oxygen. Before proceeding to verify this, it will be well to examine more carefully the properties of the inflammable gas, and to do this it is necessary to collect the gas in greater quantity than hitherto. It is prepared by acting upon dilute sulphuric or hydrochloric acid with zinc or iron, the former metal being generally used.

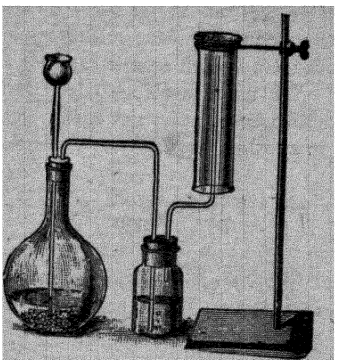


FIG. 228.—Hydrogen being lighter than air can be collected by upward displacement.

The apparatus suitable for the preparation of hydrogen in this way is described in the experimental work (p. 336). Owing



FIG. 229.—Crystals of zinc sulphate.

to the slight solubility of hydrogen in water, it can be collected in the same way as oxygen over the pneumatic trough. If, when the chemical action in the flask has completely stopped, the liquid is filtered from the still undissolved zinc, as previously explained, and then partially evaporated in a basin and after-

wards allowed to crystallise, a quantity of clear colourless crystals is formed. These crystals melt if heated in a tube, give off water, and leave a white powder. They consist of a compound formed from the zinc and part of the sulphuric acid and are known as **zinc sulphate** (Fig. 229).

We may therefore state that *sulphuric acid and zinc form hydrogen and zinc sulphate.* Or, the same fact may be expressed in another way:

Sulphuric Acid	when acted upon with	Zinc gives	Zinc Sulphate	and	Hydrogen.

Water of crystallisation.—Many crystals behave like zinc sulphate on being heated, *i.e.* they lose water which they previously contained, and become converted into a powder. The water contained in a crystal and evolved on heating is known as **water of crystallisation.** Some substances, like blue vitriol (copper sulphate), change in colour when their water of crystallisation is driven out by heat; but the colour can be regained by adding water.

Properties of hydrogen.—Having now a means of obtaining hydrogen in considerable quantity, it can be observed that it is a colourless, odourless gas, considerably lighter than air; it burns, but does not support combustion, and it forms a highly explosive mixture when mixed with air. It is now necessary to obtain and examine the compound which is produced by the burning of hydrogen—that is, the oxide of hydrogen.

96. WHEN HYDROGEN BURNS WATER IS FORMED.

1. Water is formed by burning hydrogen.—(*a*) Arrange a flask as before for the production of hydrogen. Pass the gas through a tube containing chloride of calcium in order to dry it thoroughly. Allow it to burn under a retort which is kept cool by a stream of water flowing in at the tubule and out at the end of the neck (Fig. 230). Observe the formation, on the outside of the retort, of a clear liquid which collects and drops into a beaker placed to receive it. By this means sufficient of the liquid can be obtained to identify it, especially if several students add together the liquids formed in their experiments.

(*b*) Take the density, freezing point (a mixture of sodium sulphate and hydrochloric acid forms a convenient freezing mixture), and

boiling point of the liquid formed by burning hydrogen. You will

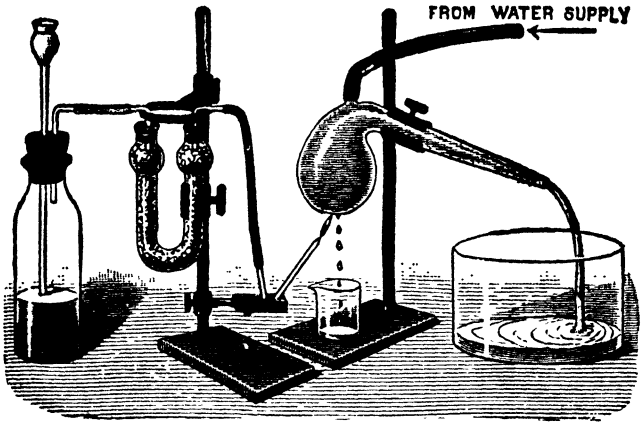


FIG. 230.—The water formed when hydrogen burns in the air can be collected and examined.

find these are 1, 0° C., and 100° C. respectively, and these results are sufficient to enable it to be stated that the liquid is identical with pure water.

ii. Analysis of water. —This may be done by means of an electric battery for generating the electric current, and a *voltameter*. The latter is most simply made by closing the bottom of a funnel by means of a tightly-fitting cork through which pass two platinum wires with small plates of platinum attached to the ends remaining in the funnel (Fig. 231). Over these plates are supported two glass test-tubes of equal capacity, and the tubes and part of the funnel are filled with water to which has been added a little sulphuric acid, as otherwise the liquid offers great resistance to the passage

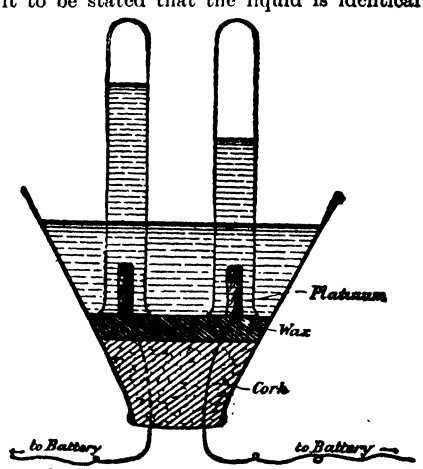


FIG. 231.—A voltameter in which water can be analysed by the electric current.

of the electric current. The wires from a battery of three or four cells are connected with the ends of the platinum wires, and as soon as the connection with the battery is complete, provided there is clean metal at every junction, bubbles of gas are seen to rise from each platinum plate, and to ascend into the tube and displace the contained liquid. After the experiment has gone on for half an hour, the gases may be tested and their volumes measured. It will be found that the volume of one gas is double that of the other, and that the gas of which there is the larger amount is *hydrogen*, whilst the other is *oxygen*.

Formation of water by burning hydrogen.—When a jet of burning hydrogen is brought into contact with a cold surface, such as a cold glass, the product of combustion, the oxide of hydrogen, is condensed. If after a sufficient quantity of the liquid has been collected it is examined, it is found (*a*) to have a density of 1, (*b*) to freeze at 0° C., and (*c*) to boil at 100° C. These are the physical characteristics of water and of no other substance, so we are justified in stating that this liquid, formed when hydrogen burns, is water.

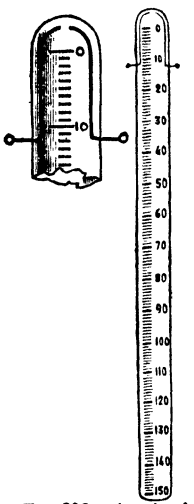


FIG. 232.—A simple form of eudiometer. The arrangement of the platinum wires is shown in the enlarged top of the eudiometer.

Previous experiments have indicated that water contains hydrogen and oxygen, so that it can now be said that: **Hydrogen in burning produces water, which is, therefore, an oxide of hydrogen.**

Proportions of oxygen and hydrogen in water.—The proportions in which the oxygen and hydrogen combine during the formation of water must now be considered. This may be done in either of two ways—viz. by finding the weights of the gases, or by finding the volume of the gases, which combine. For the latter purpose it is necessary to measure out definite volumes of oxygen and hydrogen, cause them to combine, then measure the volume of gas which remains uncombined and ascertain which gas it is. This is usually done in a piece of apparatus known as a **eudiometer** (Fig. 232). In its simplest form this consists of a long glass tube closed at one end and graduated in equal volumes, usually cubic centimetres, by divisions marked on the

glass. Through opposite sides of the tube, at the closed end, pieces of platinum wire are passed and fused into the glass, being so arranged that they do not quite touch one another. Outside the tube the platinum wires are bent into loops to which wires from an electric coil may be attached so that an electric spark can be passed between the ends of the platinum wires inside the tube to explode the mixed gases. The effect is not essentially electrical. The spark simply served the purpose of raising the temperature of the mixture to the point of ignition. Using an instrument of this kind, it can be proved that **2 volumes of hydrogen combine with 1 volume of oxygen to form water.**

Such a process as this—the formation, that is, of a compound from simpler materials—is known as a **synthesis.**

The same fact may also be found by the analysis of water—that is, by breaking the liquid up into its components, which can be done by means of an electric current. The apparatus shown in Fig. 231, and called a **voltameter**, may be used. Water slightly acidulated to render it a conductor of electricity has an electric current passed through it between electrodes of platinum. The electric current tears apart the constituents of water and liberates them separately, one from each electrode. At the electrode where the current enters the liquid, called the **anode**, oxygen is set free, while hydrogen bubbles from the **kathode.** If the gases are collected in two test-tubes inverted over the electrodes, it will soon be obvious that the decomposed water yields two volumes of hydrogen to each volume of oxygen. Such decomposition or breaking up of chemical compounds by means of an electric current is called **electrolysis.** This method of analysis is applicable to many other chemical compounds as well as water. Electrolysis is not only of scientific interest, but of great commercial and industrial importance.

97. COMPOSITION OF WATER BY WEIGHT.

1. Action of hydrogen on heated copper oxide.—Arrange an apparatus like that shown in Fig. 233, in which a flask *A* for the making of hydrogen is connected with the bottle *B* containing strong sulphuric acid. The passage of the hydrogen through the strong

acid completely dries the gas. A small amount of the black oxide of copper is placed in the hard glass tube *AC*, which is about 1·5 centimetres in diameter, and fitted with the tube *D* in the manner shown. When you are sure that pure dry hydrogen is escaping from the open end of *D*, heat the oxide of copper in the tube by means of the burner *E*. In a few minutes moisture will be seen to collect in the tube *D*, and presently to drop into the dish put to collect it.

Examine the residue in the hard glass tube, and note its change to a reddish colour; this is due to the presence of copper.

Rearrange the apparatus, using the U-tube *G* in the place of the tube *D*. The U-tube *G* containing lumps of calcium chloride is employed to collect the water formed.

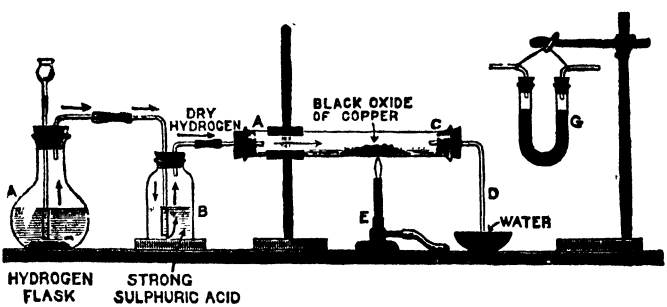


FIG. 233.—When hydrogen is passed over heated oxide of copper it extracts the oxygen, with which it forms water, and leaves copper behind.

ii. Composition of water.—Put some copper oxide into the tube *AC*, and carefully weigh the tube. Similarly the weight of the U-tube *G* and its contents must also be determined. As in the last experiment, when you are sure pure dry hydrogen is escaping from the open end of *G*, heat the copper oxide by means of the laboratory burner. Be sure that all the water formed is collected by the U-tube. If any condenses at the end of the hard glass tube *AC*, drive it over by heating the tube at this place.

Allow the tube *AC* to cool. Disconnect it at *A* and *C*, and again determine its weight and that of its contents. Notice the diminution in weight. Also weigh the U-tube again, and observe its increase in weight.

Relative weights of the constituents of water.—To find the composition of water by weight—that is, the weights of oxygen and hydrogen which combine to form water—it should be noticed that we only require the weights of two out of the

three substances concerned, *i.e.* if we know the weights of hydrogen and water (or of oxygen and water), the weight of the oxygen (or hydrogen) is readily calculated. The experiment is done by finding the weights of the oxygen and water, and for this it is best to use, not oxygen itself, but some oxide which readily gives up its oxygen to the hydrogen, so that by weighing the oxide before and after the experiment we can ascertain the weight of oxygen which it has lost. The oxide used for this purpose is usually oxide of copper, a black powder, which it has been seen is obtained when air is passed over red hot copper. Pure dry hydrogen is passed over the heated oxide, and it combines with the oxygen of the oxide to form water and leaves the copper behind.

By weighing the water produced and subtracting from it the weight of the oxygen used, the weight of the hydrogen can be found. If the experiment is carefully performed it is found that **water is formed of eight-ninths its weight of oxygen with one-ninth its weight of hydrogen.**

This experiment, also taken in conjunction with other experiments upon the volumes of the gases, proves further that any volume of oxygen is 16 times heavier than the same volume of hydrogen.

98. HARD AND SOFT WATERS.

i. Solution of chalk in water.—Pass the gas obtained from an apparatus like that shown in Fig. 234 containing pieces of marble on which hydrochloric acid has been poured, into clear lime-water until the solution is clear again, after first becoming milky. Boil a part of the clear solution in a test-tube. Observe the return of the milkiness. What is the cause?

Boil in a flask a little natural water which is hard because of the presence of chalk; observe that the water becomes milky.

ii. Soap solution to test the hardness of water.—Dissolve some Castile soap in spirits of wine and label the solution "soap solution." Add a few drops of soap solution to distilled water. Shake the mixture. Notice a lather is easily formed.

iii. Temporary and permanent hardness of water.—(*a*) Add soap solution to some of the clear solution obtained by passing carbon dioxide for some time through lime-water (Expt. 98 i.). Shake the mixture, and observe that a lather is formed only after a considerable quantity of soap solution has been added.

(*b*) Similarly show, that it is difficult to make a lather with a solution of calcium chloride.

(c) Boil portions of the solution in the last two experiments before adding the soap solution. Soap lathers in the first solution quite easily after it has been boiled: boiling makes no difference to the second solution.

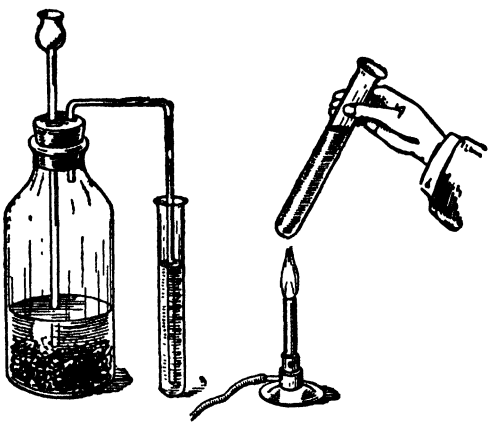


FIG. 234.—When carbon dioxide is passed into clear lime-water, the water turns milky, and afterwards clears. The milkiness is again obtained by boiling the lime-water.

iv. **Hardness of sea-water.**—Try whether the soap solution forms a lather when shaken up with sea-water. Does boiling make any difference?

Natural waters.—The great solvent power which water possesses has been studied already. It is in consequence of this that perfectly pure water does not occur naturally. At the moment of its formation rain is pure water, yet no sooner is it formed than it begins to dissolve various substances. In its passage through the air it takes up varying amounts of the gases which make up the atmosphere, such as oxygen and the gas called carbon dioxide. When the surface of the earth is reached, the water dissolves portions of all the soluble ingredients of the soil and the underlying rocks. The most soluble bodies are, of course, dissolved to the greatest extent. The solvent power of water is considerably increased by the presence of the carbon dioxide it dissolves, partly from the air and partly from the soil.

The sea, from long contact with the earth and the soluble substances in it, is charged with dissolved matter, so that its water is salt. The fresh waters of rivers and springs have only a small proportion of the substances which give the strong taste to sea water.

The substances dissolved are chiefly sulphates, carbonates, or chlorides of calcium, sodium, and magnesium, the quantity present varying from 0·05 grm. to 3 grm. per litre. River water always contains less dissolved matter than spring or well water. We can understand why this is so by remembering that rivers are chiefly fed by surface water, that is, water which has not been in contact with the mineral substances through which spring water passes on its way to the surface.

The impurities in river water may be of two kinds, **suspended** and **dissolved.** The former, which consist of insoluble substances and cause the muddy appearance of rivers after rain, may be removed by filtration. Dissolved impurities are separated from the water when it is evaporated. If the escaping vapour be condensed, pure water is obtained.

The chief substance in sea water is common salt, and in some countries, where this compound is not found in mines, as it is in England, the sea water is evaporated and the crystals of salt as they form are removed and dried. The liquid remaining contains magnesium chloride. In England it is the custom sometimes to flood the salt mines with water, allow it to remain till saturated with salt, and then to pump the salt solution, or brine, to the surface and evaporate it in large flat iron vessels called **salt pans.** The salt is thus obtained either as a fine white crystalline powder or in large crystals according to the rate at which the evaporation proceeds and to the time during which the crystals are allowed to remain in the liquid after they are formed.

When a natural water, other than sea water, is strongly charged with dissolved substances it is called a **mineral water.** Waters which contain a compound of sulphur and hydrogen, called sulphuretted hydrogen, are spoken of as **sulphur waters**; if some compound of iron is the substance which has been taken up in large quantities, we have **chalybeate waters** formed. **Effervescent waters** contain a great amount of dissolved carbon dioxide.

Hard and soft waters.—It is a fact familiar to every one that soap lathers very easily in some waters and not at all

in others. If rain-water be used, the lathering takes place with great ease, while with the water which is supplied to some towns a lather can only be made with difficulty; and if we attempt the same process in sea-water there is no lathering at all. **Those waters in which soap lathers easily are said to be soft. When this is not the case the water is spoken of as hard.** The explanation is a simple one. Water dissolves materials out of the rocks below the soil, and often takes up, among other things, compounds of calcium and magnesium, which unite with soap, forming a new compound of an insoluble kind; and, in consequence, there is no lathering until all the calcium and magnesium have thus combined with soap, after which the solution and lathering of the soap begins. The soap which combines with the dissolved materials is, of course, wasted.

Temporary and permanent hardness—Hard waters differ among themselves. Some can be softened by mere boiling, and when this is the case the hardness is said to be **temporary.** If the hardness is not removed after the water has been boiled, and the water requires the addition of a chemical to soften it, such hardness is termed **permanent.** As has already been mentioned, the presence of carbon dioxide in water gives it the power of dissolving substances which would be otherwise insoluble in it. Chalk, known to chemists as calcium carbonate, is insoluble in pure water, but in water in which there is carbon dioxide it dissolves to a considerable extent. As soon as this dissolved gas is got rid of, which can be done by boiling, the chalk, being no longer soluble, is thrown down upon the sides of the vessel. It forms in this way the incrustation which is found on the insides of kettles and boilers (Fig. 235).



FIG. 235.—The deposit, or fur in a kettle consists chiefly of chalk.

Permanent hardness is due to the presence of dissolved calcium sulphate and other compounds. Since these substances are soluble in pure water, mere boiling will not get rid of them. Washing soda, which is a form of sodium carbonate, softens such water as this by causing the formation of calcium carbonate in the place of the calcium sulphate.

Distillation of water.—If the steam which is formed by boiling water containing any dissolved substances be con-

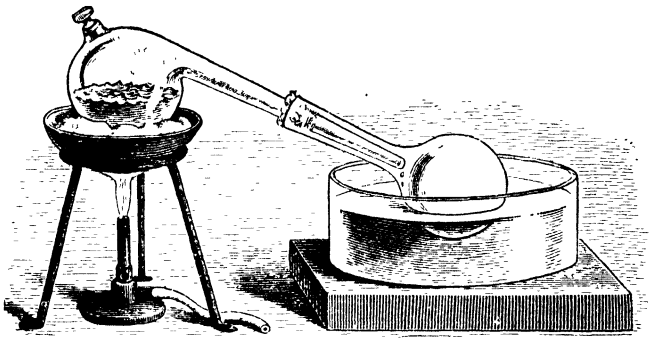


FIG. 236.—A simple arrangement for distilling water.

densed, the water formed is quite pure. To obtain pure water from any kind of water, then, whether fresh or salt, all that has to be done is to boil it and condense the steam which is given off. The dissolved materials are all left behind in the vessel in which the boiling takes place. An arrangement for condensing steam or vapour is shown in Fig. 226. The steam driven off from the water in the retort passes into a flask kept cool by resting upon a basin of cold water, and is thus condensed.

CHIEF POINTS OF CHAPTER XXV.

Water is a clear liquid with a blue-green colour. It boils at 100° C., when it is converted into steam. It freezes at 0° C., becoming ice. Its density at 4° C. is 1. It possesses great solvent power.

Preparation of hydrogen.—Hydrogen is best prepared by acting upon a dilute acid with a metal. Sulphuric acid and zinc have been found to be suitable.

Properties of hydrogen.—It is a colourless, odourless gas, considerably lighter than air. It does not support combustion but itself burns in air. It forms an explosive mixture when mixed with air or oxygen.

Production of oxide of hydrogen.—By collecting the product of combustion when hydrogen is burnt, and examining it, it is found to be a clear liquid with density 1, boiling point 100° C., freezing point 0° C. The product, or oxide of hydrogen, is thus seen to be really water. *Hydrogen in burning produces water, which is, therefore, an oxide of hydrogen.*

Composition of water by volume.—This can be determined by means of a *eudiometer. Two volumes of hydrogen combine with one volume of oxygen to form water.*

The analysis of water is effected by passing an electric current through water contained in a *voltameter.*

Composition of water by weight.—Pure dry hydrogen is passed over heated oxide of copper. The hydrogen combines with the oxygen of the oxide, forming water, and leaves the metallic copper behind. The water formed is collected and weighed. The copper oxide is weighed before and after the experiment. Its loss in weight shows the amount of oxygen in the water formed. The difference between the weight of the water formed and the oxygen it contains tells us the weight of the hydrogen in the water. When carefully performed, the experiment shows that *water is formed of eight-ninths its weight of oxygen and one-ninth its weight of hydrogen.*

Natural waters generally contain dissolved materials. When the amount is very large the water is called a *mineral water.* Natural waters containing sulphuretted hydrogen are called *sulphur* waters; those containing compounds of iron are called *chalybeate* waters; and those containing much carbon dioxide are spoken of as *effervescent.*

Hard and soft waters.—Those waters in which soap lathers easily are said to be *soft.* When such is not the case the water is spoken of as *hard.* Hard waters which can be softened by boiling are said to have *temporary* hardness. If the water cannot be thus softened, but requires the addition of a chemical, its hardness is said to be *permanent.*

QUESTIONS ON CHAPTER XXV.

1. Describe three different methods by which hydrogen may be obtained from water.

How would you prove that air contains one of the constituents of water?

2. What are the chief impurities in common water? How would you obtain pure water? Make a sketch of the necessary apparatus and explain the use of its several parts.

3. What chemical reaction takes place when a small piece of metallic sodium is thrown into water?

How would you test whether the solution left is acid or alkaline, and how would you prepare from it some common salt in the crystalline state?

4. Enumerate the differences in physical properties exhibited by sea and fresh water respectively. How can a specimen of common salt be prepared from sea water?

5. Blue vitriol heated over a lamp becomes white; on being treated with water the blue colour is restored. Explain the reason of these changes.

CHAPTER XXVI.

CARBON AND SOME OF ITS COMPOUNDS.

99. FORMS OF CARBON.

i. Carbon is obtained in organic substances.—Heat a series of organic substances, such as meat, wood, potato, egg, etc., in a crucible, and notice in all cases the production of a black residue, consisting largely of carbon. Heat more strongly and observe that it burns away, leaving an almost colourless ash.

ii. Properties of carbon.—Examine and write down the properties of as many of the following forms of carbon as you can obtain: diamond, blacklead, wood-charcoal, bone-black, and soot.

iii. Charcoal is porous.—(*a*) Show that charcoal floats in cold water. In boiling water charcoal sinks after a time, and then will not float again unless thoroughly dried. This is because air is driven out of the charcoal by the warmth of the water.

(*b*) Fill a large test-tube with ammonia gas (see p. 395) over mercury, and show that it can be absorbed by introducing a small lump of charcoal. This is a striking proof of the porosity of carbon.

Forms of carbon.—Carbon is a substance which is very widely distributed in nature, being present in all living matter, and in most products resulting from vital activity.

Carbon occurs, combined with other substances, in many rock masses, being a constituent of all the minerals known as carbonates. Combined with oxygen as carbon dioxide, it occurs in the atmosphere and dissolved in spring waters.

Diamond.—In the pure state carbon exists in various allotropic forms (p. 379). Of these the purest and the most valuable is the **diamond**. This form of carbon is crystalline, and very hard, being capable of scratching all other materials. Its refractive index (p. 224) is very high, and on this depends its brilliancy as a gem. Diamond is proved to consist of carbon by burning it in air or oxygen, when only **carbon dioxide** results.

Blacklead or **graphite** is another form of almost **pure** carbon, with properties totally different from those of the diamond. It is opaque and black, and so soft that it will mark paper. It is really a crystalline form of carbon, although good crystals are not very common. It occurs naturally in mines, chiefly in California, and was formerly largely obtained from Cumberland. Besides its use for lead pencils, it is also used as a lubricant.

Amorphous varieties.—Other forms of more or less pure carbon in an **uncrystallised** or **amorphous** state are **coke**, and **gas carbon**, which result from the heating of coal; **lampblack**, which is the carbon deposited by oils, etc., burning in an insufficient supply of oxygen; and **wood charcoal**, obtained by heating wood in closed retorts or in stacks under earth.

Charcoal has the power of absorbing colouring matter, and on the latter account it is used for decolorising solutions coloured by organic matter. **Animal charcoal** is really a misleading term, as the quantity of carbon present is usually only about 10 or 12 per cent., the remainder being chiefly bone ash.

Both animal and wood charcoal are very porous substances, and they have the power of absorbing gases to a large extent. Wood charcoal is used considerably on the Continent for heating purposes. Both kinds are useful in destroying noxious vapours.

Coal contains large quantities of carbon, especially the harder or anthracite coals, where the quantity may reach 94 per cent, being, however, only about 65 in brown coal or lignite.

Whenever any of the kinds of carbon burn freely in a good supply of air or oxygen, carbon dioxide is the compound formed, thus affording evidence that the three varieties are only allotropic forms (see p. 379) of the simple substance carbon.

100. CARBON DIOXIDE PRODUCED BY BURNING AND BREATHING.

1. When carbon is burnt carbon dioxide is formed.—(*a*) Heat strongly a piece of charcoal in a closed hard glass test-tube and show that without air it does not burn.

(*b*) Suspend a piece of glowing charcoal in a bottle containing lime-water. Shake up and show that the lime-water is turned

milky. Carbon dioxide can always be distinguished by this action upon lime-water, for it is the only common, colourless, inodorous gas which turns lime-water milky.

ii. **Carbon dioxide is produced by a burning candle.**—(*a*) Burn a candle or taper in a clean dry white bottle (Fig. 214). After the flame has been extinguished, withdraw the taper. Pour a little freshly made lime-water into the bottle and shake it up. Notice the milkiness of the lime-water.

(*b*) Cut a long thin chip of wood, hold it in the flame of a laboratory burner until it burns brightly, then thrust it into a cylinder or bottle, the bottom of which is covered with lime-water to the depth of about an inch. When the stick ceases to burn, withdraw it and shake the lime-water.

iii. **Carbon dioxide produced by breathing.**—(*a*) Blow through a piece of glass tube into some clear, freshly-made lime-water con-



FIG. 237.—The air breathed out from the lungs contains carbon dioxide, and will turn clear lime-water milky.

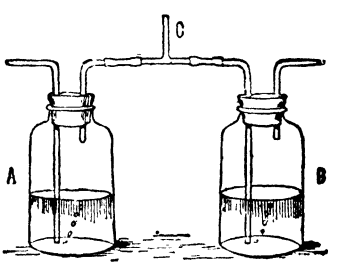


FIG. 238.—Air drawn into *A* does not turn the lime-water milky; but when blown through *B* from the lungs the lime-water in *B* is turned milky.

tained in a wine-glass or tumbler. Milkiness is at first produced, but if the blowing is continued long enough it disappears.

Fill a jar with water and invert it in a basin of water. Blow air from your lungs into the jar by means of a tube. When the jar is full of air place a glass plate under it and lift it out of the water. Show that the air will extinguish a lighted taper.

(*b*) Repeat the two preceding experiments by blowing air from a bellows instead of from the lungs. Notice that this unbreathed air has not the same effects upon a lighted taper or lime-water as breathed air.

iv. **The air contains carbon dioxide.**—Pour some clear lime-water into a blue dinner plate, or some other shallow vessel of a dark colour. Leave it exposed to the atmosphere for a little while. Notice the thin white scum formed on the top. The carbon dioxide in the air has turned the top layer of liquid milky.

v. **Breathing changes the character of air.**—Fit two bottles with corks and tubes as shown in Fig. 238. See that the corks are air-

tight. Put some clear lime-water into each bottle. Place the tube *C*, or an india-rubber tube leading from it, in your mouth. When you suck at the tube, air is drawn in through the glass tube which dips into the lime-water in the bottle *A*. When, however, you blow instead of sucking, your breath passes out through the tube which dips into the lime-water in the bottle *B*. Notice that the lime-water in *A* remains clear, but that in *B* is rendered milky by the air you breathe out. You thus see that fresh air has little effect upon lime-water, but breathed air quickly turns clear lime-water milky.

vi. Oxygen from plants.—Take a bunch of fresh watercress, or water weeds, and put it into a beaker or glass jar very nearly filled with water saturated with carbon dioxide. Cover the plants with a funnel nearly as wide as the jar, as shown in Fig. 239. Fill a test-tube with water and invert over the funnel. If properly managed there should at first be no gas in the test-tube. Place the jar in bright sunlight for an hour or two and then examine it. You will notice bubbles of a gas have collected at the top of the tube. Test the gas with a glowing splinter of wood. It is found to be oxygen.

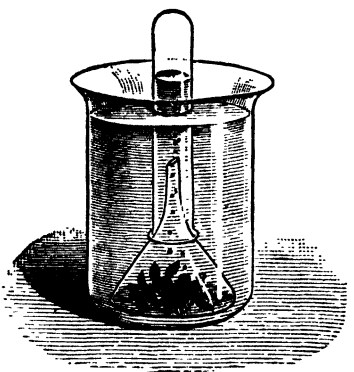


FIG. 239.—Green plants in bright sunlight can decompose carbon dioxide. They keep the carbon for themselves and liberate the oxygen.

vii. Plants in sunlight and in darkness.—Repeat the whole experiment, but instead of putting the bottle in bright sunlight place it in the dark. Observe that in such circumstances no bubbles of oxygen are formed.

viii. Carbon in plants.—Take some green portions of a plant (leaves will do) and heat them on a piece of tin plate over a laboratory burner. Note that they become charred, showing the presence of carbon in them.

Production of carbon dioxide by burning.—When things such as candles, oil, gas, and wood are burnt either in the air or in pure oxygen, a gas is produced which has the power of turning lime-water milky. All these substances contain, in one form or another, a constituent called carbon. As you have seen in previous chapters, the gas produced when these substances burn is carbon dioxide, that is, the gas obtained by burning carbon in air or oxygen. In fact, whenever a substance rich in

carbon burns in a plentiful supply of air or oxygen, this carbon dioxide is produced. Knowing how many fires there are in houses, furnaces, engines, and so on, it is not difficult to understand that at every hour of the day very large quantities of carbon dioxide are formed, which escape, sooner or later, into the air.

Carbon dioxide is given off in breathing.—If a person blows with the mouth into clear lime-water the lime-water is turned milky. This is another important fact. It is clear that carbon dioxide escapes from our mouths in breathing; and so it does from every animal. Not only, then, do all cases of ordinary burning result in the addition of carbon dioxide to the air, but also every act of breathing. It does not matter how small an animal is, all the time it is alive it is continually adding to the atmosphere a certain amount of the colourless, odourless gas which puts out flames and turns lime-water milky.

Purifying action of plants.—That there is always a certain amount of carbon dioxide in the air can be proved by exposing fresh lime-water in a shallow vessel. Very soon the lime-water becomes covered with a thin white layer of chalk, which is formed by the combination of the carbon dioxide in the air with the lime in the lime-water. One reason why there is never very much carbon dioxide in the air out of doors is because there is an agency continuously at work getting rid of this gas. This agency is the green parts of plants which occur everywhere.

When fresh watercress is put into a bottle completely full of water containing carbon dioxide in solution, and the bottle is inverted in a basin of water without allowing air to get into the bottle, it is found that, when the bottle and its contents are exposed to bright sunlight, bubbles of gas collect at the top of the bottle. These bubbles, when tested, are found to be pure oxygen. If, however, the bottle with the cress in it is kept in the dark no bubbles of oxygen collect. Or, if a bottle of water in which carbon dioxide is dissolved be put in sunlight, without any watercress, no oxygen collects in the top of the bottle.

In other words, two things are necessary for the formation of the bubbles of oxygen collected from the green plant as described. They are (1) the vegetation, (2) the sunlight. The same conditions have been found to hold true always, thus proving that **green plants in the presence of bright sunlight have**

the power of turning oxygen out of carbon dioxide. They keep the carbon for themselves, and it helps them to grow. If the experiment were carefully performed it would be found that the watercress had increased in weight after being exposed in these circumstances to bright sunlight for some time.

101. PRODUCTION OF CARBON DIOXIDE FROM CHALK BY THE ACTION OF AN ACID.

i. Preparation of carbon dioxide.—Into a flask or bottle, fitted like that in Fig. 240, place some chalk, or small pieces of marble. Place the delivery tube in a glass cylinder or a jar with a wide mouth. A disc of cardboard, through which the delivery tube passes, should rest on the top of the jar. Pour dilute hydrochloric

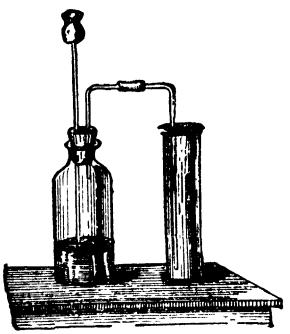


FIG. 240.—Apparatus for the preparation and collection of carbon dioxide.

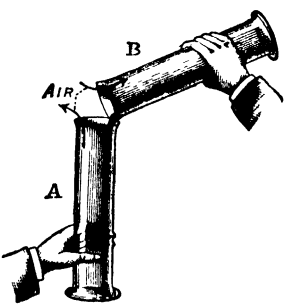


FIG. 241.—Carbon dioxide gas is heavier than air, and can therefore be poured from *B* into *A*, like a liquid.

acid down the funnel. During the effervescence a gas is given off and collects in the jar. When a burning taper is extinguished immediately it enters the jar, take out the delivery tube and put it into another jar. Cover the first jar of gas with a disc of card. In the same way collect several jars of the gas.

ii. Properties of carbon dioxide.—(*a*) Notice that the gas is (1) invisible and without taste or smell; (2) extinguishes a lighted taper; (3) must be heavier than air or it could not be collected in the way described.

(*b*) Pour the gas from one jar (*B*) into another (*A*), as shown in the diagram (Fig. 241), and test both jars by a lighted taper. It will be seen that the lower jar contains the carbon dioxide.

iii. Acid solution formed by carbon dioxide.—Pour a little water made blue with litmus into a jar of the gas and shake it up. Some of the gas dissolves, and the colour of the solution turns red. Boil the solution; the carbon dioxide is driven off, and the blue colour is regained.

iv. Action of carbon dioxide on lime-water.—Pass the gas from the delivery tube through some lime-water. Observe that a milkiness is produced, owing to the production of a white powder or *precipitate*, which disappears after a short time.

Boil the solution thus obtained, and notice that the milkiness again appears.

Filter the milky solution, and so obtain the white powder on a filter paper. Add a few drops of dilute hydrochloric acid to the powder. Notice the effervescence. Test the gas which is given off; it puts out a flame.

Preparation of quantities of carbon dioxide.—Though carbon dioxide is produced naturally by burning and breathing, and it also escapes from the earth in some regions, there are more convenient ways of obtaining the gas. Experiments show that when an acid is added to chalk, marble, or limestone, a gas is given off which puts out flames, turns clear lime-water milky, and possesses all the properties of carbon dioxide—it is, in fact, carbon dioxide. The best way to prepare bottles or jars of the gas is to place pieces of chalk or marble about the size of peas into a bottle fitted like that in Fig. 240. Dilute hydrochloric acid is poured down the thistle funnel, and when it comes into contact with the marble, the gas is given off. Enough acid is poured in to cover the lower end of the funnel, so the gas cannot escape up the funnel; the carbon dioxide passes through the other tube in the cork. The gas given off is heavier than air, and can therefore be collected as shown in Fig. 240. As the gas accumulates in the jar, the air is pushed out at the top. After several bottles or jars have been filled, the properties of the gas can easily be examined.

Properties of carbon dioxide.—An examination of the gas shows that it is colourless and has no smell. As it is heavier than air it can be poured downwards just like a liquid (Fig. 241).

Carbon dioxide is slightly soluble in water, and the solution which is thus formed turns a blue litmus paper red, just as acids do. For this reason the solution of carbon dioxide in water is often called **carbonic acid**, and the carbon dioxide itself is sometimes spoken of as **carbonic acid gas**. The gas has the

property of extinguishing the flame of a taper or match, and is consequently called a non-supporter of combustion.

Action of carbon dioxide on lime-water.—If carbon dioxide is passed into lime-water a milkiness is seen, but if the passage of the gas is continued the milkiness by and by disappears. If the clear solution which results after the disappearance of the white powder or precipitate is boiled, the milkiness again makes its appearance. The reason of this is that the white substance of which the precipitate is formed dissolves in water which has become saturated with carbon dioxide. When the clear solution, which appears after the solution of the powder, is boiled, the carbon dioxide is driven out of it, and the liquid again becomes pure water. The precipitate reappears, because it will not dissolve in water.

The chemical change when carbon dioxide is passed into lime-water.—What are the facts taught by the experiments described? When an acid is added to the white powder formed when carbon dioxide is passed into lime-water, a brisk effervescence is noticed and the colourless, odourless gas given off is found to put out a flame. But this is just what happens when the acid is dropped upon chalk, and putting the facts together they suggest that the white powder is really chalk, so that carbon dioxide gas combines with the lime in the lime-water to form chalk. We can write:

$$\text{Carbon Dioxide} \overset{\text{unites with}}{\longrightarrow} \text{Lime} \overset{\text{to form}}{\longrightarrow} \text{Chalk.}$$

From this statement it will also be seen that chalk consists of lime and carbon dioxide. Further evidence of the truth of this will be learnt later.

Uses of carbon dioxide.—The fact that carbon dioxide is a non-supporter of combustion is made use of in many forms of chemical fire extinguishers, which generally contain solutions of the gas under pressure or a means of producing large quantities of it by the action of acid upon a carbonate solution. As burning cannot take place in this gas, the flames are therefore extinguished when the gas reaches them.

The solubility of carbon dioxide in water is increased by pressure, and the sparkling nature of the various aerated waters, like soda-water, is due to the carbon dioxide with which they have been charged at high pressures, which escapes when the

pressure is reduced to that of the atmosphere by opening the bottle. As the pressure is very great, the bottles used as fire extinguishers have to be very thick.

Carbon dioxide is also produced during fermentation, the "rising" of bread being due to the escape of the gas which is generated by the fermentation, under the influence of the yeast, of the saccharine matters formed from the starch.

Carbon dioxide will not support life, and is sometimes used for suffocating stray dogs and cats.

By cold and pressure carbon dioxide may be liquefied and also solidified, forming a soft white substance, which when mixed with ether forms a powerful freezing mixture, the temperature sinking to about $-100°$ C.

Occurrence of carbon dioxide.—Carbon dioxide has been already stated to exist in the atmosphere, and to be produced by the oxidation of animal and vegetable tissues; also that under the influence of sunlight it is reconverted by the green parts of plants into its constituents, of which the carbon is used by the plant in the formation of new tissue. The gas is found in many natural gaseous emanations, and is frequently present to a large extent in the gases of caves and underground passages, where, owing to its high density, it tends to accumulate if formed by fermentation or other natural processes. In expired air, carbon dioxide is present to the extent of about 4·7 per cent. Although such air is not again respirable, this is probably due to the diminution of the oxygen, for it is doubtful whether carbon dioxide has any direct poisonous effect. The proportion of carbon dioxide may be increased even to 20 per cent. without immediate serious effects if only the quantity of oxygen be increased simultaneously.

102. PRODUCTION OF CARBON DIOXIDE BY HEATING CHALK.

1. **Change produced by heating chalk.**—Place a little powdered chalk (*not* blackboard chalk) on a piece of platinum foil, and heat it strongly for some minutes in the flame of a laboratory burner. If platinum foil is not at hand, heat a lump of chalk on a piece of coarse wire gauze for some time. Or, powdered whiting may be heated on a piece of thin tinplate. The latter is, however, a long

process. After heating, shake the powder on to a damp red litmus paper. Observe that the red litmus paper is in places changed to a blue colour.

ii. Action of lime upon litmus.—Test some wet lime with litmus paper. Observe that it changes the colour of red litmus to blue.

iii. Powder obtained by heating chalk.—Shake up in ordinary tap-water some of the powder obtained in Expt. 102 i. by heating chalk on platinum foil. Filter, or allow the powder to settle; then taste and notice the peculiar taste of lime-water.

iv. Solution of lime in acids.—Dissolve some lime in hydrochloric acid, and evaporate the solution to dryness. Note the formation of a white solid, which rapidly absorbs moisture from the air and liquefies. It has been previously used in experiments under the name of *calcium chloride*.

v. Composition of chalk.—Place a few small pieces of chalk in a test-tube. Add dilute hydrochloric acid until the effervescence, due to the production of carbon dioxide, ceases (see Expt. 101 i.). Filter the solution remaining in the test-tube and evaporate it. Notice the substance left; it is not chalk, but calcium chloride.

vi. Slaking of lime.—To a lump of fresh lime add a little cold water. Observe that the lump gets very hot and swells up.

Chalk undergoes a change when heated.—It is easy to prove by putting some powdered chalk upon a piece of moist red litmus paper that this substance is unable to change the colour of the paper. If, however, some powdered chalk be strongly heated on a piece of platinum foil in a laboratory burner and then placed on a piece of moist red litmus paper, the red colour is changed to blue. The chalk undergoes some change when heated, or it would not acquire this new property. The same chemical change as occurs when a little powdered limestone, or some chalk, is heated on platinum foil takes place on a large scale in the limekiln. In other words, when chalk and limestone are strongly heated they are changed into quicklime. The change is brought about by driving carbon dioxide out of them.

CHALK	when heated splits up into	LIME	and	CARBON DIOXIDE.

Changes produced by adding acid to chalk.—When hydrochloric acid is added to a little chalk a brisk effervescence occurs and a colourless, odourless gas which turns lime-water milky is given off.

If hydrochloric acid is poured upon a known weight of chalk, in such a way that all the gas evolved is got rid of, it is found that

the proportional loss of weight of the chalk is equal to the loss of weight on heating, thus showing that chalk contains a definite proportion of carbon dioxide gas.

If the solution remaining, after the effervescence of the chalk with the acid has ceased, is filtered and evaporated, a new substance known as **calcium chloride** is obtained. So that, from chalk and hydrochloric acid, it is possible to produce carbon dioxide, calcium chloride, and water (which is driven off by evaporation).

Substances which, like chalk, evolve carbon dioxide when acted upon by an acid are known as **carbonates**, and numerous carbonates exist, all possessing similar characteristics. Many of these on heating also give off carbon dioxide, the residue being known to be an oxide of a metal, so that they consist of carbon dioxide and a metallic oxide.

This leads to the idea that lime is also the oxide of a metal, and this view is now known to be correct, the metal being named **calcium**. Lime, therefore, is **calcium oxide** and chalk **calcium carbonate**.

Lime.—Lime produced by heating chalk or limestone—both of which are carbonates—is a white solid, which, if heated sufficiently, glows and emits a brilliant white light. It is on this account employed for the production of the limelight, where a small, hard cylinder of lime is strongly heated in an oxy hydrogen or oxy-coal-gas flame.

When water is added to freshly burnt lime, or **quicklime**, as it is termed, the water combines with it with the evolution of a large amount of heat, which is enough to boil the water if the quantity of lime is large. This can be seen at any time when bricklayers are preparing lime for making mortar. This addition of water to lime is called **slaking** it, and the altered lime is known as **slaked lime**. Lime dissolves to a slight extent in water, forming lime-water.

103. CARBON MONOXIDE.

i. Preparation of carbon monoxide.—In a hard glass tube *A B*, place some iron filings. Pass over these filings a current of carbon dioxide, dried in the usual manner. Heat the filings, and let the end *B* of the tube dip under a strong solution of caustic potash, which

absorbs any unchanged carbon dioxide, a tube containing the same liquid being inverted over the end (Fig. 242). Observe that the bubbles of gas which come from the end of the tube are *almost* completely absorbed, but that a gas slowly collects and fills the tube. Take out the tube and apply a light to the gas. Observe that it burns with a blue flame.

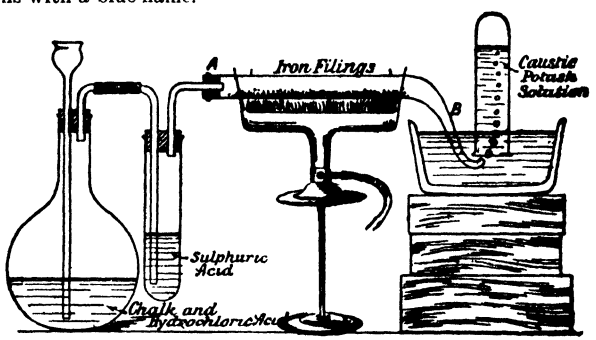


FIG. 242.—Preparation of Carbon Monoxide.

Lower oxide of carbon.—If a coke or charcoal fire be observed, flames of a peculiar blue colour are usually seen playing over the top. As, however, the only substances concerned in the combustion are carbon and oxygen, and since carbon itself does not burn with a blue flame, while carbon dioxide is quite incombustible, it appears probable that there must be some other compound of carbon and oxygen produced which burns with this blue flame. Further, it seems probable that the oxide is a lower oxide, *i.e.*, contains less oxygen, as if a compound with more oxygen existed we should expect the carbon dioxide to burn and form this oxide. **Carbon monoxide**, as this gas is called, can be prepared by depriving the dioxide of part of its oxygen. When passed over heated iron filings carbon dioxide is partly reduced to this lower oxide, while the iron oxidises or rusts. The mixture of dioxide and monoxide resulting from the action of the iron on the carbon dioxide is collected over a solution of caustic potash (Fig. 242). The caustic potash absorbs the whole of the unchanged dioxide, while the carbon monoxide gradually fills the collecting vessel.

CARBON DIOXIDE	and	IRON	give	CARBON MONOXIDE	and	IRON OXIDE.

Changes in a coke fire.—The formation of carbon monoxide in coke fires may now be readily explained. By the burning of the carbon in the lower parts of the fire, where the air enters and sufficient oxygen is present, carbon dioxide results, but

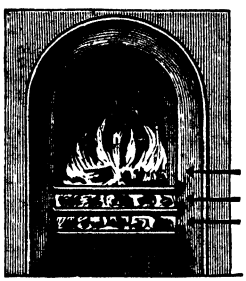


FIG. 243.—Chemical changes in the burning of carbon in an ordinary coal fire.

when this passes over the heated charcoal in the upper parts where insufficient oxygen is present, it parts with some of its oxygen, forming carbon monoxide. Thus,

CARBON DIOXIDE combines with CARBON to form CARBON MONOXIDE.

When it again meets sufficient oxygen at the top of the fire, this monoxide burns forming the dioxide. Carbon monoxide may in fact be prepared by passing the dioxide over heated charcoal instead of heated iron, but the temperature required is higher than that necessary in the experiment described on p. 362.

The gas so obtained is colourless and very poisonous. It acts as a direct poison, forming a compound with the hæmoglobin of the blood. For this reason the ventilation of rooms heated by coke stoves, and the complete removal of the products of combustion, should be always well secured.

CHIEF POINTS OF CHAPTER XXVI.

Carbon is present in all living matter. When organic substances are moderately heated, a black residue largely composed of carbon is left. If the temperature is raised, the carbon burns away and an almost colourless ash is left.

Carbon exists in several **allotropic forms** (p. 379). Two of these, *diamond* and *graphite*, are crystalline. Non-crystalline or *amorphous* carbon is known in varying degrees of purity as *coke*, *gas carbon*, *wood charcoal* and *animal charcoal*.

Burning.—When candles, gas, wood, or any substance which contains carbon burns in the air or oxygen, carbon dioxide is formed. Similarly, carbon dioxide results from the combustion of any form of carbon in air or oxygen. Carbon dioxide is also given off in the breath of animals. From these sources large quantities are continually passing into the atmosphere.

Breathing is really a kind of burning; the oxygen taken into the body by the lungs unites with the carbon of the body to form carbon dioxide, and with its hydrogen to form water; these products are expelled from the body partly in breathing.

The presence of carbon dioxide in the air may be shown by exposing lime-water in a shallow vessel. The lime-water soon becomes coated with a thin layer of chalk.

The green parts of plants have the power in bright sunlight of breaking up carbon dioxide; the carbon they keep to build up their solid parts, and oxygen is given off into the air.

Animals give off carbon dioxide (which passes into the air) and take oxygen from the air. Plants give out oxygen and take carbon dioxide from the air; in this way both are supplied from the air with the substances they need.

Preparation of carbon dioxide.—When chalk or limestone is heated it loses about 44 per cent. of its weight. This is due to the loss of carbon dioxide. The residue left behind is *lime*. The same gas is evolved when either chalk or limestone is acted upon by hydrochloric acid.

Carbon dioxide is a heavy gas which does not burn or support combustion. It is slightly soluble in water, the solution acting as a weak acid which turns blue litmus to a port-wine colour. This solution may be regarded as *carbonic acid*. The solubility of carbon dioxide in water is increased by pressure. By cold and pressure carbon dioxide may be liquefied, and also solidified to a soft white substance which, when mixed with ether, forms a powerful freezing mixture.

Lime is a white solid which is unchanged by heating. When heated intensely it glows and emits a brilliant white light. It dissolves in hydrochloric acid to form *calcium chloride*. When wet, lime turns a red litmus paper blue. When freshly made it is called *quicklime*. Quicklime combines with water, being changed into *slaked lime*.

EXERCISES ON CHAPTER XXVI.

1. Describe and explain the changes which take place when (*a*) limestone is burnt in a kiln; (*b*) water is added to some freshly burnt lime.

2. Charcoal burnt in air or oxygen in a bottle containing lime-water produces a white precipitate in the lime-water. This white precipitate, if collected and mixed with hydrochloric acid, dissolves with effervescence. What experiments would you make in order to compare the gas thus obtained with that obtainable by a similar process from the breath?

3. Describe the apparatus you would use for the production and collection of carbon dioxide gas; name the materials required, and describe the properties of the gas.

4. Four bottles are given you containing oxygen, hydrogen, nitrogen, and carbon dioxide respectively. By what experiments would you distinguish these gases from one another?

5. Describe the appearance of the three different forms of carbon. By what experiments would you prove that they are really only different forms of the same element?

6. What do you know about the chemical action that goes on in a "gasogene"—*i.e.* the common soda-water machine? Why does the "soda-water" rush out when the handle is pressed down?

7. Describe an experiment showing that water and carbon dioxide are formed in the combustion of a candle. How can ordinary coal-gas be made to burn with a non-luminous flame?

CHAPTER XXVII.

COMMON SALT. HYDROCHLORIC ACID. CHLORINE.

104. COMMON SALT.

i. Common salt is soluble in water.—Shake a little common salt with water in a test-tube for some time. Allow the undissolved salt to settle, and taste a little of the clear liquid. The taste of the solution is proof enough that common salt is *soluble in water.*

ii. Common salt crystallises in cubes.—Evaporate the clear solution and heat some of the dry residue in a test-tube.

iii. The solution of salt in water is neutral.—Test another portion of the clear solution of common salt with blue and red litmus papers in succession. The colour of neither paper is altered: hence, the solution of common salt in water is neutral.

iv. Salt crystals contain no water.—Heat a little *dry* salt in a test-tube. Notice the crackling and the absence of water on the side of the tube.

v. Action of strong sulphuric acid on salt.—Place a little dry salt in a test-tube and pour upon it some strong sulphuric acid. Gently warm the test-tube. Observe that a gas is given off, the part of which in the tube is colourless, though on coming into contact with the air at the mouth of the tube the gas fumes strongly. Plunge a lighted match into the tube, and notice that the flame is extinguished. Place a piece of moist blue litmus paper at the mouth of the tube, and observe that it is turned red, but the paper is not bleached.

Properties of common salt.—Common salt crystallises in six-sided solids, or cubes. When the crystallisation is brought about by evaporating a solution of salt, the crystals are very small. Some natural crystals, known as **rock salt**, are, however, of a considerable size. There is no water of crystallisation in the crystals, and consequently when they are heated no steam is given off. The crackling noticed when salt crystals are heated in a tube is spoken of as **decrepitation**, and is due to the breaking up of the crystals into smaller pieces.

Common salt dissolves in water, and the amount dissolved by cold water is almost as great as by warm water. The solution has no action on litmus papers, and is said to be **neutral.**

Rock salt is found in the earth in some countries as layers of varying thickness. The largest amount is found in the salt mines of Austria, but considerable quantities are found in this country, in Cheshire. Common salt is also present in large quantities in sea water, but is mixed with other substances.

Salt is used largely to prevent the decomposition of meat, and enormous quantities are employed in the manufacture of sodium carbonate.

105. HYDROCHLORIC ACID.

1. **Preparation of hydrochloric acid gas.**—Fit up the apparatus shown in Fig. 244. Remove the india-rubber stopper of the flask

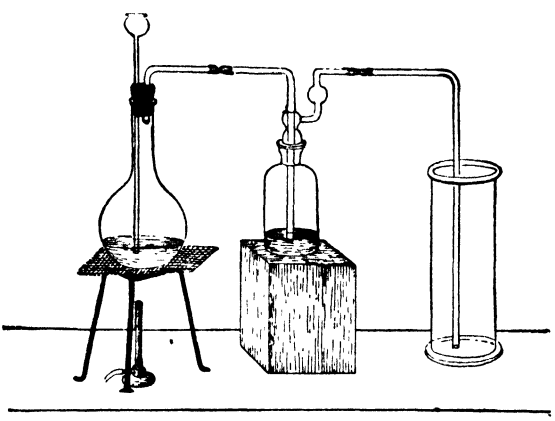


FIG. 244.—Preparation and collection of hydrochloric acid gas.

and place in it a small quantity of rock salt in small pieces, or some thoroughly dried common salt. Pour some strong sulphuric acid into the wash-bottle shown in the middle of the illustration. Re-insert the india-rubber stopper into the flask and pour down the acid funnel enough of a mixture[1] of strong sulphuric acid and water to

[1] One part of acid and one of water are convenient proportions. Be careful gradually to pour the acid into the water and not the water into the acid when mixing them, keeping the mixture well stirred throughout the process.

cover the salt in the flask. Gently warm the flask. Collect jars of the gas which is evolved (after it has bubbled through the strong sulphuric acid in the wash-bottle and so become freed from water vapour) in gas jars by downward displacement in the way the illustration makes clear. When each gas cylinder is full, which you can tell by holding a blue litmus paper just below the top of the outside of the cylinder until it is turned red, cover it with a ground-glass plate, with the ground side underneath. As the cylinders are filled set them on one side for examination as presently described. Collect four jars of gas in this way.

ii. Properties of hydrochloric acid gas.—(*a*) Raise the glass plate from the first jar and plunge a lighted taper into the gas. The flame is extinguished and the gas does not burn. Quickly replace the glass plate.

(*b*) Into the same jar drop a piece of moistened blue litmus paper and replace the glass plate. The paper is turned red, showing the gas has acid properties. Notice carefully that the paper is *not* bleached.

(*c*) Observe the fumes which the gas forms with the air when the glass plate is removed from a cylinder full of the gas. This is due to the very strong power of absorbing moisture possessed by hydrochloric acid gas.

(*d*) Firmly pressing the glass plate, invert a cylinder full of the gas and place it upside down in a basin of water. Remove the plate when the mouth of the jar is under water, and notice that water rushes up and completely fills the jar. If the water does not completely fill the jar, it shows that the air in the jar was not altogether displaced by the gas when you should have filled it.

iii. Solution of hydrochloric acid gas in water.—Modify the apparatus shown in Fig. 244. Remove the gas cylinder and the delivery tube which dips into it. Pour out the sulphuric acid from the wash-bottle, which thoroughly wash and half fill with water. If there is still enough salt and sulphuric acid in the flask, again warm it gently and allow the evolved gas to bubble into the water. Notice that it is completely dissolved. The solution of hydrochloric acid gas formed in this way is the "hydrochloric acid" of commerce.

iv. Some metals turn hydrogen out of hydrochloric acid.—Test the action of hydrochloric acid on zinc and iron.

v. Other proofs that hydrochloric acid contains hydrogen.—(*a*) Prepare hydrochloric acid gas in the manner described in Expt. 105 i. in the flask *A* (Fig. 245). Pass the gas so obtained over heated copper oxide in the hard glass tube *BC*. Observe that water collects in the test-tube *D*, and that the copper oxide is converted into a green substance. As water is formed the hydrochloric acid must evidently contain hydrogen.

(*b*) Collect a tube full of hydrochloric acid gas over mercury, and quickly introduce into it a piece of clean sodium. Allow it to stand, and observe that the volume of the gas becomes ultimately reduced to one-half the original volume (correction being made for the differences in temperature and pressure), while, further, the sodium

gets covered with a white powder which you may satisfy yourself is salt. Test the gas left with a lighted taper and see that it has the properties of hydrogen.

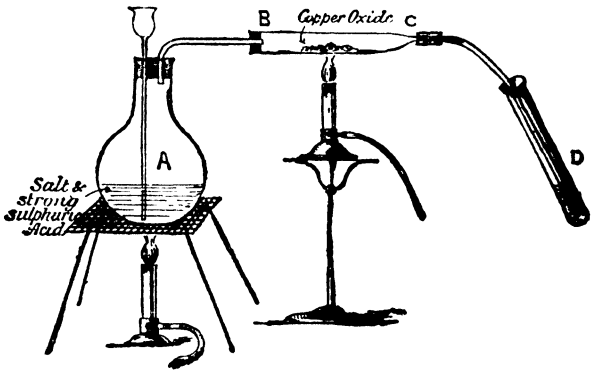


FIG. 245.—Hydrochloric acid gas can be shown to contain hydrogen by passing it over heated copper oxide and observing that water is formed.

vi. **Composition of common salt.**—Pass hydrochloric acid gas into a solution of caustic soda (known to chemists as sodium hydrate) until the solution has no effect either on a blue or red litmus paper. Distil the resulting solution and examine both the liquid which distils over and the residue left in the retort. Observe that the liquid is ordinary water and the residue common salt, which you can prove by tasting it.

Hydrochloric acid.—When common salt is heated with strong sulphuric acid a gas is given off which forms steamy fumes in the air. The gas readily dissolves in water, and the solution constitutes the hydrochloric acid of commerce. The solution was, because of its preparation from salt, originally known as "Spirits of Salt." Another common name given to it is Muriatic Acid. The "salt gas" itself is called hydrochloric acid gas. It is colourless, will not allow things to burn in it, nor will it burn itself. As is seen by its action on blue litmus paper, it is strongly acid. It is heavier than air, and can consequently be collected by downward displacement.

Composition of hydrochloric acid.—It has already been learnt that when the metal copper is heated strongly in air it combines with the oxygen of the air to form a black compound

called copper oxide, which compound is evidently entirely composed of copper and oxygen.

Now, if hydrochloric acid gas is passed over heated copper oxide which can be conveniently arranged in a tube, the black oxide is changed into a green substance, and water is formed at the same time. Moreover, if the green substance be acted upon with strong sulphuric acid, hydrochloric acid gas is again formed, just as when salt is similarly treated.

As water is formed, the hydrochloric acid gas passed over the heated copper oxide must evidently contain the hydrogen which is necessary for the formation of the water, for (p. 311) the copper oxide contains none. In fact, the simplest explanation of the experiment is that the hydrogen of the hydrochloric acid gas combines with the oxygen of the copper oxide to form water, while the copper combines with the other constituent of the hydrochloric acid (which you will presently learn to call **chlorine**) to form a green substance known as chloride of copper.

When metallic sodium acts upon hydrochloric acid gas contained in a tube over mercury, it is found after a short time that the volume of the gas is reduced to one-half and that the gas left in the tube is pure hydrogen. At the same time the sodium combines with the other constituent of the hydrochloric acid gas to form a white solid, which proves on examination to be common salt. These important experiments teach two facts:

1. That hydrochloric acid is composed of two constituents, hydrogen and another substance, which will be proved immediately to be chlorine, combined together in equal proportions by volume.

2. Common salt is a compound of the metal sodium and the second constituent of hydrochloric acid, chlorine.

Chemical behaviour of hydrochloric acid.—Many metals, such as zinc and iron, when dropped into a solution of hydrochloric acid gas in water (which will in future be spoken of as hydrochloric acid) break up the acid, evolving hydrogen gas and combining with its chlorine to form a chloride, or,

Zinc	in action with	Hydrochloric Acid	forms	Zinc Chloride	and gives off	Hydrogen.
Iron	in action with	Hydrochloric Acid	forms	Iron Chloride	and gives off	Hydrogen.

When hydrochloric acid is brought into contact with sodium hydrate, also called caustic soda, a double chemical change, or **double decomposition,** takes place, common salt or sodium chloride and water being formed (Expt. 105 vi.).

Hydrochloric Acid	in action with	Sodium Hydrate	forms	Sodium Chloride	and	Water.

106. CHLORINE.

i. Preparation of chlorine. – (*a*) Pour a little strong sulphuric acid upon some black oxide of manganese contained in a test-tube. Shake the tube until a paste is formed. Gently warm the tube, and after a minute or two examine the gas which fills it. It has no colour. Place a moistened blue litmus paper in the gas, and notice that, though the acid fumes redden the paper, it is not bleached.

(*b*) Repeat the experiment, but first mix the black oxide of manganese with an equal quantity of dry table salt. Again examine the gas evolved after heating the tube. It has a greenish-yellow colour. A moistened litmus paper has its colour completely removed, or it is bleached. The greenish-yellow gas is *chlorine.*

ii. Preparation of larger quantities of chlorine.—Fit up the apparatus shown in Fig. 244, and used in Expt. 105 i., for the preparation of hydrochloric acid gas. Remove the india-rubber stopper from the flask, put in some black oxide of manganese, and then enough strong hydrochloric acid to cover it. Shake the powder and acid together until no dry patches can be seen on looking through the bottom of the flask. Replace the india-rubber stopper, and be sure that the acid funnel dips beneath the surface of the liquid. Place in the wash-bottle an amount of *water* (not sulphuric acid as in preparing hydrochloric acid gas) like that shown in the figure. Gently warm the flask. Chlorine is at once given off. Collect several jars by downward displacement. You can easily see when the jars are full, as the gas is coloured. Cover the jars with dry ground-glass plates. **Caution.**—*It is very distressing to breathe chlorine; this experiment should therefore be done either in the open air or in a fume-cupboard with a good draught.*

iii. Bleaching power of chlorine.—In the first jar place a piece of damp Turkey-red cloth, a moist litmus paper, some coloured flowers, some writing done in ordinary ink, and a moist piece of newspaper. All the articles except the piece of newspaper are bleached. Chlorine bleaches vegetable colours. Printers' ink is not a vegetable colour.

iv. Chlorine combines spontaneously with some metals. – Powder a small piece of antimony and gently warm the powder, on a piece of notepaper, over the flame of a laboratory burner. Lift the glass

plate off a second jar of chlorine and sprinkle the warm antimony powder into the chlorine. The metal inflames on coming into contact with the gas, forming a chloride of antimony.

Or, drop a sheet of Dutch metal into a jar of chlorine. It will combine immediately with the gas with the accompaniment of flame.

v. Chlorine has a great affinity for hydrogen.—(*a*) Attach a wire to a candle. Light the candle and lower it into a jar of chlorine. Notice the candle continues to burn, but with a very smoky flame, depositing soot on the sides of the jar. The candle is composed of carbon and hydrogen. The chlorine combines with the hydrogen and leaves the carbon. Test the fumes left in the jar with a blue litmus paper and convince yourself they are acid.

(*b*) Boil a little turpentine in a test-tube and pour a drop or two of the hot liquid upon a piece of dry filter paper. When the turpentine has spread over the paper, fold it in a convenient manner, and drop it into a jar of chlorine. The turpentine at once inflames. A large quantity of carbon is thrown down, and a great deal of the steamy fumes of hydrochloric acid gas is formed. Turpentine, like a candle, is made of carbon and hydrogen.

vi. Chlorine is soluble in water.—(*a*) Invert a jar of chlorine in a basin of water, and notice that the water slowly dissolves the gas. The water rises up the jar, but not nearly so quickly as in the case of hydrochloric acid gas.

(*b*) Pass chlorine gas from the apparatus used in Expt. 106 ii. into a glass of water. Notice that the gas dissolves, the water eventually having the colour, smell, and bleaching power of chlorine gas. A solution of chlorine in water is called *chlorine water*.

(*c*) Leave some chlorine water for a day or two in bright sunshine. At the end of that time the chlorine water has lost its colour, smell, and bleaching power. The chlorine has combined with hydrogen to form hydrochloric acid, and oxygen has been given off. Test the liquid with a blue litmus paper, and see that it is acid.

Preparation of chlorine.—It has been shown that hydrochloric acid is a compound of hydrogen and a second substance; and that common salt, or sodium chloride, is a compound of the metal sodium with the same second substance. This second constituent, which is called **chlorine**, can be obtained separately either from common salt or from hydrochloric acid. It is more usual to prepare it from hydrochloric acid, and the plan adopted is to heat gently the solution of hydrochloric acid gas in water with black oxide of manganese, when chlorine is given off in large quantities in the form of a greenish-yellow gas.

The method used to get chlorine from common salt is first to mix it with black oxide of manganese and then to heat the mixture with strong sulphuric acid, when chlorine is evolved as in the previous case. This process is really the same as the

previous one, except that instead of first preparing hydrochloric acid from common salt and strong sulphuric acid (Expt. 105 i.) and then acting upon it with black oxide of manganese (Expt. 106 ii.), the two experiments are combined. The three materials are heated together, and the hydrochloric acid as it is formed is decomposed by the manganese dioxide which is present.

In manufacturing processes the compounds formed other than those sought for are called **bye-products.** The bye-products in the manufacture of chlorine are manganese chloride in the first method described above and manganese sulphate in the second.

Chlorine is heavier than the air and is usually collected by downward displacement, though it is sometimes collected over a strong solution of salt or over warm water.

Properties of chlorine.—Chlorine is a gas with a greenish-yellow colour, from which fact[1] it gets its name. It has a disagreeable smell, and the gas, if breathed, causes distressing symptoms, which have been described as like those of an exaggerated cold in the head. The gas is soluble in water and being heavier than air is collected usually by downward displacement. Its chief characteristic is its power of bleaching moist vegetable colours. Strictly speaking chlorine does not bleach. What happens is that the chlorine combines with the hydrogen of the moisture (which must be present for successful bleaching) to form hydrochloric acid and liberates the oxygen. This oxygen unites with the colouring matter to form a new chemical compound *which has no colour*, or, as chemists say, the oxygen oxidises the colouring matter. It is because of this power of liberating oxygen from water that chlorine is so useful as a disinfectant. The liberated oxygen combines with the noxious material, and, by oxidising it, renders it harmless.

The ease with which chlorine combines with hydrogen is seen not only by its action upon moisture, but in other ways. A lighted candle will continue to burn when plunged into chlorine, though with a very smoky flame. A candle is composed of hydrogen and carbon, and the flame continues, though with diminished brightness, because of the heat generated by the combination of the chlorine gas with the hydrogen of the candle to form hydrochloric acid gas. The carbon set free in the process is deposited as soot. The same explanation holds

[1] χλωρὸς (chlōros), green.

true for the spontaneous combustion of warm turpentine in chlorine gas.

Chlorine readily combines with metals to form **chlorides.** If finely divided iron, copper, antimony, and other metals be sprinkled into dry chlorine gas they at once combine with it, the heat of combination being sufficient to cause them to inflame. This happens more readily if the metals are first warmed.

Synthesis of hydrochloric acid.—When equal volumes of hydrogen and chlorine are mixed and exposed to sunlight, or the electric light, they combine together, with great violence to form two volumes of hydrochloric acid gas. There is, therefore, no diminution in volume. The case is simpler than that of the combination of hydrogen and oxygen to form steam. Two volumes of hydrogen combine with one volume of oxygen to form two volumes of steam, or the three volumes of the mixed gases are reduced to two after the combination has taken place. While

HYDROGEN 1 vol.	combines with	CHLORINE 1 vol.	to form	HYDROCHLORIC ACID GAS 2 vols.
HYDROGEN 2 vols.	combine with	OXYGEN 1 vol.	to form	STEAM 2 vols.

107. CHLORINE COMPOUNDS.

1. Action of chlorine on caustic potash.—(*a*) Pass chlorine for some time through a *boiling* solution of caustic potash. Allow the solution to crystallise, and examine the crystals which first separate out. They will be found to consist of *potassium chlorate*, which evolves oxygen when heated, either alone or with a little manganese dioxide.

(*b*) Pass chlorine through a *cold* solution of caustic potash, and observe that you no longer obtain potassium chlorate.

Some compounds of chlorine.—When chlorine is passed into a hot solution of caustic potash a substance known as **potassium chlorate** is formed, while potassium chloride is also produced. The **chlorates** may be regarded as derived from an acid, which we may call **chloric acid**, just as the chlorides are derived from hydrochloric acid. Chloric acid differs from the latter acid in containing oxygen as well as hydrogen and chlorine.

Of the chlorates, potassium chlorate is by far the most important, being largely employed in the manufacture of matches and in pyrotechny. On heating the chlorate either alone or mixed with manganese dioxide, oxygen is obtained. It is to the possession of a large proportion of oxygen that potassium chlorate owes its many uses.

On passing a stream of chlorine into *cold* caustic potash, another chlorine compound is formed mixed as before with potassium chloride. This compound contains a smaller percentage of oxygen than potassium chlorate and is termed **potassium hypochlorite.** When acted on by an acid this compound liberates chlorine. The chlorine thus set free again has a vigorous bleaching action. If chlorine is passed over slaked lime instead of into caustic soda, a compound is formed which may be regarded as a mixture of the chloride and hypochlorite of calcium. This also evolves chlorine when acted upon by an acid, and is hence very largely employed for bleaching, under the name of **bleaching powder,** which is manufactured on a large scale by the method indicated above.

CHIEF POINTS OF CHAPTER XXVII

Common salt occurs in nature as the mineral *rock salt.* It is soluble in water, the solution formed being neutral. Its crystals contain no water of crystallisation. The crystals decrepitate when heated. When treated with warm strong sulphuric acid a gas known as hydrochloric acid gas is evolved.

Hydrochloric acid, originally known as "spirits of salt" and muriatic acid, is prepared by the action of strong sulphuric acid on common salt; the gas evolved is dissolved in water to form the hydrochloric acid of commerce.

Hydrochloric acid gas is composed of equal volumes of hydrogen and chlorine. The hydrogen can be turned out of a solution of hydrochloric acid by the action of metals like zinc and iron. Hydrochloric acid is neutralised by sodium hydrate to form common salt.

Chlorine is easily prepared by the action of strong sulphuric acid upon a mixture of common salt and manganese dioxide, or by the action between hydrochloric acid and manganese dioxide. Owing to the distress caused by breathing chlorine, its preparation should be done in the open air or in a draught cupboard.

Chlorine is a greenish-yellow gas, heavier than air, and soluble in water. It is much used as a bleaching agent. It has a very great affinity for hydrogen, and easily combines with most metals to form chlorides.

Various compounds of chlorine.—*Potassium chlorate* can be obtained by passing chlorine for some time through a *boiling* solution of caustic potash and allowing the solution to crystallise. The *chlorates* are salts of an acid known as *chloric acid.* Potassium chlorate is the most important chlorate; it is used in the manufacture of matches and fireworks.

Potassium hypochlorite is obtained by passing chlorine through a *cold* solution of caustic potash. If slaked lime is used instead of caustic potash, a mixture called *bleaching-powder* is obtained.

EXERCISES ON CHAPTER XXVII.

1. How is hydrochloric acid obtained? Give a short account of its chief properties.

2. What are chlorides? How may they be obtained? Give examples.

3. Describe the properties of chlorine, and state how you would obtain the gas from salt and then reconvert it into salt.

4. Under what conditions does chlorine unite with (*a*) hydrogen, (*b*) phosphorus, (*c*) sodium?

5. A lighted taper is placed in a jar of chlorine; what happens, and why?

6. How may chlorine be (*a*) obtained from, (*b*) converted into, hydrochloric acid?

7. What is the action of sulphuric acid upon salt? What are the properties of both products?

8. From hydrochloric acid how could you obtain (*a*) hydrogen, (*b*) common salt?

9. Describe in simple language any crystals which you have seen. Say *either* how these crystals are made *or* where they are found.

10. What is bleaching powder? How is it obtained, and on what does its chief use depend?

11. In what form does hydrochloric acid exist under ordinary atmospheric conditions? Explain what occurs when hydrochloric acid is allowed to come into contact with bleaching powder.

12. What is an oxidising agent? In what sense can chlorine gas be regarded as an oxidising agent?

13. How do you account for the fumes that are seen when strong hydrochloric acid is exposed to the air? What experiments prove that hydrochloric acid is composed of equal volumes of hydrogen and chlorine?

CHAPTER XXVIII.

SULPHUR AND SULPHURIC ACID.

108. SULPHUR.

i. **Melting point of sulphur.**—Draw out in the flame of a laboratory burner, a piece of glass tubing so as to make a small thin-walled tube, about two or three inches long and $\frac{1}{10}$-inch in diameter. Into this tube place some finely powdered sulphur. Tie the filled tube on to a thermometer near its bulb with a piece of fine platinum wire, and put the thermometer into a beaker of sulphuric acid which has been placed over a burner. (Be very careful not to upset the acid.) Gradually heat the acid and keep it at a uniform temperature by moving the curved stirring rod (shown in Fig. 246) up and down. Notice when the sulphur melts, and at that instant read the thermometer. This reading will be the melting point of the sulphur.

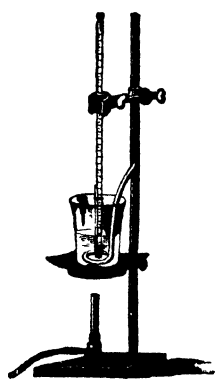


FIG. 246.—Determination of the melting point of sulphur.

ii. **Effects of heat upon sulphur.**—(*a*) Put some finely powdered sulphur into a large test-tube, using sufficient to fill the tube to a height of about 1½ inches, and heat carefully with a small laboratory burner flame. Occasionally take the tube out of the flame and shake it. When the sulphur has all melted notice that an amber-coloured liquid has been formed. Pour a little of the liquid into a beaker of water. Observe that a lump of yellow sulphur is again formed, which when broken reveals a crystalline structure.

(*b*) Continue to heat the remainder of the liquid sulphur obtained in (*a*) until the liquid boils. Carefully observe the changes in colour of the liquid. Pour a little of the boiling liquid into cold water. Examine the cooled sulphur; it is *plastic* and not unlike india-rubber.

(*c*) Notice that a yellow deposit has been formed on the cold, upper part of the test-tube in which the sulphur was heated. This is the result of the condensation of sulphur vapour. The deposit is known as *flowers of sulphur*.

iii. **Plastic variety of sulphur.**—After examining the plastic solid which results from suddenly cooling boiling sulphur, weigh a lump of it and place it on one side for a few days. After this interval examine it again and weigh it. There is no alteration in weight, but the lump has changed again to the ordinary kind of sulphur.

iv. **Crystalline varieties of sulphur.**—(*a*) Dissolve some powdered roll sulphur in carbon bisulphide. The solution must be made entirely by shaking, for *the carbon bisulphide must on no account be heated*. When the sulphur has all dissolved, pour the solution into an evaporating basin, and put it into a draught cupboard free from dust to evaporate slowly. Examine the basin after an hour or so. Observe the crystals which have formed. Take the largest and most perfect and sketch its form. This kind of crystalline sulphur is known as the *octahedral variety*.

(*b*) Place some powdered roll-sulphur in a clean, dry, evaporating basin and heat gently on a piece of iron wire gauze. When it has all melted, remove the flame and allow it to cool. As soon as a solid film has formed on the top of the liquid, pierce two holes in it and quickly pour out the remaining liquid sulphur through one of the holes. Remove the film of solid sulphur and examine the yellow, needle-shaped crystals on the sides of the basin. This kind of crystalline sulphur is known as the *prismatic variety* (Fig. 248).

Examine the crystals after a few days. Observe they are now opaque. The prismatic sulphur has changed back again to ordinary sulphur.

v. **Sulphur can be obtained from some minerals.**—Procure a little of the brass-like mineral called *iron pyrites*. Heat a fragment or two in a hard glass tube. Notice that melted sulphur gradually collects on the cold part of the inside of the tube.

How sulphur is found in nature.—Sulphur is found naturally both alone—that is, in an uncombined condition—and also united with other substances in the form of chemical compounds. The uncombined or **native sulphur** is rarely pure. It is found most abundantly in the neighbourhood of volcanoes—as, for example, in Sicily—from which island a large part of the **brimstone**, as sulphur is commonly called, of commerce is obtained. Before being placed on the market the sulphur is purified, or separated from the earthy impurities by distillation, in suitable retorts connected with large cool chambers, in which the sulphur vapour, driven off by heating, is condensed. In the first stages of the process the sulphur condenses as "flowers of sulphur," but later, when

the condensing chambers have become heated, as a liquid on the floor. This liquid is cast into the familiar "roll-sulphur" or brimstone which is well known to most people.

Sulphur occurs naturally combined with other substances in the form of **sulphides** and **sulphates**. The sulphides are compounds of sulphur with metals, among the most common being:

Galena,	containing	lead and sulphur.
Blende,	„	zinc and sulphur.
Iron pyrites,	„	iron and sulphur.
Copper pyrites,	„	copper, iron, and sulphur.

The naturally occurring sulphates are compounds of sulphur with metals and oxygen. Their constitution will be better understood after sulphuric acid has been studied. The most common mineral sulphates found in the earth are gypsum or calcium sulphate containing calcium, sulphur, and oxygen; and heavy-spar or barium sulphate, containing barium, sulphur, and oxygen.

Sulphur is also present in some of the compounds contained in animal and vegetable tissues.

Varieties of sulphur.—Sulphur is one of the few simple substances which exist in several forms. When such substances have more than one modification, all of them with the same chemical composition but possessed of different physical properties, such as density, colour, crystalline form, and so on, it constitutes an instance of what is called **allotropy**, and the different varieties of the substance are called **allotropic** forms. Sulphur, oxygen, carbon, and phosphorus all have allotropic forms. Sulphur has four allotropic forms, though it is only necessary here to mention three of them. These are **octahedral**, **prismatic**, and **plastic** sulphur. It must be carefully borne in mind that though the properties of these varieties of sulphur are so different, yet all the varieties are composed entirely of sulphur.

Octahedral sulphur.—Ordinary roll-sulphur, or brimstone, is composed of tiny crystals of this variety of sulphur compactly massed together. This can be seen by breaking a roll of sulphur in two and examining the broken ends, when crystals will be distinctly visible in the centre of the roll. But much larger crystals are obtained by dissolving powdered

roll sulphur in carbon bisulphide and allowing the solution to evaporate slowly into the air, when fairly large, perfectly formed octahedra of sulphur will be obtained. This kind of

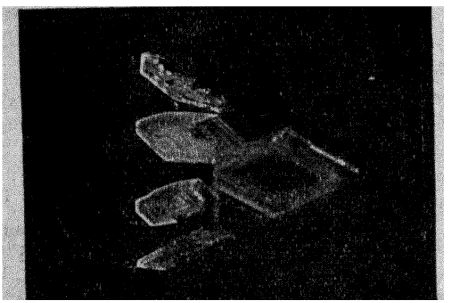


FIG. 247.—Crystals of copper sulphate.

sulphur is the most **stable** form; the other varieties gradually change into **octahedral** sulphur if left to themselves exposed to the air.

Prismatic sulphur. — The second crystalline variety of sulphur is called **prismatic** sulphur. It is obtained in the form of clear, needle-shaped crystals by carefully melting powdered roll-sulphur in an evaporating basin and allowing the liquid obtained to cool slowly. As soon as a film of solid sulphur has formed on the liquid sulphur, two holes are pierced, and the remaining liquid rapidly poured through one of them. If the film be removed and the inside of the basin examined, a number of clear, needle-shaped crystals of prismatic sulphur are seen (Fig. 248). But when the basin is examined again after an interval of a few days the crystals are no longer clear; they have become opaque owing to the transformation of each needle into a number of minute crystals of octahedral sulphur, which, as has been remarked, is the stable form of sulphur.

Plastic sulphur.—When boiling sulphur, which may be obtained by melting powdered roll sulphur in a large test-tube, is cooled suddenly by pouring it into cold water, it undergoes a remarkable change. If a piece of the sulphur, solidified in this manner, is taken out of the water and examined, it is

found to be like caoutchouc; it can be pulled about like chewing gum, and it is quite as elastic. This springy material is **plastic** sulphur. But if plastic sulphur be left to itself for a day or two it gradually changes back into octahedral sulphur, another reason for regarding the octahedral as the stable form of the element. In this process of reconversion there is no change of weight.

Effects of heat upon sulphur.—Sulphur undergoes a series of changes as it is heated. To follow the changes satisfactorily



FIG. 248.—Crystals of prismatic sulphur.

the heating must be very gradual. When powdered roll-sulphur is heated in a large test-tube it first melts, at about 114° C., into an amber-coloured liquid, which when poured into cold water solidifies into ordinary yellow sulphur. On continuing to heat the melted sulphur above 114° C., however, it gradually gets darker and darker in colour, becoming thicker and thicker in consistency, until at about 250° C. it is so viscid that the tube containing it can be inverted and the liquid will not flow. But if the temperature be still further raised, the thick liquid becomes mobile again, and by and by, at 440° C., it boils,

changing into a dark orange-red vapour. The vapour, by sudden cooling, can be changed into a yellow solid, known as "flowers of sulphur." If the boiling sulphur be poured into cold water it is converted into plastic sulphur.

109. OXIDES OF SULPHUR.

i. Burning of sulphur in air.—Place some sulphur in a deflagrating spoon (Fig. 220) and hold the spoon in a flame of a laboratory burner. Notice that the sulphur burns with a feeble, pale blue flame. Observe the suffocating fumes of "burning sulphur."

ii. Burning of sulphur in oxygen.—Refer to Expt. 92 v., and bear in mind sulphur dioxide is also formed in this case, and that the experiment shows the gas is soluble in water and that the solution has an acid reaction.

iii. Bleaching power of sulphur dioxide.—Remove the stopper from a bell-jar, and replace it by a tightly fitting cork through which a wire with a hook on the lower end is passed. Suspend a few brightly coloured flowers by a thread from the hook. In a small basin place some fragments of brimstone, and with the aid of the flame of a laboratory burner set the sulphur alight. Place the bell-jar with the flowers in it over the burning sulphur. Notice that after a time the flowers are bleached (Fig. 249).



FIG. 249.—The fumes formed when sulphur burns in air will bleach flowers

iv. Sulphur dioxide can be obtained fron sulphuric acid.—In a flask fitted with a cork and delivery tube place some copper turnings and some strong sulphuric acid, and heat over wire gauze by means of a Bunsen flame. A gas is given off, but as it is soluble in water, it cannot be collected in the same way as hydrogen and oxygen. Being heavier than air, however, it may be collected in the manner shown in Fig. 244. Obtain by this means two jars of the gas.

In (1) place a lighted taper, and in (2) a piece of moist red flannel or a coloured flower, allowing it to remain for some time.

v. Sulphurous acid and sulphites.—Using the flask you have already fitted for the preparation of the gas, allow the sulphur dioxide to bubble through a solution of caustic soda, which, we have already seen, has the power of turning red litmus blue, and is called an alkali. Observe that the gas is absorbed, and that after a

time the liquid smells of the gas and has become slightly acid. Boil the solution and observe that the smell disappears, showing that the gas is expelled by heating, while by further evaporation a white solid is ultimately left. Examine this solid, which is evidently soluble in water, and satisfy yourself that it is not caustic soda, having no action on litmus. Add to it a little sulphuric acid and observe the smell of sulphur dioxide.

Oxides of sulphur.—Sulphur forms two different compounds with oxygen, one called **sulphur dioxide**, the other **sulphur trioxide**. The latter compound contains half as much oxygen again as the former.

Sulphur dioxide.—The simpler oxide of sulphur is formed when sulphur burns in air or oxygen. The only difference in the two cases of burning is that when the sulphur combines with the oxygen of the air the combustion is feebler, and the sulphur dioxide formed is mixed with the nitrogen of the air; when the burning takes place in pure oxygen it is much more brilliant, and the sulphur dioxide formed is pure.

Sulphur dioxide is most commonly obtained in the laboratory in larger quantities by acting upon hot, strong sulphuric acid by means of copper or mercury. It can easily be shown, by heating copper in strong sulphuric acid, that

COPPER	in action with	SULPHURIC ACID	forms	COPPER SULPHATE (Fig. 247)	and	SULPHUR DIOXIDE.

In preparing the gas, the apparatus shown in Fig. 244 is employed, without the wash-bottle, and, as in the case of chlorine and hydrochloric acid gas, sulphur dioxide is collected by *downward* displacement. This is because the gas is heavier than air and soluble in water.

Bringing together the properties of sulphur dioxide, which different experiments already described have demonstrated, it may be said to be a colourless gas, with a suffocating odour, heavier than air, soluble in water—forming an acid solution—and possessed of strong bleaching powers. Its bleaching powers are made use of commercially in the preparation of straws and silks. It will not burn or support combustion.

Sulphurous acid.—The acid solution of sulphur dioxide in water is often regarded as **sulphurous acid**, the gas having entered into chemical union with the water. On boiling the water the whole of the dioxide is expelled. When sulphur

dioxide is passed into a solution of caustic soda, the caustic soda gradually loses its power of turning red litmus paper blue. On evaporation a white solid is obtained which is not caustic soda. The formation of this solid is an example of the production of a **salt**, and many other salts are formed by analogous methods. The salts produced by neutralising **sulphurous acid**, as the solution of sulphur dioxide is called, are known as **sulphites**, and the particular salt prepared in Expt. 109 v. is **sodium sulphite**. All the sulphites are characterised by their property of giving off sulphur dioxide when acted upon by sulphuric or other strong acid.

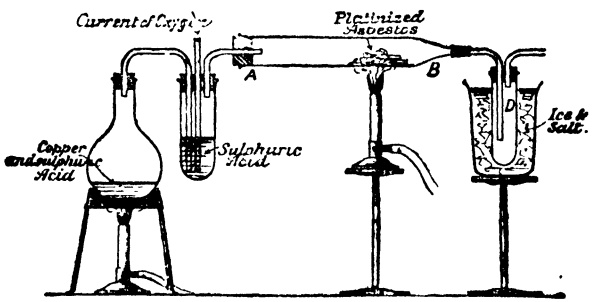


FIG. 250.—Apparatus for preparation of sulphur trioxide.

Sulphur trioxide.—Sulphur dioxide can be made to combine with a further quantity of oxygen. The process is not so straightforward as most of those already described, but it is not difficult to understand. A mixture of sulphur dioxide and oxygen is passed over heated platinised asbestos. [Platinised asbestos is prepared by dipping asbestos fibres into (1) platinic chloride solution, (2) ammonium chloride solution, and heating strongly. By this means the asbestos becomes coated with a quantity of very finely divided platinum.] White fumes are formed in abundance, and when cooled sufficiently they change into a white powder, or crystals, known as **sulphur trioxide**. The apparatus necessary for the purpose is shown in Fig. 250.

SULPHUR DIOXIDE	combines with more	OXYGEN	to form	SULPHUR TRIOXIDE.

The crystals of sulphur trioxide obtained in the above process are very soluble in water, the solution being accompanied by a hissing noise as when the oxide of phosphorus (p. 391) dissolves in water. The solution obtained is **sulphuric acid.**

SULPHUR TRIOXIDE combines with WATER to form SULPHURIC ACID.

110. SULPHURIC ACID.

i. Relative density of sulphuric acid.—Make a mark upon a small beaker about one-third of the height from the bottom. Weigh the beaker. Pour distilled water into the beaker up to the mark. Weigh again and so determine the weight of the water. Pour out the water and carefully dry the beaker inside and out. Fill up to the mark with strong sulphuric acid, being careful to upset no acid on the pan of the balance. Weigh again and determine the weight of the acid. What is the relative density of the acid ? (See p. 62.)

ii. Heat developed during solution of sulphuric acid.—Pour strong sulphuric acid into a graduated measuring jar with a spout. Place a measured quantity of water, say 250 c.c., in a jug or large beaker. Take the temperature of the water. Add 25 c.c. of the acid to the water, pouring it out in a continuous small stream (for if the acid splashes it will injure you or your clothes). Keep the water stirred with the thermometer all the time; read the highest temperature recorded by the thermometer. Similarly add a second 25 c.c. of acid and again read the temperature. Then add a third 25 c.c. of acid, always keeping the mixture stirred, and again read the temperature.

Water should never be poured into strong sulphuric acid because of this generation of heat.

iii. Sulphuric acid has a strong affinity for water.—(*a*) Into a small weighed beaker, or wide-mouthed bottle, put enough strong sulphuric acid to reach to about one-third of the height. Weigh again and determine the weight of the acid. Stick a piece of gummed paper outside the beaker to mark the level of the acid. Leave the acid exposed to the air for a day and then weigh again. Notice the increase in weight and also that the level of the liquid is now above the paper. The increase in weight and volume is due to water absorbed from the air.

(*b*) Pour the strong acid on sugar. The charring which is observed is due to the absorption of water and the separation of carbon.

iv. Acid properties.—(*a*) Observe again the acid properties by tasting the *very* dilute solution and testing it with blue litmus paper.

(*b*) Place a little of a dilute solution of sodium hydrate in a beaker; test it with a red litmus paper. Now add, little by little, some of

the acidified water. A point will be soon reached when the solution has no effect on litmus. The acid has *neutralised* the sodium hydrate. If the solution be evaporated nearly to dryness and allowed to stand crystals will be obtained, known as *sodium sulphate.*

v. Action upon metals.—Refer to Expts. 95 i. and 105 iv.

vi. Test for sulphuric acid and soluble sulphates.—To dilute sulphuric acid or a solution of any soluble sulphate, such for example as the sodium sulphate obtained in Expt. 110 iv., add a few drops of barium chloride or barium nitrate solution. Notice the dense white precipitate which is formed. The precipitate cannot be got rid of either by boiling or by the addition of acids.

Sulphuric acid.—Its simple properties and its action on metals have already been studied. With some metals, *e.g.* zinc, it reacts when cold and dilute, liberating **hydrogen** and forming a **sulphate** of the metal; with others, *e.g.* copper, it has no action until heated, when it produces a **sulphate,** but with the liberation of **sulphur dioxide.**

With substances known as alkalies, of which sodium hydrate is a typical example, sulphuric acid also forms sulphates, that resulting from its action upon sodium hydrate being known as sodium sulphate.

A similar action takes place when the **base,** such as magnesia or magnesium oxide, neutralises sulphuric acid. In this case **magnesium sulphate** is produced.

CHIEF POINTS OF CHAPTER XXVIII.

Sulphur is a brittle, yellow solid, which may easily be reduced to a fine powder. It is insoluble in water, but dissolves in carbon bisulphide. It melts at about 114° C. to a clear, yellow, mobile liquid, which, when poured into cold water, solidifies to ordinary yellow sulphur. On further heating, the yellow liquid becomes darker in colour, and more viscid, until at about 250° C. it will not run out, even though the vessel containing it is inverted. At still higher temperatures the liquid again becomes thin and mobile and finally boils, evolving a dark orange-red vapour, which condenses either to an orange liquid or to a ye low powder. If the boiling sulphur be poured into cold water it solidifies to a solid resembling caoutchouc.

This elastic solid is called *plastic sulphur.* If left in contact with air it returns to ordinary sulphur in a few days without any change of mass. The yellow powder into which sulphur vapour condenses, without passing through an intermediate liquid state, is called *flowers of sulphur.* Ordinary commercial sulphur is called *roll sulphur.*

Crystalline sulphur.—The crystals left when a solution of sulphur in carbon bisulphide is allowed to evaporate have an octahedral form. Those obtained from melted sulphur are needle-like and

called prismatic. The prismatic sulphur crystals will, if left alone, gradually change back to the octahedral variety.

Allotropy is the property some simple substances, like sulphur, possess of existing in different forms which are known as *allotropic forms*. The chief allotropic forms of sulphur are the octahedral, prismatic, and plastic.

Oxides of sulphur.—*Sulphur dioxide* is formed when sulphur burns in air or oxygen. It is also given off when copper is heated with strong sulphuric acid. It is a gas with a pungent smell, which does not burn or support combustion, and has the power of bleaching vegetable colours. It dissolves in water to form *sulphurous acid*. Sulphurous acid forms salts called *sulphites*.

Sulphur trioxide.—By suitable means sulphur dioxide can be made to combine with more oxygen to form a higher oxide known as sulphur trioxide. This oxide dissolves in water, with a hissing noise accompanied by the evolution of much heat, to form *sulphuric acid*.

Sulphuric acid is a heavy, oily liquid which boils at 335° C., giving off choking, pungent, white fumes. It mixes with water in any proportion with the evolution of much heat. It absorbs moisture very readily, and is consequently used for drying gases. For the same reason it chars any organic substance it comes in contact with. Sulphuric acid forms with alkalies salts called *sulphates*.

EXERCISES ON CHAPTER XXVIII.

1. Describe the changes which sulphur undergoes when heated.

2. What is plastic sulphur, and how is it obtained? How would you prove it consists solely of sulphur?

3. What happens when sulphur burns? By what other method can you obtain the product formed?

4. Give an account of the properties of sulphur dioxide.

5. Describe the appearances and properties of sulphur trioxide. How is it obtained, and what is its action on water?

6. How would you show that the gas formed when brimstone burns in air has the power of bleaching vegetable colours?

7. How is the element sulphur found naturally? Name as many chemical compounds as you are acquainted with, known as minerals, which contain sulphur as one of their constituents.

8. What do you mean by the "melting-point" of a substance? Why cannot we speak of the "melting-point" of glass? Contrast the "melting-points" of mercury, lead, and iron.

9. Give as many examples as you can of substances which, whilst differing in appearance and physical properties, are notwithstanding identical in chemical composition. Indicate in each case how chemical identity could be experimentally demonstrated.

10. What gaseous compounds of sulphur are you acquainted with? Describe how you would propose to obtain them from ordinary *flowers of sulphur* as the source of the element.

11. How would you prepare crystals of sulphur from flowers of sulphur? What experiment would establish the chemical identity of the various forms of sulphur?

12. What do you know about the action of sulphuric acid on (1) lead, (2) iron, (3) copper?

CHAPTER XXIX.

PHOSPHORUS. COMPOUNDS OF NITROGEN. NITRIC ACID AND AMMONIA.

111. PHOSPHORUS AND ITS OXIDES.

i. Properties of phosphorus.—(*a*) Observe the appearance of some ordinary sticks of phosphorus. Notice the manner in which they are kept. Carefully cut a stick *under water*, and observe the appearance of the cut surface. Note the phenomena observed when phosphorus is exposed to the air in the dark.

(*b*) Place a little phosphorus in an evaporating basin containing water, and heat slowly. Note the temperature at which the phosphorus is seen to melt. Allow it to cool.

(*c*) Shake up a small piece of phosphorus in carbon bisulphide (see Expt. 108 iv.). Notice that it dissolves. Pour the solution on to a piece of filter paper. Observe the effect.

(*d*) Repeat Expt. 88 i. Observe the dense white fumes. Add water and shake—the fumes dissolve. In the solution put a little blue litmus paper. Note that it is turned *red*.

General characters of phosphorus. — Phosphorus, like sulphur, exists in different allotropic forms, and these must be examined first. Ordinary phosphorus, which is always kept under water, will probably appear at first to be a dark yellow, or brown, opaque solid. This is not, however, its true appearance, but only that of a film which coats the exterior. If a piece of phosphorus be cut it will be found to consist of a waxy, translucent solid, of a pale yellow colour. It very readily inflames; contact with any warm object is sufficient to start the combustion. For this reason it should not be handled, and when not in use should be kept immersed in water, in which it is insoluble. Phosphorus glows in the dark, forming white fumes. Both glow and fumes are due to its gradual oxidation. Phosphorus itself, as well as the

fumes formed when it is exposed to the air, are poisonous, and hence care must be taken during its use, both on this account and on account of its inflammability.

If a piece of ordinary phosphorus is placed in water in a small basin and heated, it will be found that the phosphorus melts at about 43° C., but remains liquid below this temperature. This is frequently the case with melted solids, but during the solidification the temperature again rises to the melting point. If the water be poured off, the phosphorus frequently catches fire, as it inflames in air at a temperature below the melting point. Yellow phosphorus, though insoluble in water, is readily soluble in carbon bisulphide. If its solution in this liquid be allowed to evaporate on a piece of filter paper, the whole may ignite spontaneously. The evaporation of the solvent leaves the phosphorus deposited on the filter paper in so finely divided a state that it rapidly oxidises, charring the paper, even setting it alight. By careful evaporation of the solvent, crystals may be obtained.

Red phosphorus.—This allotropic form of phosphorus is a dark, brownish-red powder. Examine it carefully and note well the differences from the ordinary variety; substitute red phosphorus and try again the various experiments performed already. Red phosphorus will be found to be insoluble in carbon bisulphide, and to ignite only when strongly heated (240° C.). It is not luminous, neither does it oxidise when exposed to moist air. It is therefore not necessary to keep red phosphorus under water. Red phosphorus, or **amorphous phosphorus**, as it is also called, is, further, non-poisonous. The chief use of phosphorus is in the making of matches. The common red match is tipped with a mixture containing ordinary yellow phosphorus. The so-called safety match has no phosphorus on its tip, but is struck by friction against a specially prepared surface coated with a mixture containing red phosphorus.

Phosphorus and oxygen.—When phosphorus burns it forms a compound called **phosphorus pentoxide**. This is the product of combustion when either red or yellow phosphorus is used, and this fact may serve to prove that the two varieties are chemically identical. This oxide is readily obtained as dense white fumes when phosphorus is burnt in either air or oxygen; these fumes

settle into a white amorphous powder which dissolves with a hissing noise very readily in water, forming an acid solution.

If left exposed to the air the oxide absorbs moisture, being, in fact, one of the most powerful of drying agents, on account of which it is frequently used as a **dehydrator** (*i.e.* for the purpose of removing moisture from gases or liquids). For the same reason it is of service in promoting many chemical reactions, the essential part of which is the removal of the elements of water.

Phosphoric acid and phosphates.—The acid solution which is formed when phosphorus pentoxide is dissolved in water is known as **phosphoric acid.**[1] We may also obtain the acid directly from phosphorus itself by means of nitric acid, which is a powerful **oxidising agent** (*i.e.* a substance which readily gives up oxygen to other bodies).

When nitric acid and phosphorus, which may be of either variety, are gently warmed together, an energetic action results, accompanied by an evolution of red fumes. On evaporation to dryness a gelatinous product, which cools to a hard glassy mass, remains.

The phosphoric acid so obtained is a crystalline solid which dissolves in water. Like other acids, it may be neutralised by the addition of caustic soda, and a salt obtained. The salts obtained from phosphoric acid are known as **phosphates.**

Phosphorus trioxide.—Other oxides of phosphorus exist, however, besides the pentoxide, and of these the most important is that known as **phosphorus trioxide,** *phosphorus oxide,* or *phosphorus anhydride.* This oxide results, to a small extent, when phosphorus burns in air, the quantity increasing when the supply of air is insufficient. It is a white solid with a garlic-like odour, and it differs greatly from the pentoxide in its action on water, as it is only very slowly dissolved, forming an unstable acid, **phosphorus acid,** the salts of which are called **phosphites.**

Manufacture of phosphorus.—Phosphorus is employed for various purposes, but the greatest quantity is used for the production of matches, the tips of which consist of a little wax with phosphorus and potassium chlorate. The chief source of phosphorus is the residue obtained on burning bones, which is

[1] This acid results when the water is hot: when cold water is used, another acid, which gradually changes to this one, is produced.

known as *bone ash*, and consists of *phosphate of calcium*. This product is converted into phosphoric acid by treating with sulphuric acid, when the following change takes place:

CALCIUM PHOSPHATE	and	SULPHURIC ACID	form	CALCIUM SULPHATE	and	PHOSPHORIC ACID.

The calcium sulphate, being insoluble, is easily separated from the acid, which is then concentrated and heated with powdered coke in cast-iron retorts connected with pipes dipping under water, by which means phosphorus is obtained.

The crude phosphorus so prepared is purified by melting under warm water, the melted phosphorus being then cast, while still under water, in the form of round sticks.

112. NITRIC ACID.

i. Preparation of nitric acid.—Into a stoppered retort, such as is used in the distillation of water (p. 349), place 30 or 40 grams of small crystals of potassium nitrate (also known as nitre). Using a

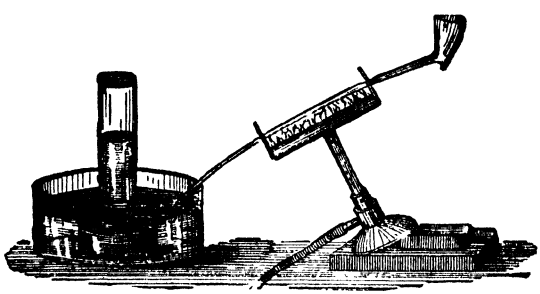


FIG. 251.—An experiment to show that nitric acid easily gives up oxygen

funnel, carefully introduce enough strong sulphuric acid to cover the nitre. Replace the stopper. Place the retort on a stand as shown in Fig. 236, and insert its neck in that of a flask which is continually kept cool by water, just as in the distillation of water. Gently heat the retort. Brown fumes are given off in abundance, and soon drops of a light yellow liquid are seen to fall into the receiving flask. When enough nitric acid has distilled over, remove the laboratory burner, and while the materials in the retort are still liquid, pour them, after removing the stopper, from the retort into an evaporating dish.

ii. Properties of nitric acid.—Test the action of nitric acid upon metals and upon litmus.

iii. Nitric acid easily gives up oxygen.—Procure a clay tobacco pipe of the pattern known as a "church-warden," and, by means of a retort-stand with a clip, support it in the position shown in Fig. 251. Let its stem dip into a pneumatic trough in the manner the illustration makes clear. Heat the middle of the stem with a Bunsen burner until it is red-hot. When this is the case, drop strong nitric acid slowly into the bowl. You will notice that a colourless gas collects in the cylinder over the pneumatic trough. When a sufficient amount of the colourless gas has been obtained in the cylinder, lift the stem of the pipe out of the water, and remove the burner. Test the gas in the cylinder with a glowing splinter of wood. The gas is oxygen.

iv. Formation of a nitrate.—Dissolve some caustic soda in water and add nitric acid gradually until a blue litmus paper placed in the solution just becomes purple. This shows that the acid has neutralised the caustic soda. Evaporate the solution till crystals appear on cooling, and allow the sodium nitrate to crystallise out. This experiment is an example of the production of a neutral salt from an acid and a soluble base, or alkali.

Nitric acid.—One of the most important compounds with which chemists are familiar is an acid which has long been known under the name of **aqua fortis.** It is a compound of nitrogen with hydrogen and oxygen. It can be prepared by synthesis (p. 343) from its constituents, but the experiment is rather one of theoretical interest than of practical importance. It is always prepared by distilling a nitrate with strong sulphuric acid. Either potassium nitrate, which is more familiarly known as "saltpetre" or "nitre," or sodium nitrate, also called "Chili saltpetre," is generally employed. The latter salt is the cheaper, and in addition yields a larger amount of nitric acid for a given expenditure of sulphuric acid, so that it is more commonly employed than ordinary saltpetre in the manufacture of nitric acid.

The changes which take place when potassium nitrate is distilled with sulphuric acid in a glass retort, as described in Experiment 112 i., may be expressed by stating that

Potassium Nitrate	and	Sulphuric Acid	form	Potassium Hydrogen Sulphate	and	Nitric Acid.

In the manufacture of nitric acid on a large scale, however, greater heat is employed. This is possible because earthenware retorts replace the glass vessel of Experiment 112 i. In

these circumstances the chemical change is taken a step further :—

Potassium Nitrate	and	Potassium Hydrogen Sulphate	form	Potassium Sulphate	and	Nitric Acid.

A further quantity of nitric acid is, in this way, obtained with the same expenditure of sulphuric acid.

For practice the student should write out statements to show the changes that take place when sodium nitrate is made use of instead of potassium nitrate.

Properties of nitric acid.—The simple properties of nitric acid have been studied in the experiments on p. 393, to which the student should refer. Its chief uses and its great activity are due to the readiness with which it gives up oxygen, and for this reason it is called an oxidiser. Not only does nitric acid give up oxygen when it is passed over a heated surface, as in Experiment 112 iii., but also in coming into contact with different substances. For example, if strong nitric acid be dropped upon heated sawdust, the organic material catches on fire. Or, if a stick of charcoal be made red-hot and plunged beneath the surface of some strong nitric acid, it will burn brilliantly at the expense of the oxygen in the nitric acid.

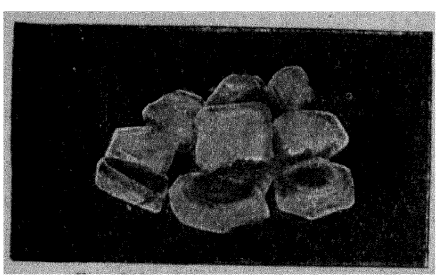


Fig. 252. Crystals of lead nitrate.

Nitric acid acts readily upon most metals. With lead it gives off red fumes and the liquid when evaporated yields colourless crystals of lead nitrate (Fig. 252). Copper is rapidly dissolved, a green liquid being produced, which becomes blue on diluting with water, together with red fumes. The liquid on evaporation yields dark blue crystals of copper nitrate. In a similar manner mercury nitrate in the form of colourless crystals can be made by dissolving mercury in nitric acid.

113. AMMONIA.

i. Preparation of ammonia.—(*a*) Place some ammonia solution in a flask and boil, allowing the gas that is evolved to pass over quick-lime or solid caustic potash in order to dry it. Collect several jars of the gas as shown in Fig. 253. Observe the gas has the same smell as the liquid and acts similarly on red litmus.

ii. Properties of ammonia.—(*a*) In one jar place a lighted taper.

(*b*) Place a second jar in water, and note the rapid absorption of the gas and the rise of the water in the jar.

(*c*) Dip a rod in hydrochloric acid and hold it over a jar of the gas. Observe the formation of white fumes.

(*d*) Shake up a jar of the gas with a little water. Examine the solution. See that it is identical with the "ammonia liquid," from which the gas was obtained, and that it loses its odour on boiling, the gas being evolved.

(*e*) Heat an ammonium salt with caustic soda or lime, or merely well mix the two in a mortar with the addition of a little water. Observe by the smell that ammonia is produced.

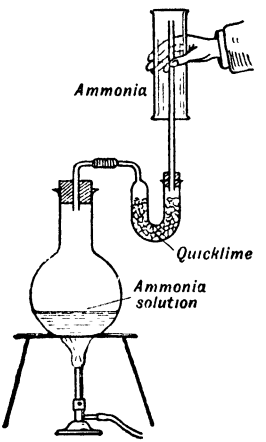


FIG. 253.—Preparation of Ammonia.

Properties and composition of ammonia.—The liquid commonly known in the laboratory as ammonia has a sharp pungent smell and colours red litmus blue. If this liquid be heated, a gas is given off which may be dried by passing over quick-lime, and collected by upward displacement, as illustrated in Fig. 253. It is found that the "ammonia liquid" is merely a solution of a gas, viz. ammonia, which is very soluble in water, turns red litmus blue, does not support combustion, and is apparently not combustible. In an atmosphere of oxygen, however, the gas readily burns, forming nitrogen and water. It hence contains nitrogen and hydrogen, and may, in fact, be proved to have the composition, three volumes of hydrogen to one volume of nitrogen in two volumes of the gas. To prove this, a long tube is filled with chlorine

gas, and by means of a funnel and the stopcock (Fig. 254) some strong liquid ammonia is run into the gas. Combination ensues between the chlorine and the hydrogen of the ammonia with the formation of white fumes, and frequently a flash of light. The stopcock is then opened under water, when the water rises and fills two thirds of the volume of the tube, the gas remaining being nitrogen. As chlorine combines with an equal volume of hydrogen to form hydrochloric acid, the volume of nitrogen in the ammonia is only one-third that of the hydrogen.

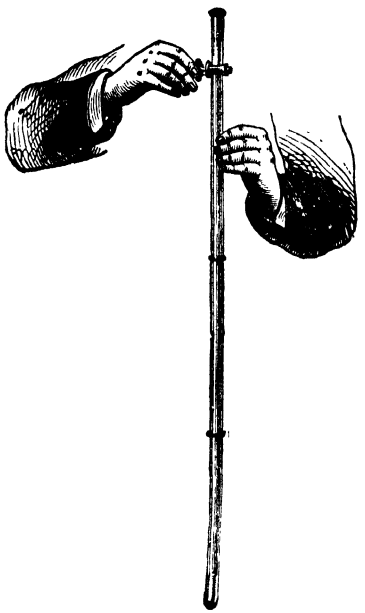


FIG. 254.—Tube for determining the composition of ammonia.

The white fumes formed in the above experiment and by the action of the gas on hydrochloric acid are fumes of ammonium chloride:

$$\text{AMMONIA and } \underset{\text{ACID}}{\text{HYDROCHLORIC}} \text{ form } \underset{\text{CHLORIDE}}{\text{AMMONIUM}}$$

The gas ammonia is very readily driven off from ammonium salts by the action of caustic soda. This method is very commonly used for the preparation of the gas.

When ammonium salts are heated they are generally driven into vapour. Ammonium chloride sublimes without decomposing. Ammonium nitrate is decomposed into a gas known as nitrous oxide, which closely resembles oxygen in its properties, and water in the form of steam.

CHIEF POINTS OF CHAPTER XXIX.

Phosphorus, like sulphur, exists in different allotropic forms.

Ordinary phosphorus is a waxy, translucent solid of a pale yellow colour. It very readily inflames and should not, in consequence, be handled, and when not in use should be kept immersed in water. It glows in the dark, evolving white fumes. It is poisonous. It melts at about 43° C. It is soluble in carbon bisulphide.

Red phosphorus is a dark brownish-red powder, insoluble in carbon bisulphide, and only ignites when strongly heated. It is non-poisonous. It need not be kept in water.

Phosphorus and oxygen.—When either ordinary or red phosphorus is burnt in air or oxygen a compound called *phosphorus pentoxide* is formed. It is a white amorphous powder which dissolves very readily in water, to form an acid solution known as *phosphoric acid.* Phosphoric acid may be obtained directly from phosphorus by oxidising it with nitric acid.

Phosphoric acid forms salts called *phosphates.*

Phosphorus trioxide is formed to a small extent when phosphorus burns in air, the quantity increasing if the supply of air is insufficient. It is a white solid with a garlic-like odour. With water it forms phosphorus acid.

Manufacture of phosphorus.—Phosphorus is usually prepared from the residue obtained by burning bones. This residue is a phosphate of calcium. By treating this compound with sulphuric acid it is converted into phosphoric acid. After concentration the phosphoric acid is heated with powdered coke in cast-iron retorts, when phosphorus distils over and is collected by water.

Nitric acid, sometimes called *aqua fortis,* is prepared by distilling potassium nitrate or sodium nitrate with strong sulphuric acid.

Being an acid it has a sour taste and turns a blue litmus paper red. It is colourless when pure and is decomposed easily when heated. It acts violently on most metals. Its great activity is due to the readiness with which it gives up oxygen.

Ammonia.—When the gas, ammonia, is dissolved in water, the solution which is produced acts as an alkali. The salts which it forms are known as ammonium salts, *e.g.* ammonium sulphate. When the solution of ammonia in water is heated the dissolved gas is again driven off and can be dried and collected.

The gas itself is colourless, has a pungent odour, and is exceedingly soluble in water. It burns in oxygen, but will not support combustion.

QUESTIONS ON CHAPTER XXIX.

1. There are two forms of phosphorus. What are the characteristic differences between them? How would you show their identity? Of what industrial importance are the two forms of this element?

2. What is formed when phosphorus burns? Give an account of the appearance and properties of the product.

3. What are phosphates, and how may they be obtained?

4. Describe, with diagram, an experiment to find the relative weights of phosphorus and oxygen which combine during the combustion of phosphorus (see also Chap. XXIX.).

5. For what purposes may the higher oxide of phosphorus be used in chemical work, and to what property is its use due?

6. Give a short account of the mode of production of phosphorus. For what purposes is phosphorus chiefly employed?

7. You are provided with some nitre and strong sulphuric acid; describe fully how you would proceed to prepare a specimen of nitric acid.

8. Why is nitric acid said to be an oxidiser? Describe an experiment to show its oxidising power.

9. State the composition and principal properties of ammonia. What happens when it is mixed (*a*) with water, (*b*) with nitric acid?

10. You are given a number of cylinders containing ammonia gas. Describe the experiments you would make in order to illustrate the chemical and physical properties of the gas.

11. Describe experiments which could be performed in order to demonstrate the points of difference between an aqueous solution of ammonia and of caustic potash respectively.

CHAPTER XXX.

CALCIUM CARBONATE AND SULPHATE; SILICA.

114. CALCIUM CARBONATE AND SULPHATE.

Occurrence of carbonates.—We have already seen that chalk consists of calcium carbonate. The production of lime from this source, together with its reconversion into calcium carbonate, have been studied already in Chapter XXVI. Calcium carbonate occurs naturally in several distinct forms, and many rock masses are largely or entirely built up of this compound. When pure it occurs crystalline as the minerals **calcite** and **aragonite,** which differ only in their crystalline forms and physical properties. Calcite is also known as *Iceland spar, calc spar,* and by other local names. It is generally quite transparent and somewhat resembles quartz, from which it can be distinguished by its inferior hardness. It is easily scratched by a knife, while quartz is unaffected. It is what is called a *doubly-refracting* substance. If a clear crystal of Iceland spar be placed upon the page of a book and the print viewed through it, two images of each word will be seen. Calcium carbonate also occurs more or less pure in the earth's crust in a great variety of forms, such as **chalk, limestone, stalactite, stalagmite, travertine,** etc., some of which have been formed by purely chemical means, others by the aid of living organisms.

Those formed by chemical means owe their formation to the power, possessed by water containing dissolved carbon dioxide, of dissolving calcium carbonate, which is again precipitated when carbon dioxide escapes.

Travertine or calcareous tufa is precipitated by springs which lose their dissolved carbon dioxide, which is necessary for the

solution of the calcium carbonate, as they flow onwards. The carbonate, being insoluble in water alone, is deposited as soon as the carbon dioxide escapes.

Stalactites and stalagmites.—The streams traversing limestone districts become saturated with carbonate of lime. In their course they often trickle through crevices in the roofs of caverns which have been formed in the limestone by the solvent power of water containing carbon dioxide. The drop of water

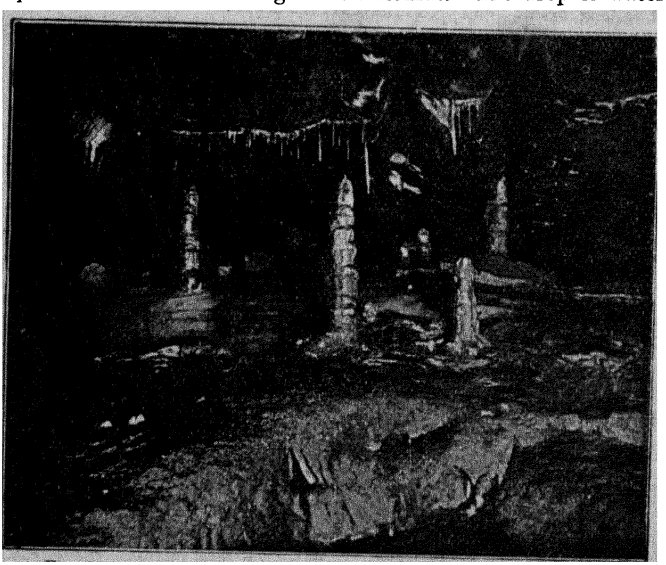


FIG. 255.—Stalactites and stalagmites in Clapham Cave, Yorkshire. (From a photograph by Mr. George Fowler.)

which is thereby exposed on the roof, is subjected to evaporation, and the escape of the carbon dioxide and loss of water cause a slight deposition of carbonate on the roof, which is continuously added to by a constant succession of drops, until eventually beautiful pendants of calcium carbonate are formed, called *stalactites*, sometimes coloured by the presence of traces of iron oxide, and often having a wild profusion of forms (Fig. 255). Further evaporation of the water takes place on the floor of the cavern, giving rise to layers of the same chemical

compound, called *stalagmites*. Chalk, limestone, and coral consist almost entirely of the remains of animal organisms.

Marble, also, is a form of calcium carbonate. It is really limestone, which, by the effect of great pressure and heat in the earth's crust, has become much harder and at the same time crystalline.

Calcium sulphate is a compound which occurs in various forms including **alabaster**, **selenite**, and **satin spar**. The first of these somewhat resembles marble but is softer. It is white or slightly tinted, and when made into vases and other ornaments, and therefore polished, it is seen to be semi-transparent or *translucent*. Selenite is found in transparent crystals, or in masses which at their edges show the leaf-like or foliated structure of the mineral. The commonest form of calcium sulphate, known as **gypsum**, occurs in irregular masses, like flint, and is a valuable mineral.

When either of these substances is heated it loses water of crystallisation and a white powder known as **plaster of Paris** is obtained. This substance is the well-known material used for mouldings, ceilings, statuary, etc. When mixed with water to the consistency of cream and left, it rapidly sets hard, heat being developed at the time of adding the water.

Calcium sulphate dissolves in water to a small extent, producing permanent hardness in the water. It is also used in paper making.

115. SILICA.

Silica.—More widely spread throughout the earth's crust even than calcium carbonate is the compound silica, which consists of the oxide of an element termed **silicon**. The element itself is only of importance on account of the compounds it forms. This oxide enters into the formation of by far the greater number of minerals and rocks, being combined with metallic oxides in the form of **silicates**; indeed, in the free state or thus combined, silica forms more than half the weight of the earth's crust. In the pure state it occurs both crystalline and amorphous. Two crystalline varieties are known, one, *tridymite*, is unimportant, while the other, **quartz**, is a frequently-occurring and highly-interesting mineral. If the quartz is

quite clear and transparent, it is known as **rock-crystal**, and is the *Brazilian pebble* from which lenses are made for eye-glasses, etc. Sometimes the oxide of a heavy metal is present, colouring the quartz.

Many **sands** are made up entirely of grains of quartz which have become more or less rounded by continual rubbing against one another in water. If the sand becomes compacted by the introduction of a cement and by the action of great pressure a **sandstone** is formed.

Amorphous Silica is found in the form of chalcedony and its varieties, jasper and its varieties, and opal.

Chalcedony is known having all sorts of tints. It is often regarded as a mixture of quartz and opal; it is familiar as the well-known red stone used in signet-rings and called **carnelian**. **Agate** is a variegated chalcedony, composed of different coloured bands. **Flint** is generally of a black or dark grey colour, and is found in nodules or bands in the chalk formations of Surrey, Kent, etc.

Jasper is an opaque, impure form of silica, of a red, brown, or yellow colour.

Opal.—This species of naturally-occurring silica always contains water in varying amounts. It is supposed to contain some quartz as well as amorphous silica. It is often used in jewellery; one variety shows a well developed iridescence, and is called **precious opal**.

CHIEF POINTS OF CHAPTER XXX.

Calcium carbonate is very largely found in the earth's crust. It occurs pure and crystalline in *calcite* and *aragonite* and in a less pure condition in *chalk*, *limestone*, *stalactites*, *stalagmites*, *travertine*, etc.

Chalk, limestone and coral are composed chiefly of calcium carbonate, and consist almost entirely of the remains of animal organisms.

Marble also is a form of calcium carbonate. It is really limestone which by the effects of great pressure and heat in the earth's crust has become hard and crystalline.

Gypsum or sulphate of calcium occurs under the forms of *alabaster*, *satin spar*, etc. When heated it yields a white powder known as *plaster of Paris* which sets hard in a few minutes on being mixed with a little water.

Silica is even more widely found than calcium carbonate. It constitutes, either free or combined with bases to form silicates, about one-half the weight of the earth's crust.

In the free state it occurs both crystalline and amorphous. Two crystalline varieties are known, one, *tridymite*, is unimportant; the other, **quartz**, is very abundant.

Many *sands* are made up entirely of grains of quartz.

Amorphous silica is found in three forms (1) chalcedony, (2) jasper, (3) opal. Carnelian, agate, and flint are common kinds of chalcedony.

QUESTIONS ON CHAPTER XXX.

1. What is silica? How does it occur in nature?

2. How do flint and quartz differ from and resemble one another?

3. Give a short account of the naturally-occurring forms of silica.

4. What is calcite? Describe its appearance.

5. What are stalactites and stalagmites? Of what do they consist and how are they formed?

6. Describe three naturally-occurring forms of calcium carbonate.

7. What effect has chalk upon water which flows through or over it?

8. Mention three minerals consisting of silica, and state by what characteristic feature you would recognise them.

9. How would you attempt to prove that limestone, marble, and calcite were essentially similar compounds? [The student is advised to consider well what experiments he has previously carried out with calcium carbonate.]

10. Limestone is dissolved in hydrochloric acid and the solution evaporated. Briefly describe the appearance and properties of the substance which you would expect thus to be obtained.

11. Describe the naturally-occurring forms of calcium sulphate.

12. What are the properties and uses of plaster of Paris?

CHAPTER XXXI.

SODIUM, POTASSIUM, AND THEIR COMPOUNDS.

116. SODIUM AND ITS COMPOUNDS.

i. Action of sodium on water.—Refer to Expt. 94 iv. p. 334.

ii. Glauber's salt or sodium sulphate.—Heat some crystals of this substance and notice the large quantity of water driven away. Prepare a hot saturated solution of sodium sulphate, pour it into a flask and allow it to cool while being kept perfectly still. Drop a small crystal into the cold solution and notice the sudden crystallisation, and at the same time the production of a large amount of heat.

iii. Caustic soda or sodium hydroxide.—Throw some lime into cold water till on stirring a milky liquid is obtained. Pour some of this liquid into a dish and add a few crystals of washing soda. Boil the mixture of lime and washing soda for five minutes and allow the muddy liquid to settle. Pour off the clear solution and evaporate it to dryness. A hard white solid, sodium hydroxide, is obtained. Notice that when it is left exposed to the air it soon absorbs moisture. Dissolved in some water the sodium hydroxide produces a solution which seems soapy when touched; it tastes very bitter.

iv. Washing soda.—Heat some crystals of washing soda and notice the large quantity of water driven away. Test some solution of washing soda in water with litmus paper. Examine the action of dilute acid upon washing soda.

Sodium.—This metal has already been referred to in two experiments. It is a soft silvery-coloured metal which so rapidly rusts or oxidises on exposure to air that it must be kept in bottles under mineral oil which contains no oxygen. It is easily cut with a knife. It burns in air or in oxygen with a yellow flame. Sodium combines with other elements so powerfully that it is only set free with difficulty. Davy used a strong electric current to electrolyse some melted caustic soda. Sodium

is now prepared by the electrolysis of a fused mixture of common salt with other chlorides.

Sodium is a very widely occurring element. Its commonest compound is that with chlorine or common salt, see Chap. XXVII. Combined with nitric acid it forms Chile saltpetre found in Chile and Peru (see p. 408), and it is also a constituent of some kinds of rock.

Common salt.—See Chap. XXVII.

Glauber's salt, or sulphate of soda, is made by heating common salt with sulphuric acid. It is the substance remaining in the flask when hydrochloric acid is prepared, as on p. 367.

Glauber's salt contains no less than 56 per cent. of water of crystallisation, which it gives up readily when exposed to the air, a white powder forming on the colourless transparent crystals. When heated the crystals melt, and at 100° lose all their water and become converted into the white *anhydrous* powder.

Sulphate of soda is found in some natural waters, and was formerly much used in medicine.

Washing soda, or sodium carbonate, crystallises in large transparent crystals containing nearly 63 per cent. of water. The crystals rapidly lose this water on exposure to air or when

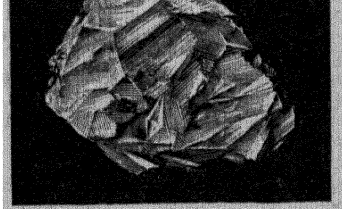


FIG. 256.—Fresh Crystals of Soda. (From a photograph by Mr. H. E. Hadley.)



FIG. 257.—Crystals of soda after exposure to the air. (From a photograph by Mr. H. E. Hadley.)

heated, and fall into a fine white powder. Washing soda is very soluble in water, 100 parts of which at 100° dissolve 45·5 parts of the crystals. When the fingers are dipped into the solution it is found to possess a soapy nature.

Washing soda changes litmus, previously reddened with acid, to its original blue colour. It is thus capable of neutralising acids, and in the process carbon dioxide escapes freely.

The manufacture of washing soda is a very great industry, known as the alkali manufacture. The explanations of the processes employed, of which there are two, are not easy to understand at this stage. It will be sufficient to say that in both cases the soda is made from common salt. In the first or *Leblanc* process, the salt is heated with sulphuric acid, and the sodium sulphate so produced is strongly heated with a mixture of limestone and coal. The soda is dissolved out from the resulting substance, called *black ash*, by hot water. In the second and later process washing soda is made by passing carbonic acid gas and ammonia through a strong solution of common salt.

Sodium bicarbonate is made by passing carbon dioxide gas into a solution of washing soda. It forms transparent colourless crystals which produce a cooling effect upon the tongue. The crushed crysta's are generally sold, as a white powder, for making effervescing drinks. When strongly heated the bicarbonate gives off carbon dioxide gas and is changed into the ordinary carbonate.

Caustic soda, also known as sodium hydrate or sodium hydroxide, is produced when sodium is placed in water (p. 334). The solution then obtained, if evaporated, becomes thicker and oily in appearance. If sufficiently strong the solution will crystallise. In practice caustic soda is made by a much cheaper process. Washing soda and quicklime are boiled together and then allowed to settle. The clear liquid above the white mud (consisting of chalk) is a solution of caustic soda. This solution is concentrated and poured into moulds to solidify into the sticks in which form it is sold.

Caustic soda is a strongly alkaline substance which eats into or "burns" away the skin. It rapidly absorbs moisture from the air, finally becoming liquid if not preserved in a closed bottle. It also strongly absorbs carbon dioxide gas from the air and gradually becomes coated with a layer of sodium carbonate.

Caustic soda is used in the manufacture of soap, paper, and starch, and for many laboratory purposes.

117. POTASSIUM AND ITS COMPOUNDS.

i. Action of potassium on water.—Repeat Expt. 94 iv., p. 334, using potassium in place of sodium. Similar results will be obtained, but the production of hydrogen is more energetic and *the experiment requires still more caution.*

ii. Potassium chlorate.—Refer to Expts. 91 i. ii. and iii., pp. 323-4, for experiments with this substance.

iii. Caustic potash or potassium hydroxide.—Repeat Expt. 116 iii., p. 404, using potassium carbonate in place of washing soda. Notice the very close resemblance between sodium and potassium hydrates.

iv. Nitre, saltpetre, or potassium nitrate.—This substance has already been used for the preparation of nitric acid. We may show that it contains a large quantity of oxygen in several ways. (*a*) Heat some crystals of nitre in a test-tube and, when melted, drop into the liquid a small piece of charcoal. The charcoal burns with great energy. The experiment may be repeated with a small piece of sulphur with similar results.

(*b*) Prepare a saturated solution of nitre in water. Dip a filter paper into the solution and dry it by waving the paper over a lighted Bunsen burner. When dry apply a light to the edge of the paper. It will be found that the paper burns away quickly as far as it has been dipped into the solution, a sparkling line of fire taking the place of a flame. Paper thus prepared is called *touch-paper.*

v. Potassium permanganate.—(*a*) Drop two or three crystals into a test tube of water and notice the fine purple colour produced as the crystals dissolve while falling through the water.

(*b*) Heat a small quantity of the solid permanganate in a test tube and test for oxygen escaping by means of a smouldering cedar splint.

Potassium.—This metal resembles sodium very closely in all its properties. The chief difference, so far as the elementary student is concerned, lies in the fact that because of its stronger attraction for oxygen it is still more difficult to obtain than sodium. It was discovered in the same way as sodium and is also prepared on the large scale by a similar method.

Potassium is a soft bright metal which tarnishes even more rapidly than sodium. It is therefore preserved under mineral oil or naphtha which contains no oxygen. It may be easily distinguished from sodium by the fact that when thrown on water the heat produced is sufficient to set fire to the hydrogen set free. The hydrogen flame is tinged with violet by the presence of the vapour of potassium.

Potassium carbonate is the chemical name for the substance obtained from wood ashes, whence its old name of *pot-ashes.*

The ashes are **lixiviated** with water, that is, water is allowed to dissolve as much as possible of the ashes and the clear liquid drawn off and evaporated. At the present time most of the potassium carbonate is obtained from Stassfurt where are found large deposits of **sylvine** (a double chloride of potassium and magnesium), and of **kainite** (a complicated compound of sulphates and chloride).

The pure carbonate forms colourless crystals which, on heating, yield water and a white anhydrous powder which rapidly absorbs moisture from the air.

Potassium chlorate.—This substance has been already referred to in Chapter XXIV. for the preparation of oxygen, and also in Chapter XXVII. under "chlorine compounds."

Caustic potash or potassium hydroxide resembles caustic soda very closely except that it absorbs moisture from the air somewhat more readily. It is produced when potassium is thrown into water. Like caustic soda it must be kept from the air since it has a very strong attraction for carbon dioxide gas and forms with it potassium carbonate. It is sold in sticks and is only distinguished from caustic soda by careful tests. Caustic potash is made by boiling a solution of potassium carbonate to which lime is gradually added (cf. caustic soda, p. 406). It is used for the same purposes as caustic soda, such as soap making, etc.

Saltpetre, or nitre, or potassium nitrate occurs in the soil in India and other countries. Chile saltpetre or sodium nitrate, from which most of the potassium nitrate is obtained, is found in Bolivia and Peru.

From the soil the potassium nitrate is prepared by simply stirring it with water, concentrating the clear solution and crystallising. To obtain it from the sodium nitrate, potassium chloride is added to the solution, when potassium nitrate and sodium chloride are formed. The former crystallises out before the latter, which is more soluble in water.

Potassium nitrate forms colourless transparent crystals which, when heated, melt to a clear liquid and afterwards give off oxygen.

If some charcoal or sulphur be thrown on some melted nitre either takes fire and burns with great rapidity.

Nitre is chiefly used in making gunpowder and fireworks. The sodium nitrate cannot be used for this purpose because on exposure to the air it absorbs moisture and becomes wet.

Gunpowder is made by carefully mixing together 75 parts of nitre, 15 parts of charcoal, and 10 parts of sulphur. These substances are finely powdered, then mixed by hand, and finally while wet, thoroughly ground together in a powder mill. The cake thus obtained is broken up between copper plates, the pieces are subjected to pressure, which produces granulation, and finally are glazed by being placed in revolving drums.

The power of gunpowder arises from the sudden formation of a very large volume of gas consisting of carbon dioxide and nitrogen.

The composition of gunpowder may be ascertained by first dissolving out the nitre with boiling water, the powder being placed on a filter paper. The sulphur is next extracted by means of carbon bisulphide, and the carbon or charcoal remains.

Potassium permanganate is manufactured by heating together the black oxide of manganese, potassium chlorate, and caustic potash. The hard dark-coloured residue is boiled with water and allowed to settle. The clear liquid is evaporated and crystallised.

Potassium permanganate forms small dark reddish-purple crystals which are soluble in water, producing a very deeply coloured solution. This solution is much used in the laboratory on account of the ease with which it liberates oxygen either on being heated alone or with sulphuric acid. It is a powerful deodorant and disinfectant. Sodium permanganate, which closely resembles it, is known as Condy's Disinfecting Fluid.

CHIEF POINTS OF CHAPTER XXXI.

Sodium is a soft metal having a powerful attraction for oxygen. It must be preserved under mineral oil. It decomposes water setting the hydrogen free. It burns in oxygen or in air with a yellow flame.

Common salt.—See Chapter XXVII.

Glauber's salt or sodium sulphate is made by heating together a mixture of common salt and sulphuric acid. It is chiefly produced in the manufacture of washing soda and is remarkable for its solubility in water and for the large amount of water of crystallisation it contains.

Washing soda or sodium carbonate forms large colourless crystals which are very soluble in water and contain a large amount of

water of crystallisation. It neutralises acids, carbon dioxide being set free. It is made in immense quantities in the *alkali industry*.

Caustic soda or sodium hydroxide is a very powerful alkali. It forms a soapy liquid when dissolved in water which burns the skin, whence its name. It is largely used in the manufacture of soap, paper, etc., and is a useful substance in many laboratory processes.

Potassium closely resembles sodium. It acts violently on water and the hydrogen it sets free takes fire.

Potassium chlorate.—See Chapter XXVII.

Caustic potash or potassium hydrate is so much like caustic soda that it is not readily distinguished from that substance.

Saltpetre, or nitre, or potassium nitrate occurs in the soil in hot countries. It is a crystalline salt having a cool taste. It is chiefly used for making gunpowder because of the large quantity of oxygen which it contains.

Gunpowder is a mixture of nitre, charcoal, and sulphur. It owes its power to the large volume of carbon dioxide gas and nitrogen suddenly set free when it burns.

Potassium permanganate is a dark reddish-purple crystalline substance which contains and easily liberates a large quantity of oxygen. It is this oxygen which makes it a powerful disinfectant when dissolved in water. In solution it forms a deep violet-coloured liquid.

EXERCISES ON CHAPTER XXXI.

1. Describe what happens when (*a*) sodium, (*b*) potassium are thrown into water. What substances are produced in each case ?

2. Give the chief properties of Glauber's salt.

3. Write a short account of the manufacture and properties of washing soda.

4. For what purposes is caustic soda employed ?

5. How could you clearly distinguish between specimens of potassium carbonate and potassium chlorate ?

6. For what purpose is saltpetre chiefly used ? What happens when saltpetre is strongly heated in a tube and a piece of sulphur is thrown into the heated substance ?

7. Suggest a method of ascertaining the percentage of saltpetre in a sample of gunpowder.

8. How can you obtain oxygen from potassium permanganate ? What are the uses of this substance ?

CHAPTER XXXII.

SOME COMMON METALS.

118. GENERAL CHARACTERS OF METALS.

i. Some familiar metals.—(*a*) Examine specimens of lead, copper, iron, zinc, tin, silver, gold, sodium, and mercury. Notice that while most are heavy solids, sodium floats on water, and mercury is a liquid.

(*b*) Observe they are all opaque and possess a "metallic lustre." Compare with crystals of iodine to satisfy yourself that some non-metals possess a lustre. Examine a very thin sheet of gold between two glass plates; it is transparent.

ii. All metals are not elements.—Examine brass, gun-metal, pewter, German-silver, bronze, solder. These are *alloys* or mixtures of metallic elements. The metallic characters are not lost by mixing metals together.

Metallic elements.—Though most people have some idea of what a metal is, it is a somewhat difficult matter to say exactly what constitutes a metal. The fact is, that the elements cannot be sharply divided into two classes with all those that possess metallic properties in one division, and those in which such characteristics are absent in the other. It is an easy thing to enumerate the respects in which, say, the metal gold and the non-metal oxygen differ from one another, but there are many elements, like arsenic, which possess properties common to both typical metals and characteristic non-metals.

But, as a rule, metals may be said to possess the following properties, though the student must remember that exceptions are known to nearly every one of the statements.

Characters of metals.—1. Metals possess a peculiar lustre, which is commonly referred to as "metallic." But iodine and graphite, which are undoubtedly not metals, also have a lustre very like metals.

2. Metals are opaque. Yet gold can be rolled into sheets which are transparent to light.

3. Metals are very dense, or possess a high specific gravity. Sodium and potassium are unmistakable metals, yet they float upon water.

4. Metals, as a rule, are good conductors of heat and electricity.

5. Metals unite with oxygen, or with oxygen and hydrogen, to form bases, which neutralise acids.

Alloys.—All metals are not elements. Many of the familiar metals in common use are mixtures of metallic elements called **alloys.** When one of the metals present is mercury the alloy is known as an **amalgam.** The following table shows the metals present in some common alloys:

Alloy.	Metallic Elements Present.
Bell-metal, .	Copper and tin.
Brass, . .	Copper and zinc.
Bronze, . .	Copper and tin with some zinc and lead.
German silver,	Copper and nickel.
Gun-metal, .	Copper and tin.
Pewter, . .	Tin and lead.
Solder, . .	Tin and lead.
Type-metal, .	Lead, tin, and antimony.

119. LEAD.

i. **Properties of lead.**—(*a*) Scrape a piece of lead and examine its metallic surface; incidentally notice that it is much softer than steel.

(*b*) By the method described on p. 60, determine the relative density of lead.

(*c*) Examine lead in the forms of sheets and wires. What properties must lead possess in order to be available in these forms?

ii. **Heating lead in air.**—(*a*) Melt lead in an iron spoon and pour the liquid metal into a mould made in sand.

(*b*) Melt more lead and keep the liquid metal stirred with an iron wire. Observe that it combines with the oxygen of the air to form yellow oxide of lead (litharge).

iii. **Lead obtained from its compounds.**—(*a*) Strongly heat some red lead on a piece of charcoal by means of a blow-pipe. Notice the globules of the metal which separate.

(*b*) Heat a mixture of lead oxide and powdered charcoal in a

crucible by means of the flame of a foot blow-pipe. Again notice the separation of metallic lead.

iv. Lead is soluble in nitric acid.—Upon some fragments of lead in a test-tube pour some moderately strong nitric acid. Notice the brown fumes and the gradual solution of the lead. When the lead has dissolved, add water to the solution and gently evaporate on a sand-bath till nearly dry. Place on one side for a short time and notice the salt, lead nitrate, which separates in crystals.

Lead.—Lead has a bluish colour, which can be seen by examining an untarnished surface of the metal. The bright metal, however, soon tarnishes on exposure to air. It is 11½ times as heavy as water. It is very malleable and fairly ductile. It melts at 326° C. into a silvery fluid, which can be easily cast in moulds. Melted lead combines with the oxygen of the air to form yellow **litharge**. If this be strongly heated it will combine with a further quantity of oxygen to form the **red oxide**. Though it easily dissolves in nitric acid, forming lead nitrate (p. 394), hydrochloric acid and sulphuric acid have little, if any, action upon it.

In nature, lead occurs combined with the element sulphur as the ore **galena**, which is lead sulphide. But many other ores are known.

The metal is made into sheets and pipes and extensively employed by the plumber. Alloyed with other metals it occurs as pewter, solder, and type-metal (p. 412).

120. IRON.

i. Properties of iron.—(*a*) Determine the relative densities of steel, wrought-iron, and cast-iron by the methods described on p. 62.

(*b*) Recall and, if necessary, repeat the experiments dealing with the rusting of iron on p. 307, that concerning the burning of iron in oxygen on p. 326, and the action of acids upon it.

(*c*) Revise what has been studied of the magnetic properties of the metal (p. 243).

Iron.—Iron is by far the most important metal known to man. The discovery of how to obtain it from the minerals in which it occurs was probably the most valuable ever made. Though it is very abundant in nature, it is rarely found uncombined. It has been found in certain strange masses, called meteorites, which drop upon this planet from inter-stellar

space. In the earth it is found combined with oxygen, forming oxides, as **magnetite** and **haematite**; as oxides combined with water in *limonite* and *gothite*; with sulphur, as the sulphides, **iron pyrites** and *magnetic pyrites*; with carbon dioxide, as the carbonates, **clay iron-stone** and *chalybite*; and other less important ores.

Three kinds of iron.— Iron is known and used in three forms, **wrought-iron, cast-iron,** and **steel.** The first is almost the pure element, but cast-iron contains also varying amounts of carbon and silicon. Steel contains the same elements as cast-iron, but the amount of carbon is considerably smaller. The different uses to which these varieties of iron are put depends upon the difference in properties they possess.

Wrought-iron is very tough, and can easily be beaten out into plates. For those articles which are made by hammering the iron into shape, wrought-iron will evidently be used.

Cast-iron is, on the other hand, brittle and easily melted, and consequently is employed in all cases where the article is made by running the molten metal into moulds.

Steel has different properties according to the processes through which it has passed. If it has been heated and then cooled very quickly, it is extremely hard but very brittle: but if cautiously heated and cooled very much more slowly, it is no longer brittle but elastic. This latter process is called **tempering.** Another very important property which steel possesses is that it can be made into a magnet which will keep its magnetism for a very long time. All the magnetic needles used in telegraphing, and in electrical instruments of other kinds, are made of steel.

The relative densities of different kinds of iron are seen from the following table:

Steel, not hammered.	7·82
Iron, bar.	7·79
Iron, cast, ...	7·21

Oxides of iron.—Iron combines with oxygen in several proportions. The following table shows the proportions of the masses of iron and oxygen present in each case:

Ferrous oxide, . .	56	of iron with	16	of oxygen.
Ferric oxide, . .	112	,,	48	,,
Tri-ferric tetroxide, .	168	,,	64	,,

The first of these does not occur in nature as a mineral. It combines with acids, forming the series of salts known to chemists as *ferrous* salts, one of which, **ferrous sulphate** (green), is known naturally as the mineral **copperas.**

The second, ferric oxide, is fairly abundant in nature. It constitutes the beautifully crystallised mineral **speculor iron ore,** found in Elba. It also makes up the mineral *haematite*, which goes under the names of **kidney ore** and **pencil ore** in the Furness district of Lancashire, according to the shapes which it naturally assumes.

121. COPPER.

i. Properties of copper.—(*a*) Examine copper in the form of bars, sheets, and wires. Notice its colour. From these varieties what properties of the metal can be adduced ?

(*b*) Hold one end of a copper wire between your finger and thumb and put the other in the flame of a laboratory burner You will soon be convinced that copper is a good conductor of heat.

ii. Heating copper in air.—(*a*) Heat a piece of sheet copper over the flame of a laboratory burner. Notice the formation of a black film. This is the black oxide of copper.

(*b*) Refer to the experiment on p. 344, and recall the action of hydrogen gas upon this black oxide when the latter is heated.

iii. Action of copper on acids.—Try the action of the common acids upon fragments of copper placed in test-tubes.

iv. Iron can displace copper from its solutions.—Into a solution of blue vitriol (copper sulphate) plunge the blade of a knife. Observe the deposition of copper on the blade. The iron of the knife passes into solution and takes the place of the copper.

v. Alloys of copper.—Examine specimens of brass, bronze, bell-metal, and gun-metal. Compare them with copper.

Copper.—Copper is a red metal which is found naturally in an uncombined or " native " condition. It also occurs as **ruby copper,** which is an oxide of the metal, *cuprous oxide* ; and more abundantly as **copper pyrites,** a compound of the metal with iron and sulphur. **Copper-glance** and **malachite** are also well-known ores of copper.

Copper is hard and does not, when cold, change in dry air, though in moist air it slowly becomes covered with a green compound, due to the combination of the metal with carbon dioxide and water. When it is heated in the air copper combines with oxygen to form black oxide of copper.

The metal is nearly nine times as heavy as water. It is very ductile, and can be drawn out into very thin wires; its great malleability enables it to be rolled into very thin sheets, which are known as Dutch leaf. It is a good conductor of heat and of the electric current. Everybody is familiar with the use of copper wires to conduct electric currents from one place to another.

Moderately strong nitric acid dissolves copper forming with it a blue liquid, copper nitrate. Hydrochloric acid does not act upon the metal, nor does dilute sulphuric acid, but when copper is heated with strong sulphuric acid the metal is dissolved with the formation of copper sulphate, sulphur dioxide being evolved. The alloys of copper are both numerous and useful. As an examination of the table on p. 412 shows, copper is mixed with; tin in varying proportions to form bell-metal and gun-metal; with zinc to make brass; with nickel in the manufacture of German-silver; and with tin, zinc, and lead to produce bronze.

122. MERCURY.

Properties of mercury.—(*a*) Determine the relative density of mercury by means of a density bottle as described on p. 62.

(*b*) Using the apparatus employed in the experiments described on p. 63, show that a column of mercury 1 inch high balances a column of water between 13 and 14 inches high. The relative density of mercury is consequently between 13 and 14.

(*c*) Satisfy yourself by trial, (1) an iron key floats on mercury, (2) mercury does not wet glass, (3) mercury adheres to clean zinc or copper, forming what is called an *amalgam*.

(*d*) Boil a little mercury in a test-tube and show that it is volatile. Some of the mercury condenses in tiny drops on the cool parts of the tube. Take care not to breathe any of the vapour.

(*e*) If necessary, repeat the experiment of heating red oxide of mercury (p. 321).

Mercury.—Mercury or quicksilver is the only metal which is liquid at ordinary temperatures. Its appearance is familiar to everyone from its frequent use in barometers and thermometers. It sometimes occurs pure in nature. Chiefly, however, it occurs in combination with sulphur as the mineral **cinnabar** (*red*), which is found in Spain, Hungary, Tuscany, and South America. If mercury is cooled to a temperature of $-40°$ C.

it solidifies, and is then malleable. It is the heaviest liquid known, being 13½ times as heavy as water.

If it be heated to a temperature of 315° C. and air be passed over it, it combines with oxygen, forming red oxide of mercury. At ordinary temperatures the oxygen of the air has no action upon it and the metal does not tarnish by simple contact with air. It boils at 357·5°, and is converted into a transparent, colourless vapour, which is very poisonous. It dissolves many metals, *e.g.* zinc and copper, forming alloys known as **amalgams**.

Not only is it used in barometers and thermometers, but also in the manufacture of looking-glasses and in the laboratory, instead of water, over which to collect some gases soluble in water.

Hot concentrated sulphuric acid dissolves mercury, the action being exactly similar to that in the case of copper (p. 382). Nitric acid, too, dissolves it readily.

123. ZINC.

i. **Properties of zinc.**—(*a*) Examine some strips of zinc and make out as many properties as you can.

(*b*) Determine its relative density.

(*c*) Heat some pieces of zinc in an iron ladle and pour the liquid metal, drop by drop, into a bucket of water. The metal is cooled in the form of granulated zinc.

ii. **Action of acids upon zinc.**—(*a*) Re-read and, if necessary, repeat Experiment 95 i.

(*b*) Pour a little dilute nitric acid upon some fragments of zinc in a test-tube. Warm the solution. Filter and evaporate to dryness—zinc nitrate is thus obtained.

Zinc.—Zinc occurs in nature combined with sulphur in the form of zinc sulphide, making the well-known ore **zinc blende**, or *black jack*. The carbonate of zinc occurs as the ore **calamine**. Zinc is a bluish-white metal about seven times as heavy as water. It does not readily combine with the oxygen of the air and is extensively used to coat iron-plates to prevent their rusting. These sheets are known as **galvanised iron**. Zinc is also one of the constituents of brass. When zinc is heated strongly in air it readily combines with the oxygen, the combination being accompanied by a greenish flame.

As has been seen in studying hydrogen, zinc is easily dissolved

by both sulphuric and hydrochloric acids, the hydrogen of the acid being evolved, and zinc sulphate, and zinc chloride, being formed.

124. SILVER AND GOLD.

i. Properties of silver.—(*a*) Examine a piece of silver and notice as many of its properties as possible. Determine its relative density. Recall the uses of silver, and in this way remind yourself that silver does not tarnish, *i.e.* combine in ordinary circumstances with the oxygen of the air.

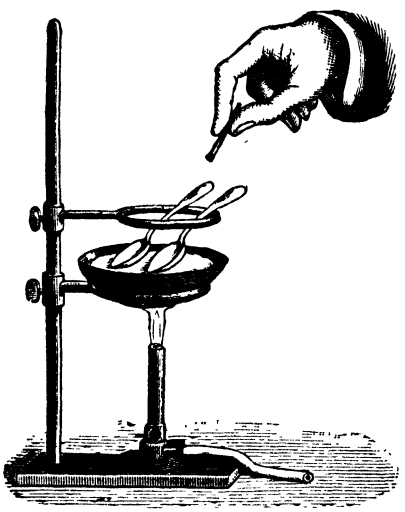


FIG. 258.—Simple method of showing that silver is a better conductor of heat than electro-plate.

(*b*) If possible, examine a sheet of thin silver leaf between two sheets of glass. Observe that when very thin it transmits blue light.

(*c*) Place a silver spoon and an ordinary electro-plated spoon upon a sand-bath, as in Fig. 258. Upon the end of each put the end of a wax vesta without any wax with it, or hold the head of a match at the end of each when they are hot. Heat the sand-bath by placing a Bunsen burner under it. The match on the silver spoon will take fire before that on the other spoon.

ii. Silver coins contain copper.—Dissolve a threepenny piece in moderately strong nitric acid. Notice the blue colour of the solution.

Refer to Experiment iii. and satisfy yourself that silver coins contain copper, and that silver is soluble in nitric acid.

iii. Action of light on silver compounds.—To a solution of silver nitrate add a little dilute hydrochloric acid, and notice the white precipitate which is at once formed. Filter the solution and so obtain the white precipitate on a filter paper. Leave it exposed to light, and note its gradual change of colour.

iv. Properties of gold.—(*a*) Roll up a sheet of gold-leaf and drop it into strong nitric acid contained in a test-tube. The gold is not dissolved even when the acid is warmed.

(*b*) Add strong hydrochloric acid to the nitric acid in the last experiment, in this way making what is called *aqua regia*. The gold is now dissolved.

Silver.—Silver is a white metal about ten and a half times as heavy as water. It does not tarnish when exposed to the air, even when heated. It is consequently much used for coinage and for ornamental purposes. It is, however, too soft to be used by itself, and is generally alloyed with copper. British coins contain about seven and a half per cent. of copper.

Silver conducts heat and the electric current more readily than any other metal. It is very malleable. When hammered into very thin leaves it is transparent to some constituents of white light, transmitting the light of wave-lengths corresponding to the blue end of the spectrum. Silver is also very ductile and can be drawn out into wires of exceeding fineness.

Acids act upon silver, producing similar effects to those which have been studied under copper. Thus, hydrochloric acid has no action on silver; nitric acid dissolves it, forming silver nitrate, and evolving oxides of nitrogen; and hot concentrated sulphuric acid dissolves silver, forming silver sulphate, and giving off sulphur dioxide.

Silver forms with sulphur a black sulphide of silver. This explains the blackening of silver in ordinary rooms lighted by coal gas. The coal gas nearly always contains slight traces of sulphuretted hydrogen, which act upon the silver with the formation of the black sulphide of the metal. Similarly, india-rubber contains sulphur, and if a silver coin is left in contact with a piece of india-rubber, the same blackening is noticed.

Several compounds of silver, notably the *chloride*, the *bromide* and the *iodide*, are blackened by exposure to light, a fact which forms the basis of photography.

Gold.—Gold is nearly always found native (p. 415) in nature, though it also occurs alloyed with other metals. Everybody is familiar with its bright yellow colour and with the circumstance that it is unacted upon by the air. It is more than nineteen times heavier than water. It is unacted upon by any of the common acids acting singly, though a hot mixture of strong hydrochloric acid and nitric acid will dissolve it. For this reason the mixture of these acids is known as **aqua regia.**

Gold is too soft in the pure state to be used either for coinage or for jewellery and is always alloyed with copper. This gives rise to the employment of the term *carat.* Pure gold is known as 24 carat gold. The British sovereign, which contains 22 parts of gold in 24 parts of the coin, is said to be made of 22 carat gold. Similarly, 9 carat gold consists of 9 parts of gold in every 24 parts of the article made of it.

It is the most malleable and most ductile metal known. Gold leaf has been made into sheets so thin that it would require more than a quarter of a million of them together to make a thickness of an inch. Gold wire of such an extreme thinness has been manufactured that two miles of it only possess a mass of 1 gram. Ordinary gold leaf, such as is used for gold lettering, is transparent to some constituents of the spectrum; it transmits green light quite readily.

125. PREPARATION OF SOME COMMON METALS FROM THEIR ORES.

Zinc.—The carbonate of sulphide when strongly heated or calcined in a current of air becomes changed into the oxide. The oxide is heated strongly with coal, which at a very high temperature removes the oxygen, leaving the zinc. The mixture of oxide and coal is heated in iron retorts and the zinc distils over into iron receivers.

Iron is obtained from its ores by reduction. A mixture of coal, limestone, and the iron ore previously calcined so that it consists chiefly of oxide, is strongly heated in a large furnace some eighty feet in height, and a supply of hot air is forced into the heated mass by means of tubes, called *tuyeres*, arranged near the base. The hot coal with the oxygen in the blast forms

carbon monoxide, a very powerful *reducing agent*. The reduction of the iron oxide takes place part of the way down the furnace and the iron gradually works downwards, starting as a pasty mass but becoming quite liquid at the bottom. Here it is found beneath a layer of melted slag, which consists of silica from the ore combined with the limestone. The melted iron is run off into branched grooves made in sand and on cooling is known as pig iron or cast iron.

Copper.—The smelting of copper ore is a complicated process. If the ore contains no sulphur it is heated strongly with coal or coke which removes the oxygen. If sulphur is present the operations are not so simple, for it is important that no sulphur be allowed to remain in the final product. The ore is roasted and melted repeatedly, and the mass of nearly pure melted copper is stirred finally with wooden poles till all trace of sulphur is removed.

Mercury.—The ores are roasted, the sulphur passes off as sulphur dioxide, the mercury distils over and is collected in a series of cool vessels.

Lead.—The ore galena, or lead sulphide, is first roasted in a stream of air. Part of the sulphide absorbs oxygen from the air and becomes sulphate. The temperature is now raised and the sulphate and the remaining unchanged sulphide act upon one another, the sulphur escaping as oxide. The lead now set free melts and flows to the bottom of the furnace.

Silver.—This metal is prepared by three processes. In the first mercury is mixed with the ore, after previously treating it with salt, and the two metals form an alloy from which the silver is obtained by distilling off the mercury.

In the second process the silver ore is heated with some lead ore. The lead and the silver together form an alloy. The lead is removed by melting and blowing air over the surface. Lead oxide is formed, melts on the surface of the heated mass, and is blown away by the blast of air, till at last the pure liquid silver appears.

In the third process the ore is heated carefully in a current of air until the silver is changed into sulphate. The sulphate of silver is dissolved in water and the silver is removed from the solution by placing in it some scrap copper upon which the silver is deposited.

CHIEF POINTS OF CHAPTER XXXII.

Metals.—Metals possess a peculiar lustre; they are opaque and very dense; they conduct heat and electricity well; they unite with hydrogen and oxygen to form bases.

Mixtures of metallic elements are called *alloys*.

Lead has a bluish colour; its density is 11·5; it is malleable and ductile; it melts at 326° C. With oxygen lead forms *litharge* and *red-lead*; with sulphur it forms *galena*. It is one of the constituents of pewter, solder, and type-metal.

Iron is the most important of metals. It is used in three forms, *wrought-iron*, *cast-iron*, *steel*. These forms have different properties. It is abundant in nature. *e.g.* combined with oxygen it forms *lodestone* and *haematite*; with sulphur, *iron pyrites*; with carbon dioxide, *clay ironstone*. Its density varies from 7·2 to 7·8.

Copper has a red colour, it is hard, and does not change in dry air. Its density is about 9. It is very malleable and ductile. It is a very good conductor of the electric current. Common minerals in which it occurs are *ruby copper*, *copper pyrites*, and *malachite*. It is dissolved by moderately strong cold nitric acid and is soluble in hot strong sulphuric acid. Numerous alloys of copper are used.

Mercury or *quicksilver* is the only liquid metal at ordinary temperatures. It is much used in thermometers and barometers. Its density is 13·5. It boils at 357·5° C. It dissolves many metals forming amalgams. Cold nitric acid and hot strong sulphuric acid dissolve the metal.

Sodium is lighter than water. Its density is 0·97. It has to be kept under naphtha because of its strong attraction for oxygen.

Zinc is a bluish-white metal. Its density is about 7. It does not easily combine with the oxygen of the air, and is extensively used to coat iron, as in *galvanised iron*. It is one of the constituents of brass. It occurs in nature combined with sulphur, and with carbon dioxide. The metal easily dissolves in dilute hydrochloric, sulphuric, and nitric acids.

Silver is a white metal. Its density is about 10·5. It does not tarnish in air, and when alloyed with copper is much used for coinage and jewellery. It is very malleable and ductile. The action of acids upon it is similar to the case of copper. Many compounds of silver are sensitive to the action of light, a fact which forms the basis of photography.

Gold is a bright yellow metal, which generally occurs native in nature. It is unacted upon by air, and is much used for jewellery and coinage. It is soluble in *aqua regia*. Pure gold is known as 24 *carat* gold. The British sovereign contains 22 parts of gold in 24 of the coin, and is called 22 carat gold. Gold is the most malleable and the most ductile of all metals.

EXERCISES ON CHAPTER XXXII.

1. What are the usual characters possessed by a metal? Name any exception you know to any of the properties you state.

2. What do you understand by each of the following terms: native gold, alloy, amalgam, carat?

3. Name three alloys in which copper occurs, and three in which lead is an important constituent.

4. What metals occur in each of the following minerals: cinnabar, haematite, galena, malachite, blende?

5. Enumerate the characteristic properties of lead and silver.

6. What metals do you know to be soluble in each of the following acids: hydrochloric, sulphuric, nitric?

7. What do you understand by galvanised iron? How is it made, and what reasons can you suggest for the process?

8. To what uses are the following metals put: mercury, gold, and zinc? On what properties of the metals do these uses depend?

9. Give a short account of the preparation from their ores of any two common metals.

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