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ELEMENTARY
PHYSICS AND CHEMISTRY

FOR THE USE OF SCHOOLS

BOOK II.
PHYSICS AND CHEMISTRY

BY

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PREFACE

THE "Specimen Schemes" of Instruction which were issued by the Board of Education with the New Code of 1900, include for schools, where facilities exist for performing simple experiments, a course in Elementary Physics and Chemistry. This book is the second of three intended to serve as reading-books as well as lesson-books in this subject for such schools.

Whilst the subject has been treated strictly on scientific lines, no pains have been spared to make that treatment as simple as possible. And whilst recognizing that very few schools are provided with laboratories for individual experiment on the part of the scholars, the book assumes that there are very few which are not able to command the few simple pieces of apparatus needed to carry out, on the part of the teacher, the easy experiments herein described. Many of them the scholars will delight in repeating on their own account at home.

The author has had many years' experience in a mixed school where girls have undergone such a course as this before, or concurrently with courses in Cooking and Domestic Economy, Botany, Physiology and Hygiene, and with the best results. Seeing that without a foundation of pure science such applied subjects as those named must be treated in a purely empirical manner, he does, without hesitation, urge mistresses of girls' schools to give their scholars the benefit of a sound foundation in Elementary Physics and Chemistry.

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BOOK II.

ELEMENTARY PHYSICS AND CHEMISTRY

LESSON I.

THE SKIN OF LIQUIDS.

Apparatus needed:—Elastic thread and small ring ; basin, matches, and piece of soap ; fine glass tubes ; mercury ; soap solution ; various wire rings and frameworks ; silk thread ; clay pipe ; candle.

You have already been shown in Book I. how that, if you have two liquids which do not mix, such as oil and water, the less dense liquid floats on the surface of the denser—not as a solid does, partly immersed in it, but entirely on it, the surface which separates one from the other being a horizontal plane. This is because of their mobility.

But if you have two liquids of equal density which do not mix, the one liquid will float in the other, not on it, and there will be no flat surface of separation. The smaller quantity of liquid will gather itself together into a ball because of the mutual attraction of its molecules—that is, their cohesion.

Now let us consider the molecules in a vessel of water or any other liquid. If we think of any molecule, which we will call A, in the body of the liquid, we see that it is surrounded on all sides by other molecules, which attract it; and because the attractions of the surrounding molecules are equal and in every possible direction, the molecule A is not set in motion.

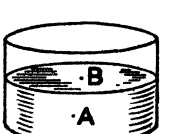


FIG. 1.

But a molecule, B, at the surface of the liquid has not molecules all round it. Those underneath it attract it, but it has no molecules of liquid above it to counterbalance their attraction, so that it is pulled downward a little towards the interior of the liquid. And this is the case with all the surface molecules.

It is somewhat like this. Suppose you had half a dozen pieces of elastic thread tied to a ring and stretched out like the spokes of a wheel, equal distances apart. So long as the pieces of elastic thread were all equally stretched the ring would be

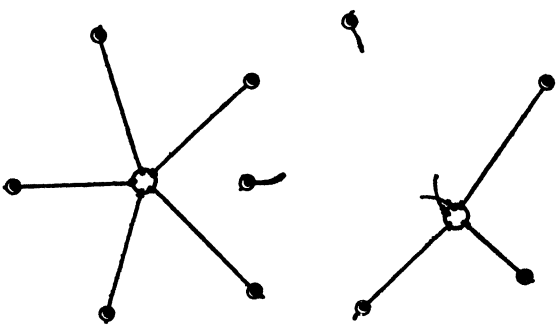


FIG. 2.

kept in the middle. But cut the threads on one side and the ring is drawn towards the other side.

There are two important consequences of this one-sided attraction of the molecules on the surface of a liquid. The first is, that the surface layer of a liquid is tougher than the rest of it, for it is not so mobile; and the second is, that the molecules beneath the surface are always trying to draw the surface molecules into the interior. That cannot be, however, or the surface of the liquid would become smaller, and that could not happen unless the vessel or its shape were changed. So that the surface of a liquid is always tending to become smaller, and it acts as if it were a skin, tougher than the rest of the liquid, always stretched, and constantly trying to get smaller, like a piece of stretched indiarubber. This apparent stretching of the surface of a liquid is called its

surface tension. It is stronger in some liquids than in others; it is strongest in water. Substances dissolved in water lower its surface tension or weaken the surface layer.

Thus you can perform the following experiment. On the surface of the water in a vessel place half a dozen match-sticks pretty close together. Touch the water between the sticks with the corner of a piece of soap. You weaken the surface layer just there, and the stronger-stretched layer outside the sticks shrinks up and drags them to the side of the vessel. A drop or two of alcohol would have the same effect as the soap.

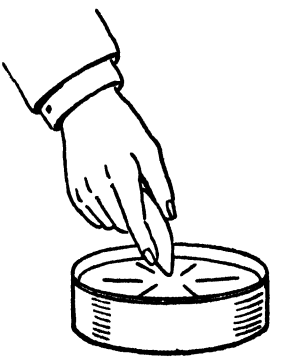


FIG. 3.

Dip a piece of clean glass vertically into water or into any other liquid which wets it, and the liquid rises up a little, close to the glass, because of the adhesion of the glass to the liquid. Dip a very fine glass tube in the liquid, and the adhesion causes a slight hollow in the surface of the liquid in the tube, and the surface tension of this concave or hollow surface drags the liquid quite a long distance up the tube. The smaller the tube the higher does the liquid mount, and of all liquids water mounts the highest. This is called *capillary attraction*, and we see its effects on all bodies having fine pores.

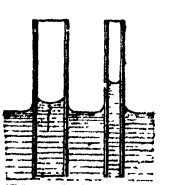


FIG. 4.—(From Ganot's *Physics.*)

Water rises into the pores of a piece of sponge or a lump of sugar or salt. In a lamp or candle the oil or melted wax rises up the wick by capillary attraction. Fill a glass with water and place in it a piece of lamp-wick, which hangs out over the side of the glass. The water will rise up one side of the wick by capillary attraction, and fall down the other by gravity. In this way the glass may even be emptied.

There is no capillary attraction between a solid and a

liquid which does not wet it. A glass plate or a narrow glass tube in mercury depresses the liquid, and so does a greased glass plate or tube in water. A needle, which has not been specially cleaned to free it from grease, will float on water, just dimpling the surface without being able to break through it.

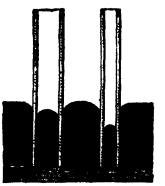


FIG. 5.—(From Ganot's *Physics.*)

If we have a thin layer or film of liquid, it has two surfaces, and there is a toughened skin on both sides. You can make a good liquid for films or bubbles in this way. Take—

Oleate of soda (or white Marseilles soap), 20 grams.

Dissolve this in ½ litre of distilled or rain-water. When dissolved add 30 grams of glycerine, put it in a litre bottle, and fill up with distilled or rain-water after shaking well. Cork it up tightly, keep it in a cool place for a week, and it will then be ready for use. Pour out as much as you want to use at any time, and do not return the used liquid to the bottle. With this liquid many instructive experiments may be made.

Make a ring of wire by passing it round a small bottle and twisting the ends together to serve as a handle. Dip this in the solution, and it will come out with a film on it. Make a little loop of silk, dip it in the liquid, and carefully place it in the middle of the film. Break the film within the loop by touching it with a hot wire or the corner of a piece of blotting-paper, and immediately the stretched film outside the loop retreats towards the ring on all sides, and the silk loop is drawn out into a circle.

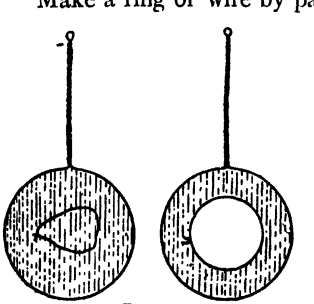


FIG. 6.

Make an oblong of two stout wires, about 4 inches long, and two silk threads. Across it tie another loose silk thread, with another attached to its middle, and provide one of the

wires with a loop so that it can be handled easily. Fig. 7 will show you how it is made. Dip the whole in the soap

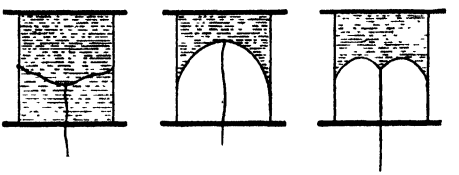


FIG. 7.

solution, and draw it out carefully by the loop with a film covering it. Break one half of the film, and immediately the cross thread is drawn back in the form of a semicircle. You may pull on the loose end and draw it down again, thus stretching the film, but on releasing the end the cross thread is again drawn back.

If you can get a lamp-glass which is smaller at one end than the other, as in Fig. 8, you may try another very interesting ex-

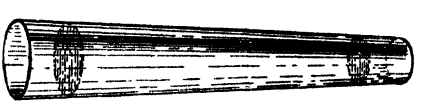


FIG. 8.

periment. Wet the whole of the glass inside with the soap solution, and then dip in the larger end, and bring it out with a film on it. Now you see this film slowly travel along inside the glass to the smaller end. Why is this? Because the surfaces of the film are trying to become smaller, just as a stretched sheet of indiarubber would, and the only way in which it can be done is by moving along the glass towards the smaller end.

You may make of iron wire, or get, many skeleton solid figures, as in Fig. 9, each furnished with a little handle, and dipping them in the soap solution, bring them out again with most curiously shaped and interesting films upon them.

For the blowing of bubbles you may use a clay pipe, but there are many better things than that to do it with. Pieces of glass or metal tubing of various sizes will do very well, and for

very large bubbles you should use a child's toy trumpet or a piece of tin bent to that shape.

Dip the large end of the trumpet into soap solution and blow a very large bubble. Without shaking it off, hold the other end near a lighted candle, and the candle will be blown

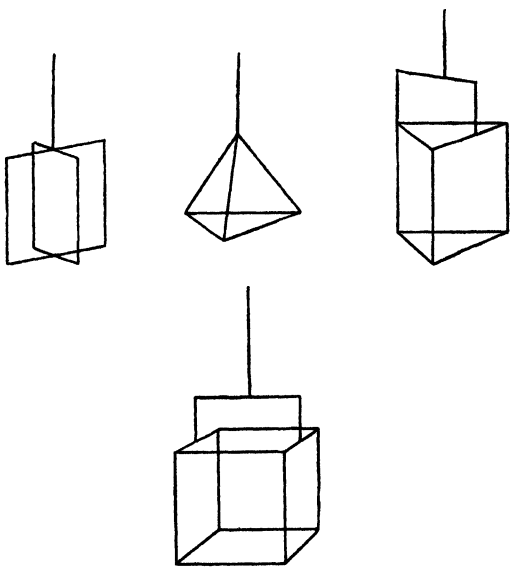


FIG. 9.

out. This is because the bubble, like the flat film, becomes smaller if it can, and in doing so drives out the air with which it is filled.

Blow a large bubble, and carefully place it on a wire ring wet with soap solution. You will very likely not succeed the first time; but try again. You may now dip a smaller tube in the solution; pass it carefully through the wall of the bubble, and blow another one inside the first (Fig. 10). As many as three or more may be blown inside each other if great care is taken.

A bubble on the wire ring may be drawn into a barrel

shape by attracting to its upper end another ring previously dipped in the solution (Fig. 11). It may even have another bubble blown inside it, which may be rolled to and fro and around the barrel-shaped bubble.

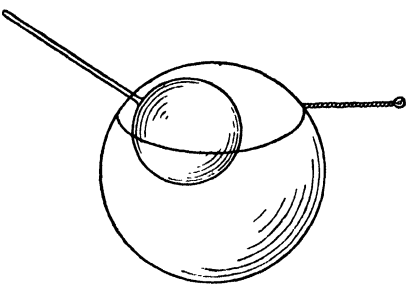


FIG. 10.

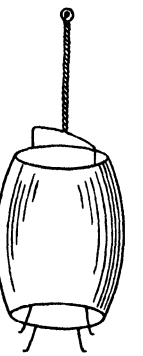


FIG. 11.

In all cases of dealing with bubbles and films, you must never touch them with anything which has not previously been dipped in the soap solution, or they will be broken.

SUMMARY.

The *free surface of a liquid* acts like a stretched skin.

Liquids pass into fine pores by *capillary attraction.*

There is no capillary attraction between a solid and a liquid which does not wet it.

A *film or bubble* of liquid has *two free surfaces,* each of which acts as if it were a stretched skin.

LESSON II.

MOISTURE IN THE AIR.

Apparatus needed:—Flasks, tubes, and Bunsen burner or spirit-lamp; balance and shallow dish; ice; sulphuric acid; U-tube, pieces of pumice-stone, and an aspirator; barometer-tube, mercury, and bent pipette; piece of indiarubber.

If you fit a flask with a cork and a bent glass tube through it, as in Fig. 12, and after placing some water in the flask, heat it until the water boils, you will see dense clouds issue from the tube, whilst the space over the water is perfectly clear. You know already that this space is full of the water-gas or vapour which the water is being turned into by the heat, and that therefore water-vapour, or *steam*, as it is called, is quite invisible. If you watch the clouds issuing from the tube, you will note that they are visible only for a short space. Not far from the tube they disappear into the air of the room. What are these clouds composed of?

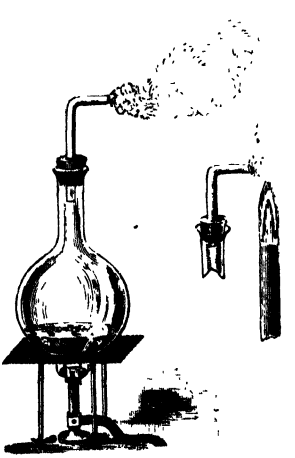


Fig. 12.—(From Newth's *Elementary Practical Chemistry*.)

Hold a piece of cold glass for a moment in this cloud, and it is covered with a multitude of tiny drops of water. Place the flame of a Bunsen burner or spirit-lamp so that the issuing clouds are strongly heated, and

they immediately disappear. They are composed of very small drops of water in the air. The invisible steam is cooled a little when it issues from the tube, and is, of course, turned into water. Heated by the flame of the Bunsen or the spirit-lamp, the drops are again turned into steam. As the clouds float away they are more slowly vaporized, and the invisible vapour disappears into the air. You can see the same thing happen to the clouds which issue from a railway engine.

Half fill a small shallow vessel with water, weigh it, and put it aside where it will not be disturbed. You had better cover it loosely with a piece of paper, in order to prevent dust falling into it. Weigh it day after day for several days, and note that it gets lighter every day, because some of the water is constantly being vaporized, and the vapour passes into the air. If you leave it long enough all the water will pass into the air. It will dry up, as we say. Why do people hang out wet linen in the air, and what becomes of the water wetting it? There is a great deal of water in grass, as you can find out by taking a handful of it and twisting it strongly in your hands. What becomes of this water when the grass is cut and spread out on the fields to dry into hay?

Because so much of the surface of the earth is covered with water—seas, lakes, rivers, and damp soil—we should expect to find there is always water-vapour in the air. We know that many lakes and ponds, and even rivers, dry up in a hot, dry season, and we know that their waters pass into the atmosphere. Can we show that this vapour exists in the air? If there is any of it in the air, and we cool it greatly, it ought to be condensed to water.

Bring into a warm room a glass containing cold water or ice, and the outside of the glass is immediately covered with water. The cold glass cools the air which touches it all round, and the water-vapour in that air is cooled too, so that it is condensed to water on the glass. For the same reason the window-glass of a warm room on a cold day becomes covered with drops of water. If the glass is cold enough, the condensed vapour freezes, and the window-panes are covered with frost.

If you hang up a flask full of broken ice by the neck in a warm room, the water-vapour in the air of the room will continuously condense on the flask; and if you place a vessel beneath it, the condensed vapour will drop into the vessel. Quite a large quantity of distilled water may be collected in this way. We may call it distilled water, because all water condensed from vapour is distilled.

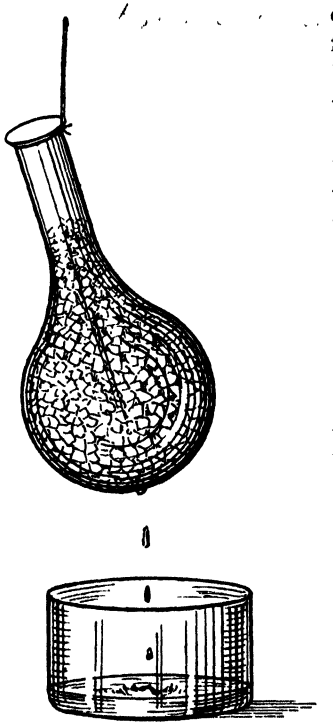


FIG. 13.

Into a little vessel put a small quantity of strong sulphuric acid. Weigh the vessel and the acid together, and place them aside where they will not be disturbed, covering loosely with a piece of paper to keep out dust. Weigh the vessel and the acid day after day for several days, and note that there is a continual increase in weight. This is because sulphuric acid is so greedy for water that it readily takes it from the air. Even in the driest and hottest weather the acid grows heavier, showing that there is water-vapour in the air at all times, even when you would least expect it.

Expose a little fused calcium chloride or a little pearl-ash in an open vessel to the air. These also take water from the air, and grow heavier and heavier and wetter and wetter every day, until they are quite sloppy.

Substances which take water from the air in this manner are called *hygroscopic substances.*

If we wish to find out exactly how much moisture there is in the air at any time, it can be done by using a hygroscopic

substance. We take a glass tube, for preference a U-tube, and fill it full of pieces of pumice which have been wetted with strong sulphuric acid. Each end of the U-tube is fitted with a cork and a bent glass tube, and the whole weighed. Then the tube is connected with a large vessel full of water, whose

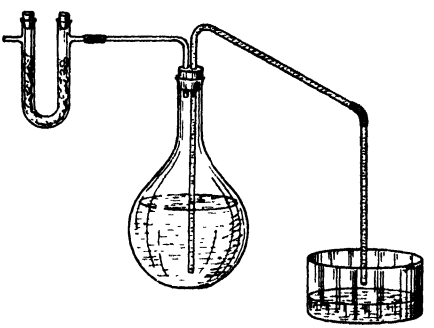


FIG. 14.

capacity is known. This vessel is made to act as an *aspirator* to draw a current of air through the tube. It is easily understood that the volume of air which is drawn through the tube is equal to the volume of water which runs out of the aspirator. All the water-vapour in the air is caught and held by the sulphuric acid, so that the increase in weight of the U-tube is the amount of water-vapour which was contained in the air passed through it.

An instrument like this, which is sometimes called a *chemical hygrometer*, will tell us that there is always water-vapour in the air, and that there is frequently more on a hot summer's day than on a cold winter's day. Why is that? The following experiment will tell us something about it.

A straight, dry barometer-tube is taken, filled with dry mercury, and inverted into a vessel of mercury in the usual way to form a barometer. There is then a space over the mercury which contains no air. Then a little water is sucked into the bend of a bent pipette like that in the figure, the point of the pipette is placed under the open end of the barometer-tube, and a drop of water is blown out of the pipette, care being

taken that no air follows it. Because water is less dense than mercury, the drop rises to the surface of the mercury in the tube, and immediately the mercury falls a short distance. It would not fall because of the weight of this little drop of water, even if the drop had not disappeared into vapour, but it fell because the water was immediately converted into vapour in

FIG. 15.

the empty space, and this vapour pressed down the mercury just as air would. A few more drops are blown into the tube, and the mercury falls a little further. But it falls no more when a drop remains on the top of the mercury. Then, as much water has been turned into vapour as the space will hold, and we say that the space is *saturated* with water-vapour at the temperature of the experiment.

If the experiment were performed at a higher temperature more water would be vaporized in the space before it was

saturated. Place underneath the open end of the barometer tube a piece of soft indiarubber, lower the tube, and press it tightly against the rubber. Whilst in this position warm gently the top of the tube containing the drop of water, and the drop will also be turned into vapour.

In a space filled with air the same result follows. Warm air will hold more water-vapour than cold air before it is saturated, just as warm water will dissolve more sugar than cold water before the solution is saturated.

Now you see why the hot air of summer often contains more water-vapour than the cold air of winter. It is because

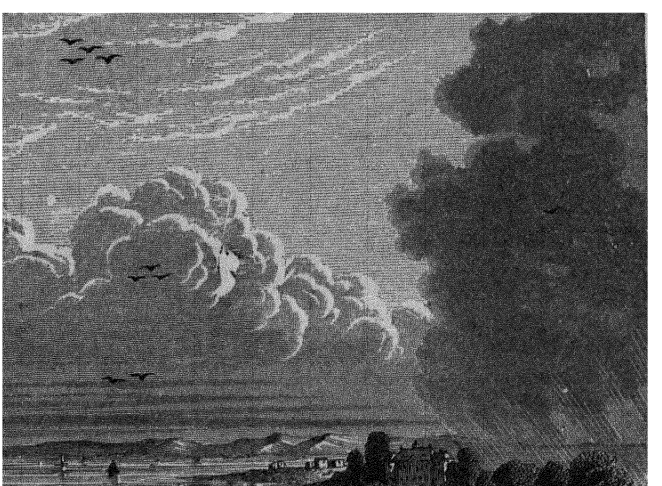


FIG. 16.—(From Ganot's *Natural Philosophy*.)

it can hold more and yet not be saturated. When the air is saturated and the temperature falls, some of the water-vapour is condensed, and a *cloud* appears. If the condensation is very rapid, the little drops join together and form larger drops, which fall as *rain*. Rain always forms from water-vapour in the air, and it is therefore distilled water, although it is not often pure,

because it brings down with it dust which is always present in the air.

When a cloud is formed near the surface of the earth, we call it *fog* or *mist*, and when the cloud-drops are formed on the surface of plants and other things, it is called *dew*.

If the water which would otherwise form dew is deposited at a temperature below the freezing-point of water, we call it

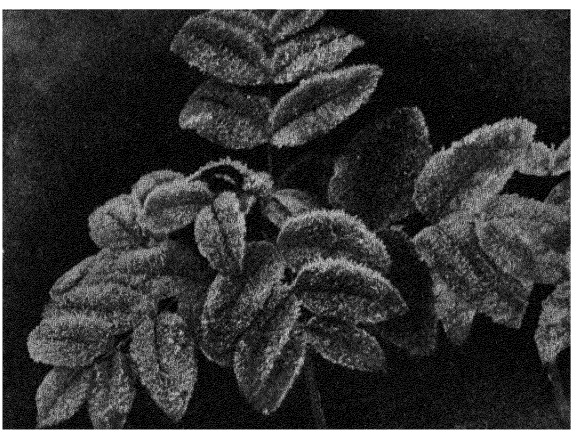


FIG. 17.

hoar frost, because objects, instead of being covered with little drops of water, are loaded with tiny ice-crystals.

Sometimes water-vapour high up in the air is cooled below the freezing-point, and is then condensed directly to crystals of ice. When these crystals cling together they fall as *snow* (Fig. 18).

It is very interesting to consider how the water of the earth is continually circulating. It rises from sea or land as vapour, and is often carried far away from its place of origin by winds. A time comes when it is condensed and forms clouds, and when these clouds descend in the form of rain, or snow, or hail. When it reaches the ground a large portion of it flows away and forms rivers, and from the rivers it again reaches the sea.

Air which is saturated or nearly saturated with water-vapour

is said to be damp, and then moist objects dry very slowly, and a slight lowering of the temperature brings condensation. When it contains much less water-vapour than is necessary to saturate it, we call it dry air, and then wet objects dry quickly because water evaporates readily, and a considerable lowering of temperature must take place before condensation occurs. It has already been explained how that the air in summer often contains more water-vapour than in winter. It may contain more in summer than in winter, and yet be drier, because its temperature being higher it is not so nearly saturated as the winter air.

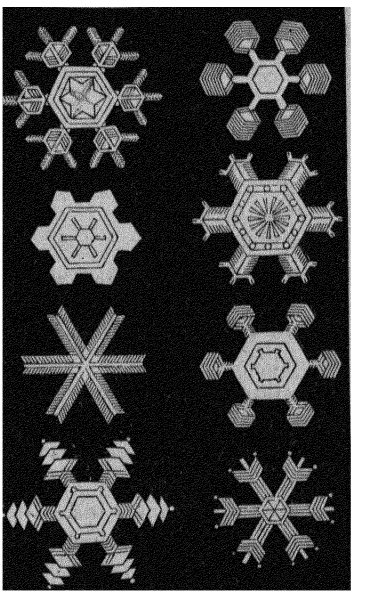


FIG. 18.—(From Ganot's *Physics.*)

Very dry air, as well as very moist air, is not healthy for us. It is in the best condition for breathing when it is about three-fourths saturated.

SUMMARY.

Steam or water-vapour is *invisible.*

Clouds are composed of small *drops of water.*

Exposed water slowly *passes into the air* as vapour.

There is *always water-vapour in the air.*

Hygroscopic substances take water from the air.

Frequently there is *more water-vapour* in the air on a *hot summer's day than* on a *cold winter's day.*

Air is saturated with water-vapour when it *can contain no more.*

Hot air contains *more water-vapour* before it is saturated than *cold air.*

LESSON III.

GRAPHIC REPRESENTATION.

Apparatus needed:—A blackboard ruled in squares will be useful.

In a certain class in a school there were 38 boys who did not all attend as regularly as they ought to have done. One week there were present in that class on Monday morning 34 boys, and on Monday afternoon 36 boys. On Tuesday morning 36 were present, on Tuesday afternoon 32, on Wednesday morning 36, on Wednesday afternoon 30, on Thursday morning 34, on Thursday afternoon 32, on Friday morning 32, and on Friday afternoon 28.

There are much better ways of stating these facts than printing them in a paragraph like the above. A plainer way of setting them down is this:—

Number of boys present on— .

Monday morning,	34	Wednesday afternoon,	30
„ afternoon,	36	Thursday morning,	34
Tuesday morning,	36	„ afternoon,	32
„ afternoon,	32	Friday morning,	32
Wednesday morning,	36	„ afternoon,	28

When printed in columns like this a person can read the figures off much more easily.

Here is another way of setting them down.

Number of boys present on—

Monday.		Tuesday.		Wednesday.		Thursday.		Friday.	
Morn.	Aft.	Morn.	Aft.	Morn.	Aft.	Morn.	Aft.	Morn.	Aft.
34	36	36	32	36	30	34	32	32	28

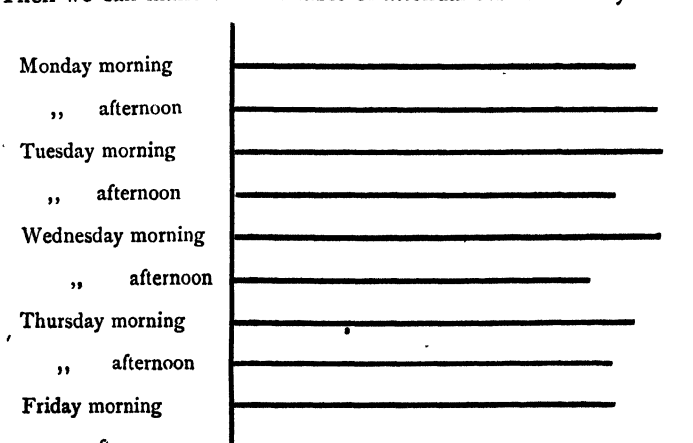
Fig. 19.

This way is neater than either of the others, takes up less room, and is much more easily understood.

Or we may add together the number of boys present on each occasion the school was opened that week, and divide the sum by the number of times the school was opened, and thus obtain the average number present every morning and afternoon.

Thus the total number of attendances is 330, and the class met ten times that week. Therefore the average attendance of that class is $330 \div 10$, or 33. This is a number very useful to know for some purposes, but it does not tell us enough. We want to know the number present on each occasion, and to set down these numbers so that we can see at a glance how large the attendance was every morning and afternoon during the week.

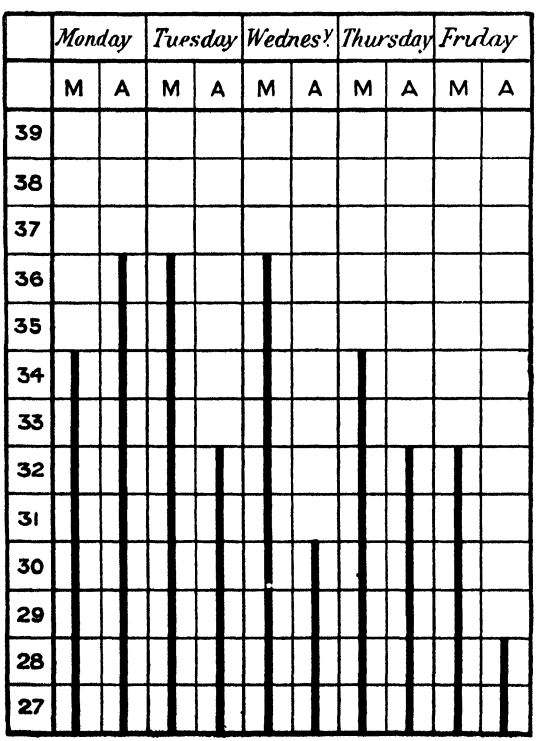
Let us try some other ways of doing it. Suppose we represent the numbers by lines—the longer the line the more boys it represents. Let us say that a line $\frac{1}{16}$ of an inch long shall stand for one boy. Then $\frac{2}{16}$ of an inch will stand for two boys, $\frac{3}{16}$ for three boys, and so on. For 36 boys we shall want a line $36 \div 16$ inches long—that is, $2\frac{1}{4}$ inches. Then we can make out our table of attendances in this way—



From this table we see, at a glance, that the attendance was largest on Monday afternoon, Tuesday morning, and Wednesday morning, and smallest on Friday afternoon. But to find the exact numbers these lines stand for we must have a rule divided into inches and parts of an inch, and each line must be measured.

We can manage without the rule and the measuring by the use of table like this. We use a piece of paper upon which a large number of small squares are drawn—we call this squared paper—and we let each square stand for a boy. And as the smallest number present was 28 and the largest 36, we do not want to use the whole 38 squares.

Number of boys present on—



I think you will say that this is much plainer and better. You see at a glance that two boys were absent even on the days when the attendance was greatest, and as the squares are numbered no rule is necessary to measure the line. Every day's attendance can be seen in a moment, so that this is a real *graphic method* of representing the numbers. The lower portions of the lines are not given at all because it is not necessary to make any measurements. We may indeed do without the lines altogether by joining their ends with a bent line. Here is the same figure without the lines.

Number of boys present on—

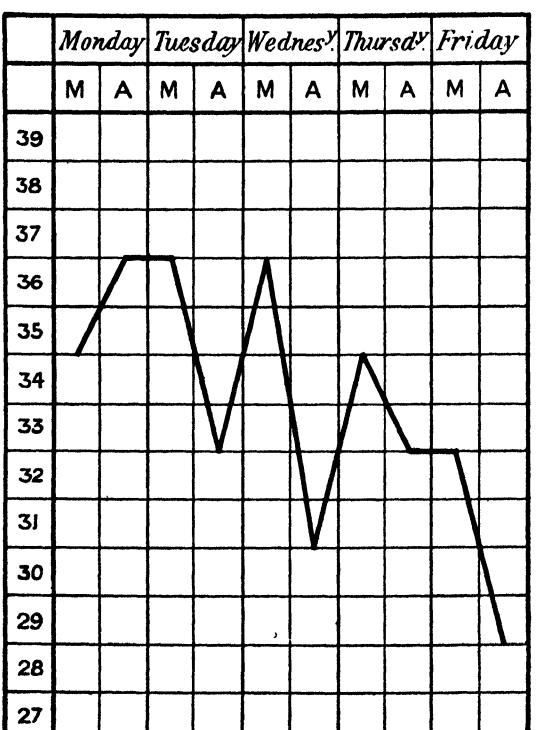


FIG. 22.

You see that in this figure the vertical lines are not drawn at all, only their tops are connected by a line which rises and falls just as and when the attendance rises and falls, and thus gives a complete picture of it. See what a drop there is on Tuesday, Wednesday, and Friday afternoons.

You must note that there are two things which vary in all these tables. The number present throughout the week varies, and the days vary. We say that they vary together, and this graphic or pictorial method of representing quantities which vary together is a very convenient one, and may be applied to any such cases. Let us apply it to a case like this, which shows how the price of coals varied during the last six months of the year 1900.

Take a piece of squared paper, and let the squares taken

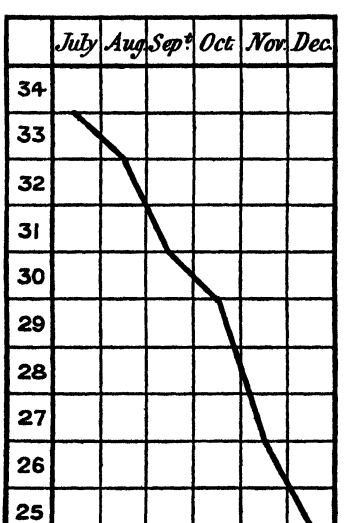


FIG. 23.

in vertical columns represent the months, and those in horizontal columns represent the prices per ton in shillings. Head the vertical columns with the names of the months, and the horizontal columns with all the possible prices.

Then we get a figure like this, and it shows us at a glance that the price of coals fell continuously from July to December, and also the amount it fell every month.

Here is another figure which represents the rate at which a boy spent his pocket-money in a certain week last year.

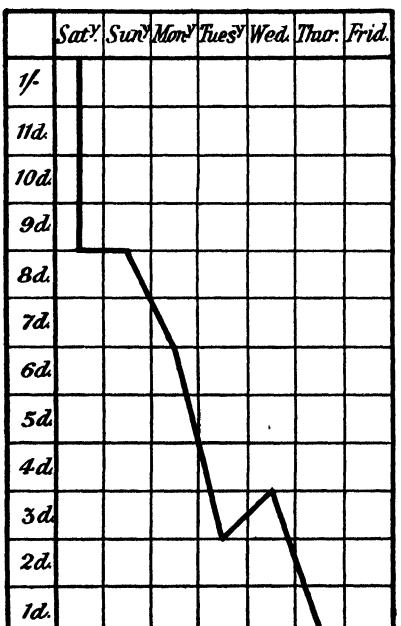


FIG. 24.

He had a shilling given him on Saturday, which he immediately began to spend, and threepence was also given him on the following Wednesday. Find out from Fig. 24 how much he spent every day.

SUMMARY.

Numbers may be represented by lines.

Two quantities which vary together may be well represented by a curve.

LESSON IV.

VARIATION OF THE BAROMETER AND THERMOMETER SHOWN BY GRAPHIC METHODS.

WHEN you have been at the seaside you have very likely noticed at the coastguard station, or in some other public place, a barometer and a thermometer exposed behind a glass screen in such a manner that any one can easily see them. These are kept most carefully from every kind of injury, and several times a day the height of the mercury in the barometer which measures the pressure of the atmosphere, as well as the height of the mercury in the thermometer which measures the temperature of the air, are noted, and a record kept of their variations. The readings, as they are called, of these instruments are regularly telegraphed from all parts of this country, as well as from other places in Western Europe, to a central office in London.

Here they are all collected and compared, and skilled men study them to find out as far as possible what the weather is likely to be during the next twenty-four hours in any part of the country. These forecasts are sent every day by telegraph to persons all over the country, and most of the newspapers publish them with much other information about the weather.

The variations of the barometer and thermometer are very well shown by the graphic or pictorial method, and almost every newspaper has a different way of showing them.

The London *Daily Telegraph* publishes every day a figure,

of which this is a copy for March 19, 1901. It shows how the barometer fell almost continuously during the four previous days. The vertical columns represent the days, the horizontal columns the height of the barometer in inches and tenths of

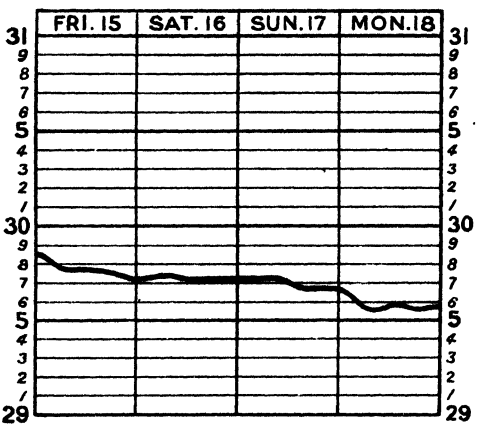


FIG. 25.—*Daily Telegraph* weather chart.

an inch. You will see from this table that the barometer stood at $29\frac{9}{10}$ inches on the morning of Friday, March 15. Twenty-four hours later it stood at $29\frac{7}{10}$ inches. During Saturday, the 16th, it was steady, and continued so until about midday on Sunday, when it began to fall again. At midnight, Monday, March 18, the height of the barometer was $29\frac{6}{10}$ inches. A continuous fall like this foretells much bad weather.

On the same day the London *Daily Chronicle* published a weather chart, of which Fig. 26 is a copy. It shows how the barometer varied for the two previous days, the observations being made and recorded four times a day—at midnight, six o'clock in the morning, midday, and six o'clock in the evening. It also shows on the same figure the variations of the thermometer on the same two days. In order to show both of these, the horizontal columns are figured, on the left

in inches and tenths of an inch for the barometer readings, and on the right in Fahrenheit degrees.

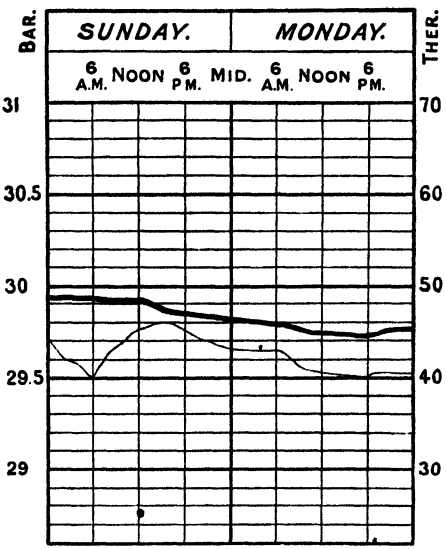


FIG. 26.—*Daily Chronicle* weather chart.

Many of the more important provincial papers also show the variation of the barometer by graphic methods. This is the way the Newcastle *Daily Chronicle* shows it daily for the

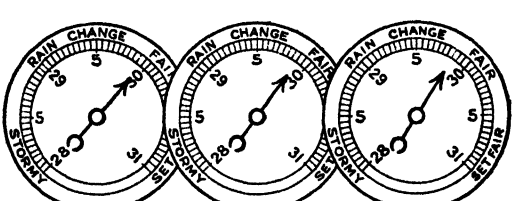


FIG. 27.—The Newcastle *Daily Chronicle* weather chart.

previous twenty-four hours. It represents the clock-face of an old-fashioned wheel barometer at three different times in the twenty-four hours.

The London *Times* shows the height of the barometer every two hours during the previous twenty-four by a series

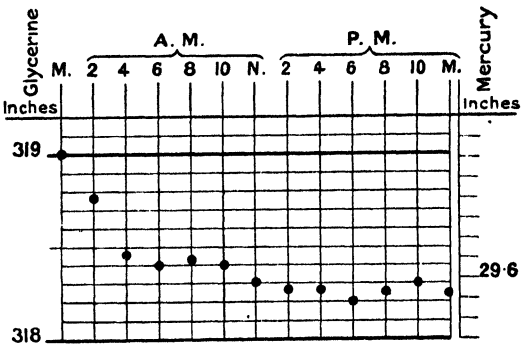


FIG. 28.—The *Times* weather chart.

of dots. And it shows the height, not only of a mercury barometer, but also of one containing glycerine.

These figures are enough to show quite clearly that the variations of the barometer and thermometer, from time to time, are most easily and clearly shown on a figure of straight lines crossing each other at right angles. But suppose we

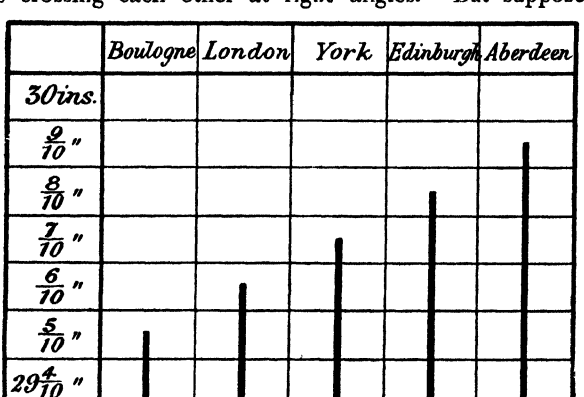


FIG. 29.

wanted to show how, at the same time, the barometer varied in different places. Suppose that on March 18, 1901, at

6 p.m., the height of the barometer in Boulogne, France, was $29\frac{5}{10}$ inches, in London $29\frac{6}{10}$ inches, in York $29\frac{7}{10}$ inches, in Edinburgh $29\frac{8}{10}$ inches, and in Aberdeen $29\frac{9}{10}$ inches. Could that be shown by a graphic method? And if it can, which is the best?

The simplest and best will be similar to most of those you have already studied. Fig. 29 will do it for you.

But suppose we want to know more than this figure shows us. Suppose we want to show not only at what places the barometer stands at a different height, and the amount of that difference, but also at what places the barometer stands at the same height. Figures such as we have often used will not do for this. A chart similar to that which is sent out every day from the Meteorological Office in London is necessary. The *Times* publishes a copy of this chart every day, and here is a copy of it for March 18, 1901.

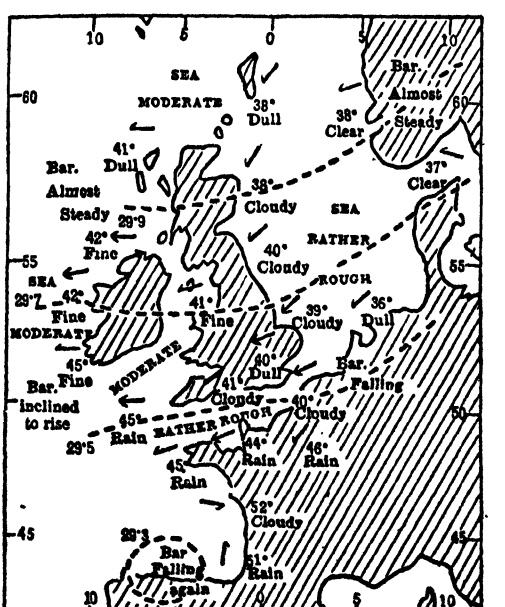


FIG. 30.—Weather chart of the Meteorological Office. (From the *Times* newspaper.)

It is a rough map of Western Europe. All places where the barometer stands at the same height are joined together by a dotted line, and marked with the height of the barometer. Thus you can see from the chart that all along a line passing up through the English Channel, the north of France, Belgium, and Holland, the barometer stood at $29\frac{5}{10}$ inches. The line of $29\frac{7}{10}$-inches passes through Ireland and England, and the $29\frac{9}{10}$ inches line through the north of Scotland. In the north of Spain the line connecting all the places where the barometer stood at only $29\frac{3}{10}$ inches is nearly a circle, and there where the pressure is least, the chart tells us that the barometer is still falling. These lines are called *isobars*.

This chart or map tells us a great deal more than the distribution of the atmospheric pressure. You will see arrows all over it; these show the direction and force of the wind at the various places. The state of the weather and the sea is also shown all over the west of Europe by words printed on the map.

SUMMARY.

The variations of the barometer and thermometer may be well shown by the graphic or pictorial method.

LESSON V.

SUBSTANCES SOLUBLE OR INSOLUBLE IN LIQUIDS.

Apparatus needed:—Glass vessels; sand, powdered charcoal; filtering funnel, stand, and papers; sugar, salt; sealing-wax, shellac, and methylated alcohol; sulphur and carbon disulphide; fat and ether.

You remember having learnt that many substances can exist in the solid, liquid, and gaseous states, and that any one state can be converted into another by adding heat to the body, or subtracting heat from it. Thus ice, which is the solid form of water, is converted into the liquid form by heating it, and liquid water is turned into gaseous steam by heating it. On the other hand, steam is converted into liquid water when heat escapes from it, and water into ice when it loses heat.

You also know that in the solid condition the cohesion of the molecules is so great that they do not easily move about amongst each other; in the liquid state the cohesion of the molecules is less, and they easily move about amongst each other; whilst in the gaseous state there is little or no cohesion, and the molecules are always trying to get further away from each other.

There are some solids whose molecules separate from each other when they are placed in certain liquids. These molecules spread through the whole body of the liquid, no matter how large the volume of the liquid may be, spreading themselves evenly throughout the whole of it. This is just what a small quantity of a gas does when introduced into a vessel, either empty or already filled with another gas. We say that such a solid is *soluble* in that liquid, or that it *dissolves* in it, and we call the liquid a *solvent* for that solid.

But there are also solid bodies which do not dissolve in any liquid.

Take a glass half filled with water, and place in it a little clean sand. Stir it up well, and then let it stand for a little while, and you will find that all the sand falls to the bottom of the glass. Taste the water, and you will observe that its taste is not altered.

Take another glass of water, and add some charcoal powder to it. Stir this up well, and let it stand. It will take a much longer time to settle than the sand did; but finally you will find it all at the bottom of the vessel, whilst the clear water above it has acquired no taste from the charcoal.

You may do the same with a small quantity of soil, but this will take a much longer time to settle to the bottom. For a very long time the water will be slightly muddy above the *sediment* at the bottom of the glass, the tiniest particles of the soil remaining *suspended* in the water for a considerable period of time.

Sand, charcoal, soil, and a great many other solid bodies do not dissolve in water. This is the same thing as saying that their molecules do not separate from each other and spread themselves throughout the water. They are *insoluble* in that liquid.

If we wish to separate the insoluble substance from the water, we can let it settle first, and then carefully pour off the water. This is called *decanting* it.

Or, if we wish to do it more quickly than by decantation, we *filter* the water.

Filtering a liquid is allowing it to pass through something which is full of tiny pores. The solid particles suspended in the liquid are larger than the pores, and are kept back whilst the liquid passes through. Anything which has plenty of small pores will do for filtering, but the most convenient thing we can use is a piece of white blotting-paper. The kind you use at school will do, although a thinner and tougher kind, called filter-paper, is better.

You take a round piece of filter-paper, and fold it on itself twice over. Then place this in a glass funnel, and open it out

to fit the funnel. If you support the funnel in a ring, and place an empty beaker underneath it, as in Fig. 31, your apparatus is quite ready. Now you pour the liquid which is to be filtered slowly and carefully on the paper in the funnel, and almost immediately you see it pass out of the funnel clear and bright into the vessel below, whilst the solid particles remain on the paper. You may open out the paper and allow it to dry, when you have the insoluble solid in the same condition as it was before you used it.

FIG. 31.—(From Thorpe's *Qualitative Analysis*.)

Now have another glass half full of water, drop into it a lump of sugar, and stir it well. In a short time the sugar disappears, and you see it no more. But if you taste some of the water, you will find that it is sweet; and if you drink it all, you swallow all the sugar likewise.

The sugar has dissolved in the water, and its molecules are now evenly spread throughout the water, so that a cubic centimetre taken from any one part contains as much sugar as a cubic centimetre taken from any other part. You cannot get the sugar back again from the water by filtering it, because now the sugar is so finely divided that it passes through the pores of the filter-paper as easily as the water does. Nothing would be left on the filter-paper.

Try the same experiment with salt and water. The salt

easily dissolves as the sugar did. It disappears, becomes evenly spread throughout the water, so that a drop from any part of it tastes salt, and you cannot get the salt back again by filtering it through blotting-paper or anything else.

No doubt you know of a great many other things besides sugar and salt which dissolve in water. You could easily make a large list of the names of such substances.

But there are some things which are insoluble in water, and yet they dissolve in other liquids. For instance, get some pieces of shellac, place them in water, and shake up well. When allowed to stand the shellac sinks to the bottom of the vessel, and the water remains clear above it. None of the shellac is dissolved; the molecules will not separate from each other in water.

But place some shellac into a vessel containing methylated spirit, and shake it up well. The shellac completely dissolves after a time, and is evenly distributed throughout the spirit. This solution of shellac in methylated spirit forms a useful kind of varnish.

Try putting some pounded sealing-wax into methylated spirit. Sealing-wax is made of shellac, to which has been added a quantity of some solid colouring matter. The spirit will dissolve the shellac, but the red, yellow, or green colouring matter will remain undissolved.

Again, shake up some flowers of sulphur in water, and none of it dissolves. But place a little into a liquid called carbon disulphide, and it dissolves at once.

You know quite well that fats will not dissolve in water. But put a piece of butter or lard into a very mobile liquid called ether, and the fat dissolves in it.

You must be very careful when you are handling methylated spirit, or carbon disulphide, or ether. Each of them is changed into a gas very easily, and this gas takes fire so readily that no light must be allowed to come near the liquids.

SUMMARY.

When one substance dissolves in another, the molecules of the *dissolved substance* are evenly distributed throughout the *solvent substance.*

A liquid can be separated from finely divided, suspended matter by *decantation* or by *filtration.*

Sugar and salt are examples of substances *soluble in water.*

Shellac and sealing-wax dissolve in *methylated spirit.*

Fats dissolve in *ether.*

LESSON VI.

SOLUTION.

Apparatus needed:—Balance, beakers, sugar, and salt; evaporating-dish, stand, and Bunsen burner or spirit-lamp; blue vitriol or alum; distillation apparatus.

WHEN sugar or salt dissolves in water it disappears, and we see it no more. We know, however, that it is still there because the solution tastes sweet or salt. We can show that it is still there in another manner.

Let us take a glass beaker and weigh it. Suppose it weighs 50 grams, and we place in it 100 grams of water. Let us also take a lump of sugar or salt which weighs, say, 15 grams. Altogether the beaker, water, and sugar weigh 165 grams. Now we dissolve the sugar or salt in the water, and weigh again the glass with the solution in it, and we shall find that it still weighs 165 grams. So that whether the sugar or salt is in the solid state, or whether its molecules are distributed amongst the molecules of the water, it still weighs 15 grams. The same

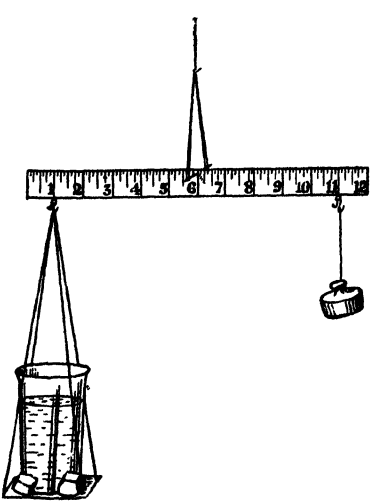


FIG. 32.

can be shown for any other substance which dissolves in a liquid.

If we wish to recover the dissolved sugar or salt, or other solid substance, we cannot do it by filtering the solution, but we can by another method. If a solution of salt is left exposed to the air for a time, the water slowly evaporates away, and the salt is left behind dry in the dish. Or if the solution is heated, the evaporation of the water is much more rapid, and the salt is recovered more quickly.

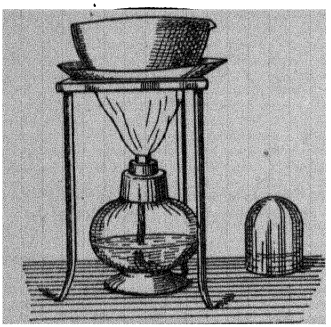


FIG. 33.

Put the solution of salt into a little evaporating-dish, place this on a sand-bath or piece of wire gauze over a Bunsen flame or a spirit-lamp, and heat it gently. The water is driven off in the form of steam, and the solid is left in the dish. Whatever amount of salt was dissolved in the water, that amount will be recovered on evaporating the water.

We cannot dissolve as much of a solid as we like in a liquid in which it is soluble. Let us try with blue vitriol or alum. We take a vessel containing water, and we add the solid to it, a little at a time, stirring it well all the while. Presently we find that the water will dissolve no more. All that we now add sinks to the bottom of the vessel and remains there, and not all the stirring in the world will make the water dissolve more of it. The water has as much as it can hold, and we say that it is *saturated* with that solid, or that it is a *saturated solution* of the solid.

Suppose now we heat it a little over the gas or spirit flame. We find that the solid which did not dissolve when the water was cold now becomes dissolved, and we may be able perhaps to dissolve still more of the blue vitriol or alum in it. But sooner or later the solution again becomes *saturated*, when it will dissolve no more.

But this *hot saturated solution* contains more dissolved solid than did the *cold saturated solution*. What will happen, then, if we let the hot saturated solution cool down until it is of the same temperature as before?

The result will be that all the additional solid which was dissolved by the water on heating it cannot remain in solution, and becomes solid again, sinking to the bottom of the vessel.

Let us pour off the liquid into another vessel. We now have again our cold saturated solution of blue vitriol or alum. Can we get any more of the dissolved solid from it without evaporating the water? Yes; but in order to do so, we must cool it still more. If we stand the vessel in ice, the temperature of the solution falls, and more of the solid comes out of solution, because at this lower temperature a saturated solution does not contain so much dissolved solid as at the middle temperature.

So we learn from these experiments that the hotter the water is the more blue vitriol or alum it can dissolve before it is saturated, and that if we cool a hot saturated solution, as much of the dissolved solid must come out of solution as will leave a cold saturated solution. And this is the case with most other solids which are soluble in water. You can try the same experiments with sugar and washing-soda.

Water is the commonest liquid, and water dissolves more common solids than any other liquid does. We meet with no water which does not contain solids in solution unless we prepare it specially by condensing steam. This process is called the *distillation* of water.

In order to *distil* water we boil it. Whatever solids are dissolved in the water are left behind, but the water itself is driven off in the form of steam; and if the steam is led into a vessel or through a tube kept cool by any means, the steam condenses again to a liquid. But this liquid is pure water. Let us take a little of it, place it in an evaporating-dish, and heat it. When all the water has evaporated, the dish is left quite empty, showing that no solid substances were dissolved in this *distilled* water (Fig. 34).

But take a little water from the tap, or pump, or river, and evaporate it from a dish. A solid mass is left behind, showing that the water we ordinarily drink contains dissolved solid substances.

Rain is distilled water, although in falling through the air it gets slightly fouled with dust which is always present in the air.

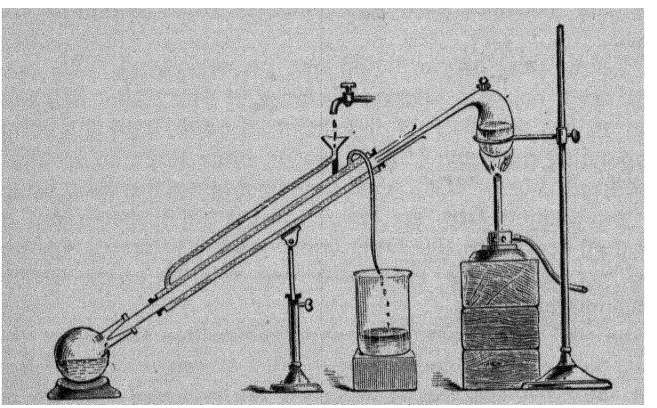


FIG. 34.—(From Reynolds's *Experimental Chemistry.*)

In towns it gets quite a large quantity of dust and soot in it. The rain falls on the ground, and a great deal of it soaks into the earth and dissolves substances from it. The water of some rivers contains much dissolved solids, and sea-water contains still more.

Sea-water contains a very large quantity of common salt. In many parts of the world the sea-water is evaporated in order to get the salt from it. Here is a description of some salt-pans on the west coast of France.

When the tide is very high certain large ponds all over the district become filled with sea-water, and each pond supplies several salt-pans. These pans have mud floors and mud partitions, which divide them into many shallow compartments communicating with each other by channels like the paths in a maze. When summer comes the water is allowed to flow into the pan in a gentle stream. Fig. 35 is a plan of one of

these pans. The water flows from the pond at A, passes along the channel AB, and enters the pan at B. Then it flows from compartment to compartment along a path marked in the figure by a dotted line, and all the while the sun's heat and the wind are evaporating the water, and the solution of salt is approaching nearer and nearer to saturation. At last, when it has reached

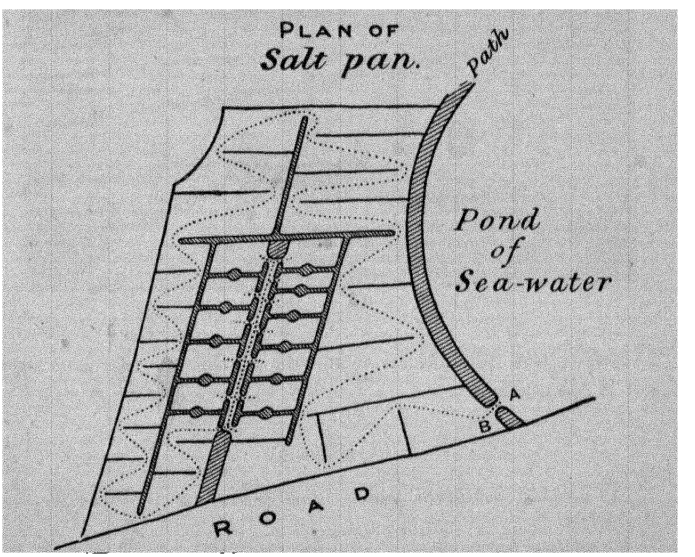


FIG. 35.

the middle compartments, it is saturated, and some of the salt comes out of solution, forming a crust on the surface of the water.

This crust is carefully skimmed off by men and women, and placed in heaps on the mud partitions, which also serve as pathways. As soon as one crust of salt is removed another forms, for the supply of sea-water is constant, and evaporation is still going on. The pans continue to yield salt in fine weather all the summer and autumn (Fig. 36).

In this country there are large beds of salt deep down in the earth in Cheshire and Durham. Mines are dug, and water

is allowed to flow down to the salt beds. This dissolves the salt, and the solution is pumped up to the surface. The water

FIG. 36.—Collecting salt from salt-pans.

is evaporated by artificial heat, and in this manner the solid salt is obtained.

SUMMARY.

A *soluble substance* which disappears in a liquid is *not destroyed.*

Dissolved sugar or salt can be *recovered* by driving off the solvent water by heat.

A *solution is saturated* when it will dissolve no more of the soluble substance.

A *hot solvent* will dissolve more than a *cold one.*

We *distil a liquid* by first converting it into vapour, and then *cooling the vapour.*

LESSON VII.

MORE ABOUT SOLUTION.

Apparatus needed:—Beakers and methylated spirit; lump sugar; distillation apparatus; beer; glycerine and ether; oil, mercury; bottle of soda-water; flask, tube, pneumatic trough, and Bunsen burner or spirit-lamp.

It is not solids only which spread their molecules throughout the volume of a liquid, and so are said to dissolve in them. We know of liquids which distribute their molecules throughout other liquids, and gases which dissolve in liquids, and mix with or, we might almost say, dissolve in other gases.

Thus let us take a vessel of water and add a little alcohol to it. Stir it up, and the alcohol mixes completely with the water or dissolves in it. Add more alcohol, and we get the same result. And however much alcohol we add to the water we get a complete mixture, so that we cannot saturate water with alcohol as we can with common salt. They easily mix together in all proportions.

Take a rather narrow cylindrical vessel, and fill it about one third full of water. Then pour very carefully on to the surface of the water, by means of a funnel, some alcohol which has been lightly coloured by a few drops of red ink. If this is done very carefully, the red-coloured alcohol will float on the surface of the colourless water, because it is lighter than water. Now put it away where it will be seen but not disturbed, and you will find that very slowly some of the water will rise up amongst the alcohol, whilst some of the alcohol will sink down amongst the water until both liquids are thoroughly mixed.

A somewhat similar experiment can be carried out with a soluble solid and water. Take a vessel of water, and carefully

drop some lumps of sugar into it. Place the vessel aside, and watch it for some days or weeks. The sugar at the bottom very slowly dissolves, and, in spite of its weight, rises up amongst the water until the solution of sugar is as complete as if it had been well stirred.

Suppose we mix together alcohol and water, is there any method by which we may separate them again? Yes; we can do so by distillation.

Put the mixture in a thin glass bottle or flask provided with a cork, through which passes a bent glass tube. The end of the tube is to be led into another empty flask, kept cool by cold water. Heat the flask over a gas or spirit flame. This is like the apparatus of Fig. 37. Whatever liquid is driven off by heat as a gas must pass through the tube and be condensed in the cold flask or receiver. But both these liquids are easily converted into gases by heat, so that if we heat a mixture of them, will not both of them be distilled? They will be, but not to the same extent. You know that water boils at 100° C. Alcohol boils at 78° C. A mixture of these two liquids boils at a temperature which is between 78° C. and 100° C., and although both of them are driven off in the gaseous condition, much more of the alcohol is distilled than the water; so that the liquid obtained by condensing the vapours contains a larger proportion of alcohol than the original mixture did. This may be distilled over and over again, getting a stronger solution of alcohol in water or a weaker solution of water in alcohol every time. Try it with beer, which is a solution of alcohol in coloured water.

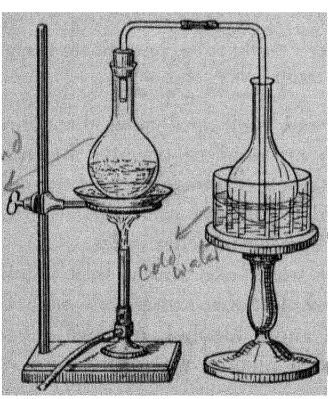


FIG. 37.

Glycerine is another liquid which dissolves in water in all proportions, whilst ether dissolves in water, but only to a slight

extent. Water is saturated with a very small quantity of ether, and ether is saturated with a very small quantity of water. If more ether is added to water than it can dissolve, it floats on the surface of the water. These two liquids do not mix together in all proportions like alcohol and water do.

But there are liquids which do not mix together in any proportions. Pour some oil into water, and stir it up well. As the mixture comes to rest the oil separates from the water, rises to the top, and forms a floating layer there. Mercury does not mix with either water or oil. You may shake up in a vessel mercury, water, and oil, and when the mixture settles, the three liquids will have quite separated from each other, and arranged themselves in three layers, the mercury at the bottom, the oil at the top, and the water between them, as in Fig. 38.

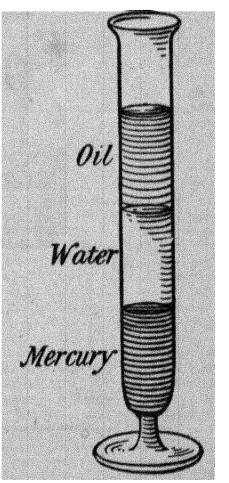


FIG. 38.

Gases dissolve in liquids, but not in all proportions. When you open a bottle of soda-water or lemonade, you find that a large number of bubbles of gas come from the liquid. It is this gas dissolved in the water which gives the drink such a pleasant, refreshing character. In making the soda-water or lemonade, the water was purposely made to dissolve as much of the gas as it would.

If you fill a glass of water from the tap or the pump, and let it stand a little, you will see bubbles of gas rise from it. This gas is air which the water had dissolved.

Water will not dissolve more than a certain quantity of air or any other gas. When it has dissolved this quantity it is saturated, and will take no more of it into solution; so that in this respect it acts towards gases just as it does towards soluble solids. There is, however, this difference. You have learnt that, as a rule, water will dissolve more of a solid the hotter it is. But it will dissolve less of a gas the hotter it is. And this shows us how to get the gas out of solution.

If we have a saturated solution of a solid in water, it must be cooled to get some of the solid out of solution. But if we have a saturated solution of a gas, such as air, in water, we can get some of the gas out of solution by heating it. And if we boil the water for a time, we can get it all out.

Fit up a flask with a good cork and a glass tube through it, bent like the tube in Fig. 39. Then, when the cork is pressed

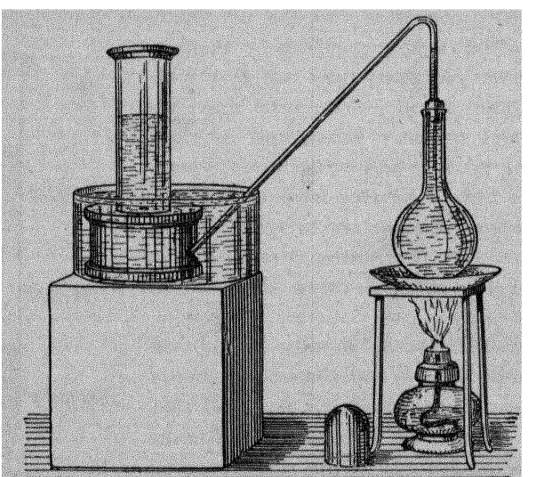


FIG. 39.

well home, nothing can come from the flask except through this tube. Take a basin three-quarters full of water, fill a gas-jar with water, and invert it under the surface of the water in the basin. Then the pressure of the air will keep it full of water.

Fill the flask quite full of water from the tap, press in the cork, and support the flask in such a way that you can get a gas or spirit flame beneath it. Place the end of the tube underneath the open end of the gas-jar, light the gas-burner or spirit-lamp, place it beneath the flask, and allow the water to boil. The air will be driven out of solution in the water, pass through the tube, and rise in the gas-jar. Some of the water in the flask will be turned to steam, but that will condense in the cold water of the basin.

Air is always present in water unless we take pains to drive it out, and this air is of the greatest importance to the life of animals and plants which live in water. Fish cannot live in boiled water, because they would have no air to breathe, and water-plants could not live in water which had no dissolved air for the same reason. All living things require air; without air no life is possible.

Most gases mix together in all proportions, but this is not the case with water-gas or steam and air. The water from seas, rivers, lakes, and damp soil is always passing away as steam into the air, so that the air has always water dissolved in or mixed with it. But just as we can saturate water with air, so we can saturate air with water-gas.

Hot air can hold more steam than cold air before it is saturated; therefore, if hot saturated air is cooled, some of the water comes out of solution. Your breath is hot, and is saturated with water-gas. If you breathe against a cold piece of glass your breath is cooled; it cannot hold so much water in solution; a portion of it comes out of solution, and you see the glass covered over with a layer of water.

In cold weather, when you breathe out air, some of the moisture comes out of solution, and you see it as a little cloud of tiny water-drops. Bring a glass of cold water into a warm room, and it cools the air around it, so that the air gives up a portion of its water, which coats the outside of the glass.

Frequently the air is saturated with water. If, then, it is cooled by a cold wind or by any other means, a portion of the water comes out of solution, and we get rain, or snow, or fog, or mist. Clouds are composed of multitudes of very little drops of water, which were dissolved in the air, but which have come out of solution on account of the cooling of the air.

SUMMARY.

Many *liquids dissolve in each other*—that is to say, they mix intimately.

Mixed *alcohol and water* may be *separated* by *distillation.*

Many *gases dissolve* in water. Gases are driven out of solution by heat.

LESSON VIII.

CRYSTALS AND CRYSTALLIZATION.

Apparatus needed:—Blue vitriol, beakers, saucer; alum, washing-soda, common salt; solution of cobalt chloride; sulphur and carbon disulphide; gum.

You already know that when a solution of a solid is heated, the solvent liquid is driven off as a gas, whilst the solid remains. You also know that when a saturated solution of a solid is cooled, some of the solid comes out of solution. But there is something more to be learned about this.

Make a strong solution of blue vitriol, a stuff which the chemists call sulphate of copper. Place 50 grams of it in 100 cub. cms. of water, and heat it over the gas or spirit flame until it is completely dissolved. Then let this hot saturated solution cool.

As you would expect, after what you have already learnt, some of the solid comes out of solution. Not, however, in the form of a powder, but in the form of blue crystals. You may say, perhaps, that you do not know what a crystal is. Well, examine the mass of blue vitriol after the liquid has been poured off from it, and you will find that it is not powdery, but consists of a solid mass of well-shaped pieces joined firmly together. Every piece is shaped and has flat sides. *All crystals have flat sides.*

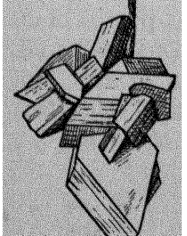


FIG. 40.

You may take the cold saturated solution which you poured off from the crystals, place it in another vessel, and allow a small piece of cinder or stone, tied to the end of a bit of cotton, to hang in it. Place it on one

side, and in a few days some of the water will have evaporated, and you will find the cinder covered with a mass of nice crystals.

To get good, well-shaped, single crystals, you must pour your cold saturated solution of blue vitriol into a saucer, and place it where it will slowly evaporate. Cover it with a piece of paper or glass to keep out the dust, and wait. Then in a short time you will find small crystals appear on the bottom of the saucer. Select the best of these, take out all the rest, and turn it over gently every day, so that it can grow equally on all sides. In this way you will obtain a well-shaped crystal.

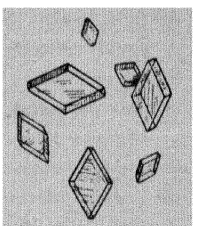


FIG. 41.

Try all this over again with alum, washing-soda, and common salt, and you will get crystals of all of them.

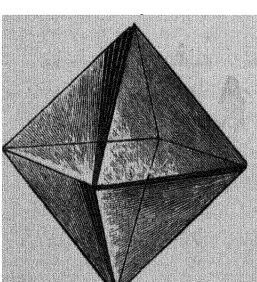


FIG. 42.—Crystal of alum.

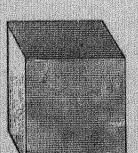


FIG. 43 [illegible] salt. (From [illegible] *Elementary Mechanics.*)

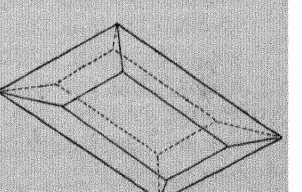


FIG. 44.—Crystal of washing-soda. (From Watts's *Dictionary of Chemistry.*)

Now take a nice clean piece of washing-soda, wash it in water, and wipe it dry with a clean duster. You see that it is quite transparent. Take a small evaporating-dish, place the washing-soda in it, and weigh both together; then place it over the gas or spirit-burner, and heat it gently.

The washing-soda first melts, then dries up, and when the dish is cold, and you weigh it again, you find that it weighs much less than before. The heat has driven something off from the washing-soda, and the white powder which remains

weighs only about one-third of what the crystals weighed. What has been driven off?

In order to answer this question get some more washing-soda, place it in a flask similar to that which is used for distilling water, heat it, and condense the gas coming from it. A liquid collects in the receiver, and you will find that this liquid is water. Washing-soda cannot exist in the form of crystals without this *water of crystallization.*

When a mass of crystals of washing-soda is left exposed to the air for a time, the water of crystallization evaporates from the outside of the crystals, leaving them covered with a white powder. This is called the *efflorescence* of the crystals.

Powder some blue vitriol, and heat the blue powder in an evaporating-dish. Water of crystallization is driven off from this, and a *white powder* is left. This white powder is also sulphate of copper, but sulphate of copper without any water of crystallization. It is *amorphous*, which means that it is without any crystalline form.

If you pour a little water on the white, powdery, amorphous washing-soda it will dissolve, and you can again get crystals of washing-soda from this solution. If you pour water on the white amorphous sulphate of copper it turns blue again, dissolves, and you can again get blue crystals of blue vitriol from the solution. You see that it is the presence or absence of water which determines whether sulphate of copper shall be blue or white.

There is another crystalline substance, called cobalt chloride, which changes its colour when you drive away its water of crystallization. In the form of crystals it is pink, but in the amorphous form it is blue. You can try a very pretty experiment with this. Dissolve a little cobalt chloride in water, and write with it on a piece of paper. As it dries on the paper it crystallizes, and the pink colour of the crystals is so faint that the writing is nearly invisible. Now hold the paper in front of a fire, and so drive off the water of crystallization, and the writing comes out quite plain in the blue amorphous cobalt chloride.

There are some crystalline substances which have no water

of crystallization. Common salt is one of these. If you heat common salt no water is driven off, and no reduction in weight takes place.

It is not only from solution in water that crystallization takes place. Place a little pounded sulphur in an evaporating-dish, and add some of that evil-smelling liquid called carbon disulphide. Stir it well until the sulphur is dissolved, and be very careful to keep it at a good distance from a fire or light, for the liquid evaporates quickly, and its vapour is very inflammable. Place the dish where the liquid can evaporate without danger, and as the evaporation proceeds, beautiful, transparent, light yellow crystals of sulphur will appear. You had better let the carbon disulphide evaporate outside the room, or, at any rate, in a draught of air, until the dish is quite dry. This will not take long.

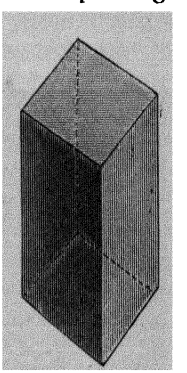


FIG. 45.—Crystal of sulphur. (From Grieve's *Elementary Mechanics*.)

Some solids are not known to crystallize at all. Thus you cannot crystallize gum from a solution in water, nor shellac from a solution in alcohol. As the solvent liquid evaporates, the solution becomes more and more viscid, but the solid does not crystallize.

Many bodies form crystals, not from solutions, but when they pass from the liquid to the solid state or from the gaseous to the solid state.

Water itself is one of these substances. When it freezes it always forms crystals. Snow is formed of little crystals of ice, and so is hoar frost. You see these crystal forms best when the moisture on the window freezes on a cold night. Even solid ice is a mass of crystals. You must have noticed when the water of the pond or in an open vessel is freezing, how crystal needles of ice shoot out from the banks until the whole of the pond is covered with them (Fig. 46).

Sulphur is another substance which crystallizes when it becomes solid after melting, as well as when it condenses from the gaseous form.

Place some sulphur in a little dish, and heat it until it melts Be careful not to let it catch fire. If it should do so, cover up the dish with a piece of slate or glass until the flames are extinguished. When it is all quite melted, let it cool, until a crust forms on the surface. Then quickly grasp the dish with a pair of tongs, break the crust at one side, and pour as much of the melted sulphur as will flow away into cold water. If you

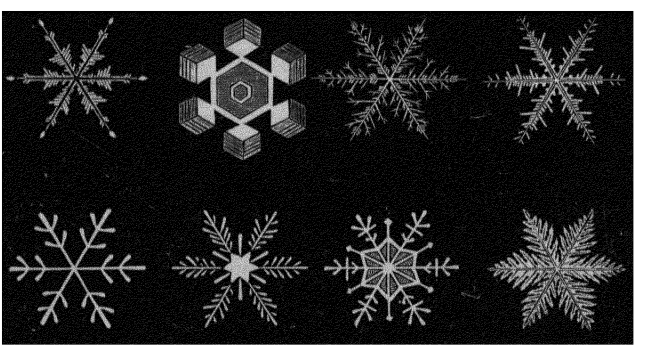


FIG. 46.—(From Ganot's *Natural Philosophy*.)

have done this at the right moment, the whole of the inside of the dish will be full of needle-shaped crystals of sulphur, all springing from the sides. If you delay too long, all the sulphur will form these crystals, but so close together that the whole will form a solid lump.

You may note that the sulphur which was cooled in water is not crystalline.

Put a piece of sulphur in the bottom of a test-tube, and heat it. First it melts, then it boils, and passes off as a gas. Higher up the tube this gas condenses to little crystals of sulphur. You will need a small magnifying-glass to see that they really are crystals.

A piece of cast iron is just a mass of crystals which formed when the melted iron solidified. Other metals form crystals under the same conditions.

We find many substances in the earth which are in the crystalline form, and we cannot tell how many of them have

been formed, whether from solution in a liquid or from fusion. The diamond is one of these. This is a mere piece of carbon in the crystalline form. Charcoal is also a kind of carbon; and if you can find out how to dissolve or melt charcoal so that it shall crystallize from solution or fusion, you will be able to make diamonds.

SUMMARY.

The solid deposited from a hot saturated solution on cooling is in the form of *crystals.*

Many crystals contain *water of crystallization* which may be driven off by heat.

The *water of crystallization* sometimes gives their *colour* to crystals; always their *shape.*

Sulphur crystallizes from its solution in carbon disulphide.

Some solids do *not crystallize* from their solutions.

Sulphur and some other solids *crystallize* on *cooling* after fusion by heat

LESSON IX.

ANOTHER KIND OF SOLUTION.

Apparatus needed:—Sodium, dish, test-tubes; zinc and sulphuric acid; flask, cork, tubing, acid funnel, and gas-jar; pieces of iron.

Be very careful how you handle sodium. Never do it with wet fingers, or you will get badly burnt.

Let us now try another experiment with a solid and water. We take a small piece, about the size of a pea, of a very soft, silvery white metal called *sodium*, and we drop it upon some distilled water in a small dish or saucer. It immediately rounds itself off into a little ball, floats, and rushes about the surface of the water, hissing all the time. It quickly gets smaller, and soon disappears. Has it dissolved in the water? Let us see if we can get it back again by evaporating the water.

We pour the liquid into an evaporating-dish, and heat it until all the water has been driven away, but no silvery white metal is left in the dish. Instead of that we have a small quantity of a white powder, which we can easily prove weighs more than the piece of sodium we used.

We will take some of this white powder, and drop it on water. It dissolves quietly enough. It is clear, then, that the white powder is not sodium, and that water does not simply dissolve sodium, but changes it into another substance.

We can learn something more about this by performing some more experiments. Drop another piece of sodium on water, and whilst it is moving about on the surface touch it with a lighted taper, and immediately it bursts into flame.

Now take a test-tube, fill it with water, and invert it into a

vessel of water, so that it remains full. Wrap a small piece of sodium in a little bit of wire gauze or blotting-paper, attach it to the end of a lead pencil, and quickly introduce it into the water, so that it is under the inverted tube. Almost immediately bubbles of a gas proceed from it, and rise up through the water to the top of the jar. When all the sodium has disappeared, so that no more bubbles rise, put your thumb over the mouth

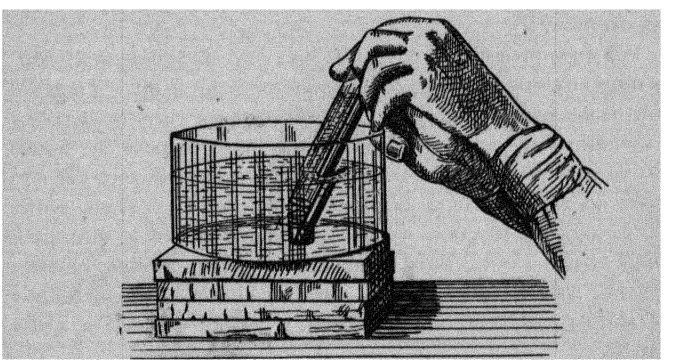


FIG. 47.

of the tube, and erect it. Now have a lighted taper ready, remove your thumb, and apply a light. The gas burns with a pale blue flame. You remember that we once boiled air out of water. But this is not air, for it burns.

So now you see that the disappearance of sodium in water is not ordinary solution, for a gas is set free which burns, and a white powder is found dissolved in the water, and this is not sodium. It is not like the solution of blue vitriol, alum, salt, sugar, and other substances in water, nor of shellac in alcohol, nor of sulphur in carbon disulphide, for in all these cases we could recover the solid by evaporating the liquid. This is a different kind of solution altogether.

Now let us try another metal and a different liquid. There is a fearfully burning liquid called oil of vitriol, or sulphuric acid. *Do not let a drop of this touch your flesh or your clothes, or you will be badly burnt.* We mix water with some sulphuric

acid—about ten times as much water as acid—and we are careful to pour the acid into the water, not the water into the acid. This we call *dilute sulphuric acid.* When it is cold—for much heat is produced when the acid and water are mixed—we pour a little of it into a small dish, and drop into it a small piece of zinc, and watch it. Bubbles arise from the zinc in large numbers, and the metal gradually disappears. The bubbles show that some gas is being set free. We will catch it presently, and examine it.

We drop piece after piece of zinc into the dilute acid until no more bubbles escape, occasionally stirring it up. Then we filter it, because it will almost certainly appear dirty on account of the zinc being impure. When quite clear we pour the liquid into an evaporating-dish, and heat it to drive off some of the water, so as to leave a hot saturated solution. Then, when this is allowed to cool, we get white crystals of a substance which is not zinc. This substance is called *sulphate of zinc*, or *white vitriol.* So that you see that here also the metal has not simply dissolved in the liquid. It has suffered a great change.

Now to examine the gas which is given off when zinc disappears in dilute sulphuric acid.

Take a small flask, and fit it with a good cork having two holes through it. Through one hole pass an acid funnel far enough to reach nearly to the bottom of the flask. Through the other pass a piece of glass tubing bent to the shape shown in Fig. 48.

Into the flask place some pieces of zinc, and replace the cork. Allow the other end of the tube to dip in the water contained in a basin. Pour some dilute sulphuric acid down the funnel, and see that the end of the funnel dips beneath the acid. By this means the gas which is set free is prevented from escaping from the flask by way of the acid funnel, and must get out through the bent tube. We immediately see that gas is being set free, for a multitude of tiny bubbles rise up from the pieces of zinc, and larger bubbles rise up from the end of the tube through the water in the basin.

We must let the first bubbles escape, because the space

over the acid in the flask is at first filled with air, and the gas first coming off is mixed with air. When all the air in the flask has been expelled, we hold an inverted test-tube filled

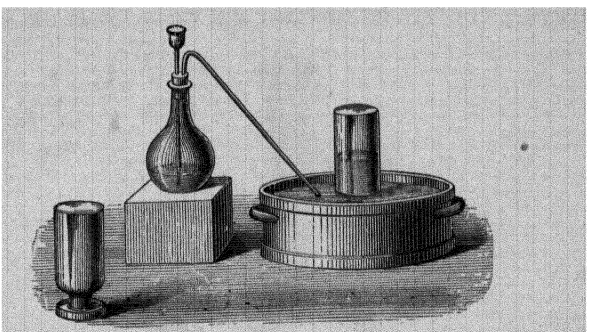


FIG. 48.—(From Thorpe's *Qualitative Analysis.*)

with water over the end of the tube, and so fill the test-tube with the gas.

When the test-tube is quite full we lift it out of the water, still keeping it inverted, and hold a lighted taper to its mouth. The gas catches fire, and burns with a pale blue flame which is difficult to see. This is how the gas acted which escaped from water when sodium was placed in it. It is the same gas.

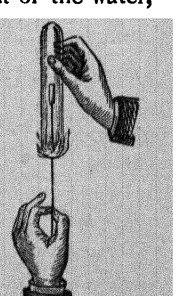


FIG. 49. — (From Reynolds's *Experimental Chemistry.*)

We catch some more in an inverted test-tube, which, however, is only half filled with water, the remainder being air. In this way we get a mixture of half this gas and half air. When we touch this off with a lighted taper, we get a vicious little explosion. If you catch the gas from sodium in a test-tube half filled with air, you can explode that mixture in the same way.

It is a very light gas, much less dense than air. Collect a test-tube full of it, and pour it *upwards* into another test-tube which has nothing but air in it at first. The lighted taper will soon show that the gas has all passed out of the first test-tube into the second (Fig. 50).

We call this gas *hydrogen.*

A great many other metals besides zinc dissolve in sulphuric acid. Try pieces of iron. Hydrogen escapes as the iron disappears, and may be collected and examined in the manner

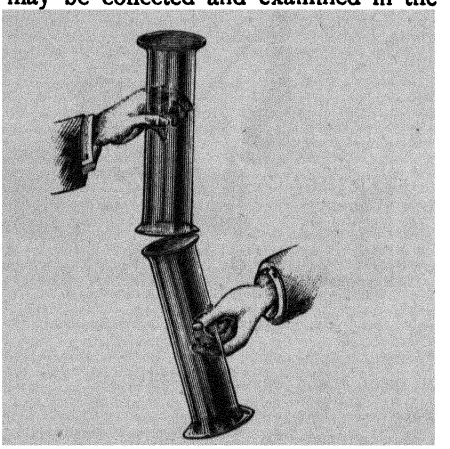


FIG. 50.—(From Newth's *Elementary Practical Chemistry.*)

already described. The liquid, however, in the flask is of a green colour, and when it has been filtered we can get from it green crystals of *sulphate of iron*, or *green vitriol*, as it is often called.

SUMMARY.

Sodium disappears in water, but it is *not* simply *dissolved.* The sodium is *changed* and a *gas is produced.*

Zinc disappears in oil of vitriol, but no zinc can be recovered from the solvent by evaporation. An *inflammable gas* is given off.

Iron also disappears in sulphuric acid, and an *inflammable gas* is produced.

LESSON X.

MORE ABOUT SOLUTION IN ACIDS.

Apparatus needed:—Copper, sulphuric acid, beaker, flask, and delivery-tube; Bunsen burner or spirit-lamp and retort-stand; pieces of zinc; nitric acid and pieces of lead; small crucible and powdered charcoal; hydrochloric acid and washing-soda.

TAKE a beaker and place in it a small quantity of strong sulphuric acid. To this add a few shavings or borings of copper, or, failing that, small pieces of scrap copper or copper wire. So long as the vessel is kept cold the copper is not dissolved; but if it is warmed, the copper dissolves in the acid, and suffocating fumes of a gas which smells like burning sulphur are given off. As soon as you smell the gas, throw the contents on to the soil—where it will do no harm—and wash out the beaker.

As we do not wish to study this gas, we will dissolve the copper in a flask provided with a delivery-tube, and so arranged as to deliver the gas into the sink-hole, whilst the tap is kept running, or, if that is not easy, into a large vessel of water, where it will be dissolved. The apparatus is pictured in Fig. 51. Put the copper into the flask, add the strong oil of vitriol, and warm the flask. When no more gas comes off, if there is no copper left undissolved in the flask, take out the cork, add some more, and replace the cork. You want to ensure that when no more gas is given off there is still some copper undissolved. Now open the flask, let it cool a little, and add as much water to it, a little at a time, as the acid which was used at first, then shake it up well.

Now filter into a clean beaker, and allow it to stand. The

clear liquid which passes through is of a bright blue colour, and as it stands in the beaker and cools, a great number of blue crystals form at the bottom of the beaker. These look like the sulphate of copper, or blue vitriol, which we have already used for our experiments. It *is blue vitriol*, or *sulphate of copper.*

We therefore see that when the metal copper dissolves in sulphuric acid, a suffocating gas, which smells like burning sulphur, is given off, whilst blue vitriol, or sulphate of copper, is formed.

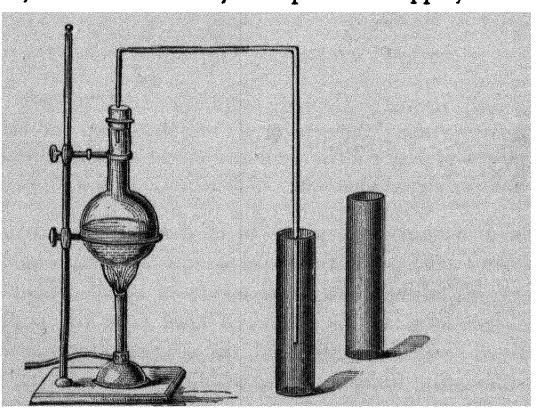


FIG. 51.—(From Furneaux's *Elementary Chemistry.*)

Now we want to know whether there really is any copper in this blue vitriol. Can we get copper from it? We will see.

Pour away the water from the crystals in the bottom of the beaker, add more water, and stir until the crystals are dissolved to a blue solution. Then place into this solution some pieces of clean zinc, and warm gently. After a time you find that the solution is no longer blue, but colourless, that the zinc has disappeared or partly so, and that a red powdery substance is at the bottom of the beaker. Take out any undissolved zinc, shake up the remainder, and filter it, and the red powder remains on the filter-paper, where it may be allowed to dry. *This red powder is the metal copper.*

What has taken place? A very simple exchange of zinc for copper. We had *sulphate of copper* in solution and *free zinc.* Zinc has taken the place of copper in the sulphate, and

now we have *sulphate of zinc* in solution and *free copper*. The copper was in the blue sulphate of copper.

Another experiment, this time with a new acid and another metal. There is a very strong acid which is commonly called *aqua fortis*, but which chemists call *nitric acid*. We pour a little of this into a beaker, add twice as much water, and then place in the dilute acid some pieces of lead. This we gently warm, and see that large quantities of a reddish-brown gas are given off. *This gas is very suffocating and poisonous.* We must therefore dissolve our lead out-of-doors, or where the fumes may pass up the chimney, or down the sink, or into some other harmless place.

By-and-by we see that as the lead dissolves a white powder is collecting in the beaker. When a sufficient quantity of this has formed, we stop the heating, and after taking out any undissolved lead, we allow the beaker to cool. When quite cold, the liquid over the white powder is carefully drained away from it, and the powder is scraped out into an evaporating-dish and dried at a very gentle heat. This powder is *nitrate of lead.*

We therefore see that when lead is dissolved in nitric acid, a suffocating, reddish-brown coloured gas is given off, whilst white nitrate of lead is formed.

Now we want to know whether there is really any lead in this nitrate of lead. Can we get lead from it?

Take first a beaker, and place in it a portion of the white powder. Add a little water to it, and stir it well. It is not very soluble in water, and it is probable that it will not all dissolve; but that does not matter. Take a piece of clean zinc, tie a string to it, and allow it to hang suspended in the solution in the beaker in such a manner that the zinc is wholly immersed. Support it in that position by a rod laid across the top of the beaker.

You will soon see the zinc covered with a black substance, and if you allow it to remain undisturbed for some hours or days, you will obtain a lovely mass of black, shining crystals of metallic lead—that lead which was contained in the nitrate of lead, and which was originally dissolved in the nitric acid.

Something has happened similar to what took place when

zinc replaced the copper in sulphate of copper. Here, zinc has replaced lead in nitrate of lead.

If you allow it to remain long enough, and if the solution is strong enough, all the zinc will be dissolved, and the mass will be wholly lead. To satisfy yourself of that, you may melt the lead in an iron spoon, and so produce the familiar soft, shiny metal.

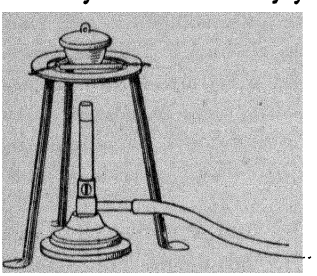


FIG. 52.—(From Tilden's *Hints on the Teaching of Elementary Chemistry.*)

Take the other portion of nitrate of lead, put it into a small crucible, and heat it very strongly over the burner so long as reddish-brown fumes come from it. Then add a teaspoonful of powdered charcoal to it, stir it up with an iron rod, and keep it red-hot for several minutes.

After a time take the crucible from the burner by a pair of tongs, and turn out its contents upon a flat stone or something of that sort. When cold some globules of a white metal will be found amongst the fragments. This metal is soft, easily cut with a knife, and will write on paper; in fact, it is *lead.*

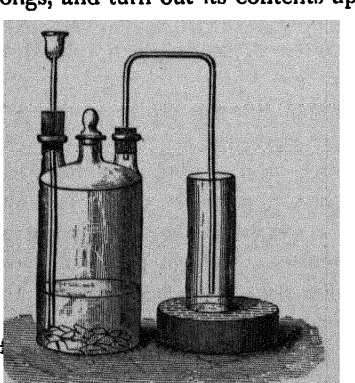


FIG. 53.—(From Miller's *Chemistry.*)

Let us now try an experiment with another acid, which is sometimes called spirits of salt, but by chemists hydrochloric acid. We place in a flask provided with a delivery-tube, like the one in Fig. 53, some washing-soda, and we add through the acid funnel some hydrochloric acid. Immediately there is much fizzing and bubbling in the flask, showing that a gas is being given off. Two or three jars full of this gas are collected.

We see that it is invisible, so that it cannot be the same gas as we see when nitric acid dissolves lead. It has no smell, so that it is not the same gas as is given off when sulphuric acid dissolves copper. If we hold a lighted taper to a jar of the gas,

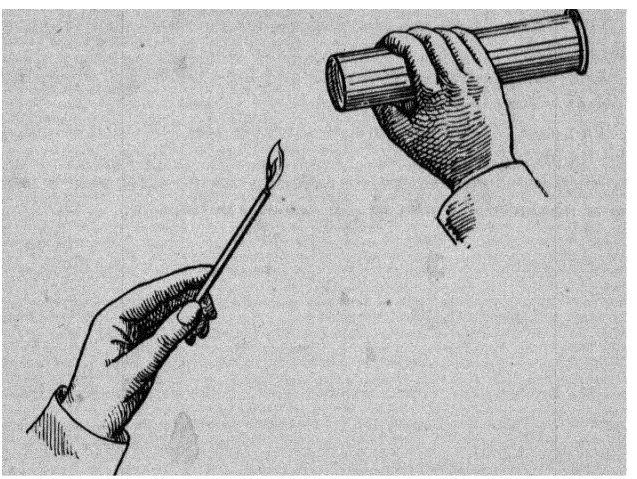


FIG. 54.

we see it is not hydrogen, for the gas does not burn, and the taper is extinguished. It is a very heavy gas, for if we pour a jar of it over a lighted taper, it extinguishes the taper just as surely as if water had been poured upon it.

Keep adding pieces of washing-soda to the acid so long as bubbles of gas are given off, and then filter the liquid, evaporate off a little of the water, and allow it to cool again. Large numbers of small crystals form in the liquid. These crystals are like crystals of common table salt. Taste them, and you will find that they *are* common salt.

So that you see that when washing-soda dissolves in hydrochloric acid, a heavy, colourless gas is given off. This is without smell, and neither burns nor allows a taper to burn in it. Also that common salt is found dissolved in the liquid after all the gas is given off.

SUMMARY.

Copper disappears on heating with sulphuric acid, and a *suffocating gas* is produced. *Sulphate of copper* or *blue vitriol* is found in solution.

The *copper is in the blue vitriol*, for zinc will turn it out and take its place.

Lead disappears in nitric acid, and a *poisonous gas* is produced. *Nitrate of lead* is formed.

The *lead is in the nitrate of lead*, for zinc will turn it out and take its place.

Washing-soda disappears in hydrochloric acid, and a *heavy gas* is produced. *Common salt* is found in solution.

LESSON XI.

HEAT.

Apparatus needed:—Balance, beakers, shot or sand; retort-stand and Bunsen burner or spirit-lamp; alcohol; a piece of metal or a small bunch of keys; thermometer.

If we place a hot body in contact with a cold one, the hot body gets colder and the cold one gets hotter, until both are at the same temperature. We explain this by supposing that something which we call *heat* has passed from the hot body to the cold one. The *thing* which has passed from one to the other is not a material thing—that is to say, it is not matter or stuff—because the body which was hot and is now colder is just as heavy as before, whilst the body which was cold and is now warmer is no heavier than before.

It is not necessary for the two bodies to be in contact. Two flasks of water, one hot and the other cold, may stand a short distance from each other on the table, and the cold flask will gain heat whilst the hot flask will lose heat.

And this is true of all bodies. If a number at different temperatures are near each other, the hotter ones lose heat and the colder ones gain heat, until all are at the same temperature.

Now, we cannot bring two bodies closer together than by mixing two liquids. Suppose, then, we take a quantity of hot water and mix it with an equal quantity of cold water. Every particle of the hot water is brought closely in contact with cold water, and therefore the opportunity is most favourable for heat to pass from the one to the other. The hand alone, without using the thermometer, will tell us that the mixture is hotter than the cold water and colder than the hot water. Let us carry out this experiment carefully.

Take two beakers, and placing them on the scale-pans of a balance, add small shot or sand to the scale-pan which contains the lighter beaker until the two are counterpoised. Now pour water into each beaker until each is about one-third full, and until each contains an equal quantity of water. Suppose there is 100 grams of water in each, and that the temperature of the water in each is 16° C. Place one beaker, with a thermometer in it, over the Bunsen burner or spirit-lamp until the water is heated to, say, 50° C.

Now pour the contents of the cold beaker into the hot one, and immediately pour the mixture back into the cold beaker, and at once take its temperature by noting the greatest height to which the mercury in the thermometer rises. It will be found to be as nearly as possible intermediate between 50° C. and 16° C.—that is to say, 33° C.

We explain this result by saying that the heat which the hot water gave to the cold, or the cold water received from the hot, was sufficient to reduce the temperature of the one

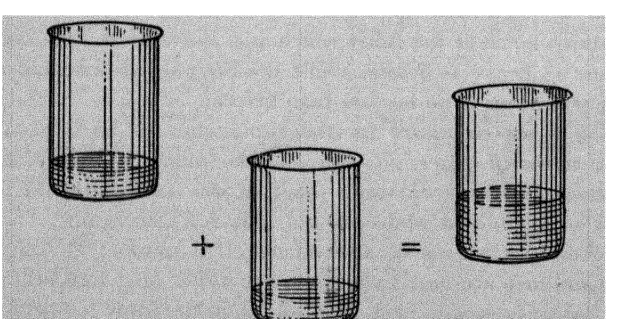


FIG. 55.—100 grams of water at 50° C. + 100 at 16° C. make 200 grams at 33° C.

from 50° C. to 33° C.—that is, 17° C.—and to raise that of the other from 16° C. to 33° C., which is also 17° C. And therefore we may say that it takes as much heat to raise the temperature of a quantity of water 17° C. as is given out when an equal quantity of water falls 17° C.

Other experiments may be tried with other equal quantities of water and other temperatures, and it will always be found

that the final temperature is just intermediate between the higher and lower temperatures. If we use other liquids we always get this same result. The hot liquid gives out heat and its temperature falls to a certain point, whilst the equal quantity of cold liquid receives this heat, and its temperature rises to the same point.

Now let us take unequal quantities of water at different temperatures—say, 50 grams of water at 45° C. and 100 grams at 15° C.—and mix them in the way already described. The

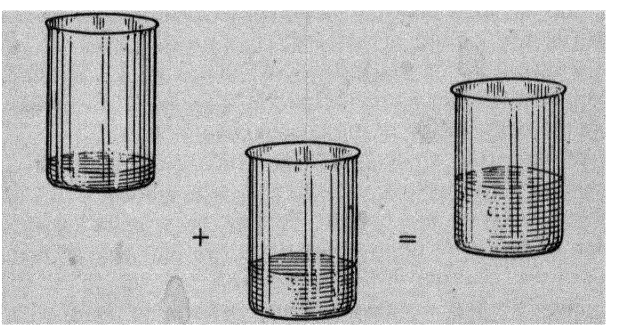


FIG. 56.—50 grams of water at 45° C. + 100 at 15° C. make 150 grams at 25° C.

final temperature will be found to be about 25° C. So that the heat given out by 50 grams of water in cooling from 45° C. to 25° C.—that is, in cooling 20° C.—is received by 100 grams of water, whose temperature it raises from 15° C. to 25° C., which is 10° C.

If we mixed 100 grams at 45° C. with 50 grams at 15° C., the temperature of the mixture would be about 35° C.—that is to say, that 100 grams of water in cooling 10° C. would raise the temperature of 50 grams of water 20° C.

Now, comparing these three experiments, we see that—

(1) The heat lost by 100 grams of water in falling 17° C. raised the temperature of 100 grams of water 17° C.

$$\text{Now, } 100 \times 17 = 100 \times 17$$

(2) The heat lost by 50 grams of water in falling 20° C. raised the temperature of 100 grams of water 10° C.

$$\text{And } 50 \times 20 = 100 \times 10$$

(3) The heat lost by 100 grams of water in falling 10° C. raised the temperature of 50 grams of water 20° C.

$$\text{And } 100 \times 10 = 50 \times 20$$

We might multiply experiments of this kind, and we should always find that the number of grams of hot water, multiplied by the number of degrees it falls in temperature, is always equal to the number of grams of cold water multiplied by the number of degrees it rises in temperature.

And we may conclude from these experiments that if we were to mix 1 gram of water at 16° C. with 1 gram at 14° C., the temperature of the mixture would be 15° C., for the heat given out by the warmer gram in cooling 1° C. would just raise the temperature of the cooler gram 1° C.

We take the amount of heat which must be received by 1 gram of water in order to raise its temperature 1° C., or the amount of heat which must be lost by 1 gram of water in order to lower its temperature 1° C., as our *unit of heat*, and we call this quantity of heat a *calorie.*

Then we may say that 2 grams of water in rising or falling 1° C. would receive or lose 2 calories; 50 grams of water in rising or falling 1° C. would receive or give out 50 calories; 20 grams of water in rising or falling 5° C. would receive or lose $20 \times 5 = 100$ calories; and so on.

Thus, returning to experiment (1), 100 grams of hot water in falling 17° C. gave out $100 \times 17 = 1700$ calories. These were received by 100 grams of cold water, and raised its temperature $\frac{1700}{100} = 17°$ C.

In experiment (2), 50 grams of warm water in cooling 20° C. gave out $50 \times 20 = 1000$ calories. This heat was received by 100 grams of water, whose temperature was thereby raised $\frac{1000}{100} = 10°$ C.

Now let us suppose that we have mixed together 50 grams of water at 80° C. with 75 grams of colder water, and that the temperature of the mixture was 50° C. What was the temperature of the colder water at first?

Heat given out by the hot water is $50 \times 30 = 1500$ calories. This is received by 75 grams of cold water.

Therefore the rise in temperature of the cold water was $\frac{1500}{75} = 20°$ C.

And therefore its first temperature was 50° C. − 20° C. = 30° C.

Again, suppose we mix 40 grams of water at 60° C. with 50 grams at 15° C., what will be the temperature of the mixture?

If $x°$ C. is the temperature of the mixture, the hot water will cool down $(60 - x)°$ C., and the cold will rise $(x - 15)°$ C.

$$\text{Then } 40(60 - x) = 50(x - 15)$$

and if you can work out this little equation, you will find that x stands for 35° C.

So far all these experiments have been done with water. A calorie is the amount of heat necessary to raise the temperature of 1 gram of water 1° C., or the amount which is given out when 1 gram of water cools 1° C. But you must not

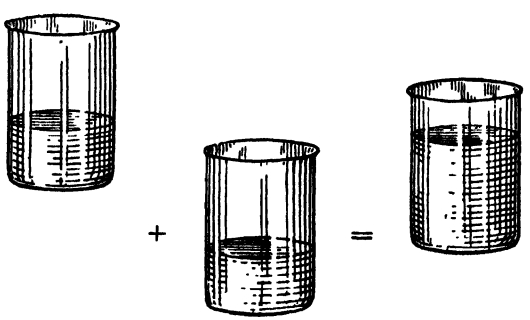


FIG. 57.—20 grams of water at 50° C. + 20 grams of alcohol at 16° C. make 40 grams of mixture at 40° C.

suppose that 1 gram of any other substance requires 1 calorie to raise its temperature 1° C., or gives out 1 calorie when its temperature falls 1° C.

Suppose we mix together 20 grams of water at 50° C. with 20 grams of alcohol at 16° C. If it were equal quantities of water we were mixing, we should know that the temperature of the mixture would be 33° C. We do know that the number of calories given out by the water in cooling is just that

amount of heat which is received by the alcohol and warms it. But the final temperature is not 33° C., but about 40° C., because 1 gram of alcohol requires not much more than one-half a calorie to raise its temperature 1° C.

It is the same with other substances. Weigh a piece of metal or a small bunch of keys, and hang them by a string in a vessel of boiling water. The keys will soon be at the same temperature as the water, namely, 100° C. Now place in a beaker a quantity of water equal in weight to the keys, and take its temperature. Suppose the temperature of the water is 16° C. Quickly transfer the keys to the cold water, stir gently with the thermometer, and note its highest reading. If we were mixing equal quantities of water at 100° C. and 16° C., we should find the temperature of the mixture to be 58° C. But in this experiment it was only 25° C., showing that the keys gave out only $\frac{3}{25}$ as much heat in cooling through 75° C. as water would have given out.

SUMMARY.

Heat passes from a hot body to a cold one. Heat is not matter.

When *equal masses of hot and cold water are mixed*, the *temperature* of the mixture is *intermediate.*

This is the case also with *other mixed liquids.*

When masses of the *same kind of liquid* are mixed, the mass of the *hotter liquid in grams multiplied by its fall in temperature* is equal to the mass of the *colder liquid in grams multiplied by its rise in temperature.*

The amount of heat given out by *one gram of water in falling one degree Centigrade*, or received by *one gram of water in rising one degree Centigrade*, is the *unit of heat*, and is called a *calorie.*

But *less heat than a calorie* is given out by a gram of any other kind of matter in falling one degree Centigrade.

LESSON XII.

HEAT ABSORBED IN THE FUSION OF ICE.

Apparatus needed:—Ice, beakers, balance; retort-stand and lamp; thermometer; ammonium chloride, common salt; photographer's hypo.

PLACE some ice in a basin, and allow it to stand in a warm room and melt. At the same time place some ice-cold water in another basin, and stand it near the ice. Have a thermometer in each basin. As the ice melts slowly the thermometer will show you that the temperature of the water resulting from the melting ice is 0° C., and that it remains at that temperature if you gently stir it with the thermometer until all the ice is melted. Not so the water in the other basin. That slowly rises in temperature. There is every reason for believing that as much heat is passing from the warm air of the room into the contents of the one basin as the other.

The result is the same even if you place the ice over the Bunsen or spirit flame, so that it is plainly receiving heat from the flame all the while the ice is melting. No rise of temperature over 0° C. occurs until all the ice is melted. What, then, becomes of the heat which passes into the mixture of ice and water? Plainly it is used in converting the ice into water without raising its temperature.

Prepare two beakers in the usual way by counterpoising them, and in one place some cold water and a few pieces of ice, in the other an equal quantity of water which you proceed to heat. Take the temperature of each, and when the ice is all melted, and whilst the water is still at 0° C., mix the contents of the beakers. The resulting temperature of the

mixture is intermediate between the temperatures of the hot and cold water, as you would expect.

Now counterpoise a beaker. About half fill it with water, weigh it to get the weight of the water, place a thermometer in it, and gently heat it. Whilst it is heating crush some ice into small pieces, wipe it as dry as possible with a clean duster, and wrap it in a piece of clean blotting-paper so as to keep it dry. The amount of ice taken should be about one-third the weight of the water.

By the time you have done this the water should be fairly hot. Note its temperature, place it on the table, and immediately tip the broken ice into the hot water. Stir with the thermometer, and the moment all the ice is melted read the thermometer again. Now, when you have again weighed the beaker and its contents, in order to find out how much ice was used, you have all the figures necessary to find out how much of the heat of the hot water was used up in melting the ice (Fig. 58).

Here are the figures of an actual experiment.

The beaker was counterpoised with small shot.

Weight of hot water = 160 grams.
weight of water and ice = 220 „
therefore the weight of ice used = 60 „
Temperature of hot water was 75° C.
Final temperature of water was 30° C.

Now, 160 grams of water in cooling from 75° C. to 30° C. give out $160 \times 45 = 7200$ calories.

This heat was used in melting 60 grams of ice at 0° C. into water at 0° C., and in raising the 60 grams of water at 0° C. to 30° C.

To raise 60 grams of water from 0° C. to 30° C. required $60 \times 30 = 1800$ calories.

The remainder of the heat given out by the hot water, namely, $7200 - 1800 = 5400$ calories, was used solely in melting 60 grams of ice.

If it takes 5400 calories to melt 60 grams of ice, it should take $\frac{5400}{60} = 90$ calories to melt 1 gram of ice.

This result is not quite correct, because some of the heat escaped into the air. To prevent this escape of heat requires elaborate precautions. When these are taken, it is discovered that 80 calories are needed to turn a gram of ice at 0° C. into a gram of water at 0° C.

We therefore see that it takes eighty times as much heat to melt ice without raising its temperature at all, as it does to

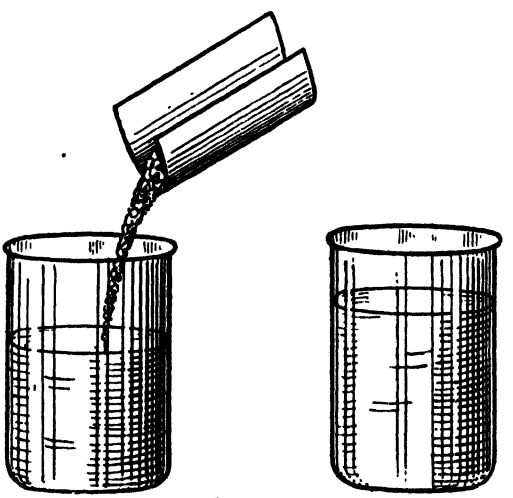


FIG. 58.—160 grams of water at 75° C. + 60 grams of ice at 0° C. make 220 grams of water at 30° C.

raise the temperature of an equal mass of water 80° C. One of the consequences of this is that ice is very long in melting when the warm weather comes after heavy frost. The sun may shine and the warm breezes blow, but day after day the thick ice on the lakes and rivers only slowly disappears, and in the mean time the water is still icy cold.

On the other hand, since it takes 80 calories to melt a gram of ice, a gram of water at 0° C. must lose 80 calories before it can freeze. On the approach of winter, therefore, the air may be icy cold some time before the water of the ponds, and especially the rivers, lose heat enough to be able to freeze.

So you see that the winter is delayed as well as the spring

by the very large amount of heat water must lose before it can freeze, and the equally large amount ice must receive before it can melt. This is one of the reasons why the sea on our shores and the water in our deepest lakes seldom freeze, even in our coldest winters, and why the large masses of ice and snow on our highest mountains never thaw, even in the hottest summers.

The quantity of heat necessary to melt a gram of ice at 0° C. without raising its temperature is called the *latent heat of the fusion of ice.*

Other solids besides ice require heat to change their state from the solid to the liquid, but of all solids ice requires the most. Thus a gram of silver must have 21 calories to melt it, a gram of sulphur nearly 10, and a gram of lead little more than 5.

It is equally the case that heat is absorbed when a solid is converted into a liquid, not by melting, but by dissolving in a solvent liquid. If you add a spoonful of common salt to a little water and stir to dissolve it, the thermometer will immediately show that the solution is colder than the water was before it dissolved the salt. Heat was necessary to dissolve the salt, so it was taken from the water. It is the rule that in all such cases of a solid body becoming liquid without the aid of heat supplied from outside, that some of its own heat becomes *latent*, and it is colder than before.

A much greater fall in temperature occurs when a salt called ammonium chloride is dissolved in water. Dissolve 50 grams of this salt in 100 grams of cold water, and note the very great fall in temperature which occurs.

By mixing together ice and substances which dissolve in water we make what are called *freezing mixtures*—that is, mixtures which are below the freezing temperature of water, and so cold that if small vessels containing water be immersed in them the water freezes.

To make a freezing mixture we take snow or pounded ice, and mix it with about one-third its mass of common salt. There is always some water attached to ice, and in this the salt begins to dissolve, and then some of the ice dissolves in

this salt and water, and so the process goes on, the heat necessary to convert both the ice and the salt to liquids being taken from themselves and from surrounding objects.

By means of this mixture, Fahrenheit obtained that degree of cold which he marked 0° on his thermometer-scale, because he thought there could be nothing colder. He then placed his thermometer in melting ice, and marked on the thermometer-stem the height to which the mercury rose, and afterwards in contact with his own body, and again marked the height of the mercury. He found that from his mark of 0° the melting-point of ice was about one-third the distance up the stem to the temperature-mark of his body-heat, which he called blood-heat. He wished at first to call blood-heat 100°, but that would have introduced a fraction into the melting-point of ice, which would have been $33\frac{1}{3}°$ in that case. So he called blood-heat 96°, and then the melting-point of ice came to be 32° F. He thought blood-heat was a fixed point; we know now that it is not.

Other substances besides common salt may be used for making freezing mixtures with broken ice or snow. Thus ammonium nitrate, or the salt called by photographers hypo-sulphite of soda, or hypo, will give much lower temperatures than can be got by using common salt.

The heat which is necessary to liquefy a solid reappears when the liquid is reconverted into a solid. Here is a very interesting experiment. Take a small flask, and fill it about one-third full of crystals of hypo. Place it over the Bunsen flame or the flame of a spirit-lamp, and gently warm it until the crystals have completely melted in their own water of crystallization. Then place it aside in a quiet place to cool. So long as it is left alone it remains liquid, even when cooled down to the temperature of an ordinary room. If, however, it be agitated, or if a crystal of hypo, however small, be dropped into it, the whole mass immediately solidifies. At the moment of solidification, the temperature of the whole rises several degrees.

This rise of temperature may be seen by placing into the cooled liquid hypo a thermometer with a tiny crystal of hypo

attached to its bulb. The thermometer shows the rise in temperature, but is held fast in the solidified salt. Do not try to wrench it out, but pour water into the flask, and take out the thermometer when the hypo has dissolved and set it free.

It does not always happen that when a solid dissolves in a liquid the solution is colder than before. Thus the white solid called caustic soda dissolves in water, but the solution is hotter than before. It is true that a certain amount of heat is used up to convert the solid caustic soda to the liquid form, but a very much larger amount is produced in this case by the union of the solid and water. You will remember that water unites with quick-lime, producing much heat, and this is the same kind of thing exactly.

SUMMARY.

The *heat* which passes into melting *ice does not raise its temperature. It is used up in turning the ice into water.*

Pour hot water on ice, and the *resulting temperature* of the mixture is *not exactly intermediate.*

It takes 80 *calories to melt a gram of ice.*

A gram of water must lose 80 *calories before it is frozen.*

When a *solid is turned into a liquid* by *dissolving, heat is absorbed* and the solution is cooled.

LESSON XIII.

HEAT ABSORBED IN THE CONVERSION OF WATER INTO STEAM.

Apparatus needed:—Thermometer, flask, retort-stand, burner; piece of thin muslin and a bicycle-pump; alcohol and ether.

It is hardly necessary to remind you that if you place some water in a flask and allow a thermometer to hang so that its bulb is just above the surface of the water whilst you heat it, the water gets hotter and hotter during the process of heating; but once having boiled, when the thermometer marks a temperature of 100° C., the water gets no hotter, although heat still continues to pass into it at the same rate as before. What, then, becomes of the heat which the water receives after it boils?

The water boils away. It becomes less and less as it boils, and if you allow the boiling to continue long enough, all the water will disappear. You know what becomes of it. You know quite well that the water is being converted from a liquid into a gas, and that this gas, or water-vapour, or steam as it is often called, may be conducted into a cool place, where it is turned back again into water. All this you know because you have seen it done.

And just as the heat which passes into ice does not raise its temperature until all the ice is melted, but is used to turn the solid ice into the liquid water, so the heat which passes into boiling water does not raise its temperature, but is used to turn the liquid water into the gaseous steam.

This heat, which becomes hidden as it were in the steam, is called the *latent heat of vaporization.* You will see later

that when steam is turned back into water all the heat which was used up to vaporize it—that is to say, its latent heat—reappears.

Other liquids have also their latent heats of vaporization. Thus alcohol, when boiling at 78° C., gets no hotter. All the heat which passes into it after that is used up in turning it into vapour, and this latent heat becomes sensible again when the vapour is reconverted into a liquid.

But water and many other liquids are converted into vapours at the ordinary temperature, and without being heated. In this case is heat necessary to effect their vaporization? Think a little.

It is a hot summer's day. The air is still, the roads are dusty, and every living thing is suffering from the heat. It is too hot for exertion, and the most pleasant thing to do is to lie down by the river's bank, or the seaside, for it is coolest there. Why is it cooler there than elsewhere? Perhaps because the water is steadily evaporating, and is therefore cooler than the land. If the water in evaporating requires heat, and you do not supply it, the heat may be taken from the water itself.

Presently a thunderstorm occurs, and for a short time there is a heavy downpour of rain. Then it clears again, and the air is deliciously cool. Is this because the fallen rain is evaporating into the air, and so cooling the surface of the ground?

In India and other hot countries they use unglazed earthenware vessels for holding drinking water, and after filling them they are put in a draught of air. Some of the water passes through the pores of the vessel to the outside, where it is evaporated, the necessary heat coming from the vessel and its contents, which are therefore cooled. In some countries also it is the practice to wrap damp cloths around bottles of wine, which are cooled by the evaporation of the water on the cloths.

It is not at all difficult to show that evaporation brought about in this way produces a lowering of the temperature. You take a thermometer, and wrap a piece of thin muslin or other fabric over its bulb. When the muslin is wetted with

water at the temperature of the room, the thermometer falls a little. The reason is that the water in the bulb evaporates, and in so doing takes heat from the thermometer, which is then colder than before. If you blow on the wet muslin with a pair of bellows or a bicycle-pump, the thermometer will fall more quickly and to a greater extent.

But why should blowing on the wet bulb have such an effect? You may blow on the dry bulb of the thermometer as long as you like, and the mercury in it will not contract to show that it is colder than before. The reason is this. When a liquid evaporates the vapour collects over the top of the liquid and around it. This space soon becomes saturated with vapour, which is the same thing as saying that it soon has as much vapour from the liquid as it can hold, and in that case no more vapour is formed. If now this vapour be blown away more vapour forms, and in this way you see that blowing at the liquid helps the evaporation, and therefore carries off more heat in a given time.

The water in wet linen which is hung out to dry passes off more quickly, and therefore the linen dries more quickly on a windy day than when the air is still. You sometimes cool your hot tea by blowing on it. The farmer is glad of a windy day

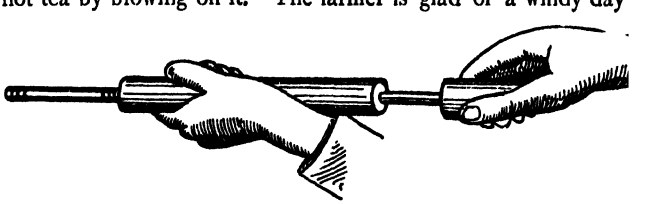


FIG. 59.

in summer to dry his grass into hay. And on a hot day a gentle breeze, although the air may be very warm, is very refreshing, because it helps to evaporate the water which is always on the outside of our bodies, and so it cools us very pleasantly.

But there are other liquids which evaporate much more readily than water does, and produce a greater amount of cooling, and some which do not evaporate at all, and therefore

produce no cooling effect. Alcohol and ether belong to the first class of liquids, and some oils, such as olive oil, to the second class.

Place on the table or on a slate separate drops of ether, alcohol, water, and olive oil. The ether is evaporated first, the alcohol next, the water next, and the olive oil not at all.

Pour a little ether on the palm of your hand. It quickly evaporates, and has a very cooling effect. Alcohol has the same but to a less extent. When your hand is wet with water, wave it in the air, and it will be cooled. But olive oil on the hands will not cool them at all, because it does not evaporate.

Here is a very interesting experiment. You take a small dish—an evaporating-dish or little porcelain crucible, or anything of that kind—and stand it in a little pool of water on a small block of wood. Into the dish pour enough ether to fill

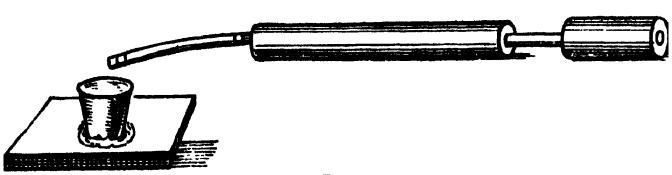


FIG. 60.

it to about one-third of its depth. Then with a pair of bellows or a bicycle-pump, or in some other manner, blow a stream of air across the top of the dish. This will remove the vapour of ether as fast as it is formed, and so allow a rapid evaporation of the ether to go on. To bring about this evaporation, the heat which becomes latent in the ether-vapour is taken from the ether, the dish, and the pool of water in which it stands. At last so much heat has been taken from the water that it freezes into solid ice, and if the dish be lifted, the ice is found to have cemented the block of wood to it.

This experiment may be performed in a slightly different manner by standing a test-tube containing a little water in a beaker containing some ether. A current of air is blown briskly through the ether, and by the swift evaporation of the ether the water in the test-tube may be frozen into a solid

block. But this manner of freezing the water uses up much more ether than the other.

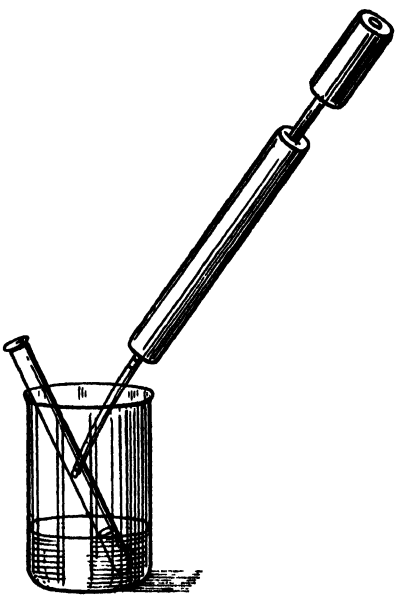


FIG. 61.

There is a very interesting instrument called the cryophorus, which is usually made of the shape shown in Fig. 62 or 63.

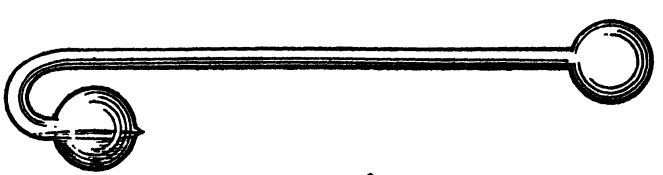


FIG. 62.

It is made in this manner. The two bulbs are blown on the tube of glass, one bulb being closed up, and the other having a small hole left in its extremity. It is now partly filled with water, and the water boiled in the closed bulb. After it has

boiled for some time, and when it is judged that all the air has been expelled by steam, the little hole is sealed up by melting

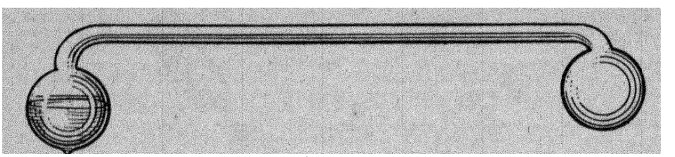


FIG. 63.

the glass round it in a very hot flame. Then the cryophorus contains nothing but water and the vapour of water.

To use it, all the water is allowed to run into one of the bulbs, and the empty bulb, or rather the bulb which contains nothing but water-vapour, is placed into a vessel containing a freezing mixture of ice and salt. This condenses the vapour in the bulb to a liquid, and fresh vapour to take its place is immediately produced from the water in the other bulb. This goes on so rapidly that in a short time the water has lost so much heat that it is frozen to solid ice—the heat, of course, being used up in converting some of it to vapour. Thus you see that placing one bulb of the cryophorus in a freezing mixture causes the water in the other bulb to freeze. The same effect would be produced if the vapour from the water were rapidly removed in any other manner.

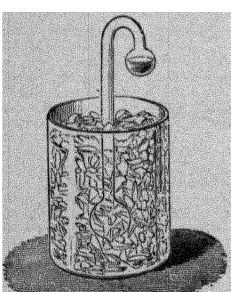


FIG. 64.

There is an instrument which is commonly called the *wet- and dry-bulb thermometer*, which indicates the state of the air with respect to its temperature and the amount of water-vapour it contains, by the lowering of temperature produced by the evaporation of water upon one of the thermometer-bulbs. You remember the experiment described in the early part of this lesson of tying a piece of wet muslin over a thermometer-bulb, which had the effect of causing the thermometer to indicate a lower temperature than before.

The wet- and dry-bulb thermometer consists of two similar thermometers mounted side by side, and a little distance apart, on the same board. The bulb of one of them is uncovered; that of the other has a piece of muslin tied over it and a thick string from the muslin dipping into a little vessel of water. The water rises up the string by capillary attraction, and keeps the muslin constantly moist.

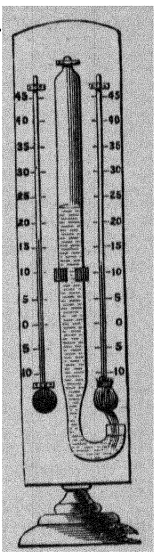


FIG. 65.—(From Ganot's *Physics*.)

As the water evaporates from the muslin it subtracts heat from the thermometer in doing so, and the latter thereby shows a lower temperature than the other dry thermometer. If the air already contains nearly as much water-vapour as it can hold, the evaporation from the muslin is slow, and the fall of the mercury in the thermometer is very slight, so that the two thermometers show nearly the same temperature. If the air is dry—that is to say, if it contains much less water-vapour than it can hold, and is therefore far removed from saturation—the evaporation from the muslin is rapid, and the wet-bulb thermometer shows a temperature much lower than the dry-bulb thermometer. In this way the hygrometric state of the air is indicated.

SUMMARY.

After boiling, liquids get no hotter. All the heat which passes into them is used up in turning the liquids into vapour.

When *any liquid is turned into vapour*, even *at the ordinary temperature, it loses heat*, and becomes cooler.

LESSON XIV.

MORE ABOUT THE LATENT HEAT OF VAPORIZATION.

Apparatus needed:—Distillation apparatus and thermometer; beakers, steam-trap; balance.

THERE are more ways than one of converting the vapours of volatile liquids like water, alcohol, ether, etc., back to the liquid state. The simplest and commonest, and the one method of which you already have some knowledge, is that of taking heat from them—that is, of reversing the process which is ordinarily made use of in converting them into vapours. Experiments on distillation have made you familiar with this method, but there is something else about it which you must now learn. Let us begin with an experiment.

Fit up a flask with a good cork and a tube passing through it bent twice at right angles, as shown in Fig. 66. Half fill the flask with water, and set it to boil over the Bunsen burner or spirit-lamp. Also get ready a beaker, half fill it with cold water, gum a strip of stamp paper on the outside of the beaker at the exact height at which the water stands, and another strip about $\frac{1}{4}$ inch above it. Place a thermometer in the beaker, and note the temperature of the water.

When the water in the flask is boiling and steam is issuing freely from the tube, place the beaker so that the end of this tube is in the water and near the bottom of the vessel. In this way you let the steam enter the cold water and condense there. When so much steam has condensed that the level of the water has risen to the upper strip of stamp paper, take out the tube, and again note the temperature of the water. There will be a very considerable rise in its temperature.

Now pour away this water, and fill the beaker up again with cold water to the lower mark. Take the temperature, which will probably be the same as at the commencement of the previous experiment. Boil some water in another vessel, and when boiling pour enough of it into the cold water to bring its

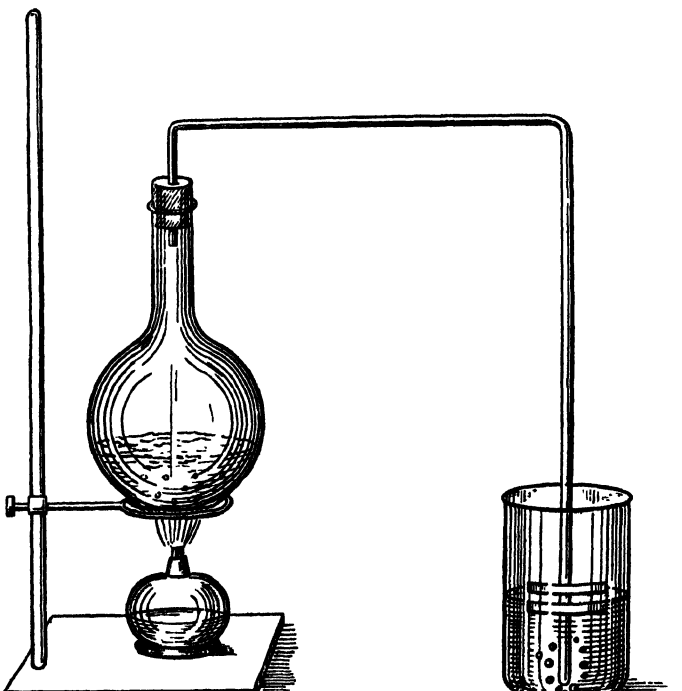


FIG. 66.

level to the upper mark. Then stir, and note the highest temperature reached. This will be very much below that attained when steam was used.

How does this difference come about? The steam had a temperature of 100° C., and the boiling water had a temperature of 100° C., and equal masses of each were added to equal masses of cold water. Yet the result was very different.

Here are the figures of an experiment actually made. The cold water was at a temperature of 16° C. Boiling water added to it raised the temperature of the cold water to 25° C., and at the same time lowered its own temperature to 27° C. But a mass of steam, equal to that of the boiling water used, raised the temperature of the cold water to 85° C., and at the same time *changed its own state from a gas to a liquid*, and lowered its own temperature to 85° C. It seems quite clear from this that the very much larger quantity of heat given out in the steam experiment was due to the change from the *gas*, *steam*, to the *liquid*, *water*.

We believe that just as much heat is given out when steam at 100° C. condenses to water at 100° C. as is necessary to convert the same quantity of water at 100° C. into steam at 100° C. This is the *latent heat of vaporization of water*.

We can find out the value of this latent heat by an experiment made in the following manner:—

Steam is produced by boiling water in a flask provided with an exit tube for the steam. But this tube should have attached to it a contrivance known as a steam-trap, to catch any water condensed in the tube through which the steam passes. It is impossible to prevent some amount of condensation in the tube, and if this condensed steam were allowed to pass into the beaker of cold water, it would vitiate the result, because the heat given out by its condensation would not pass into the cold water of the beaker.

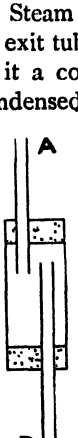


FIG. 67.—(From Watson's *Elementary Practical Physics.*)

The steam-trap may be made out of a piece of large-sized glass tubing about 3 inches long. This is fitted with a good cork at each end, and through each cork is bored a hole. The tube A from the flask in which the water is boiled just passes through the upper cork, whilst the tube B, which is a short one, through the lower cork reaches nearly to the upper cork. In this piece of tubing collects any water condensed in the tube, and it is thereby prevented from passing into the cold water.

A beaker or flask is taken, counterpoised with small shot

on the scale-pan of the balance, half filled with cold water, and carefully weighed. Its weight is noted, as well as its temperature, by a thermometer placed in it.

When the water in the flask C is boiling, and steam is issuing freely from the tube, the beaker or flask A of cold water is placed as in the previous experiment, so that the steam issues from the tube D near the bottom of the water. It is

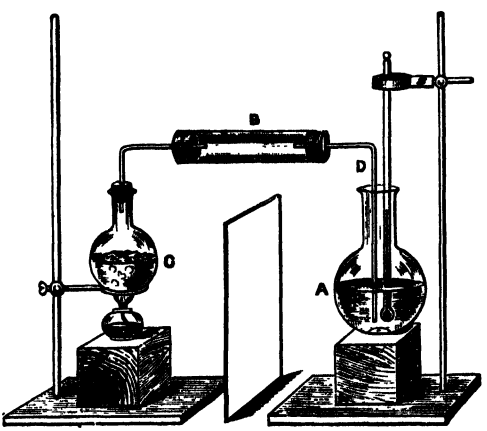


FIG. 68.—(From Wright's *Elementary Sound, Light, and Heat.*)

better if the vessel stands on and is wrapped in cotton wool to prevent loss of heat. Stir gently with the thermometer, and when the temperature has risen to about 60° C., stop the experiment, and weigh the vessel again to find out how much steam has been condensed. In this figure is shown a horizontal steam-trap B.

Here are the figures of an experiment actually made—

Mass of cold water at 16° C. = 150 grams
Mass of steam at 100° C. = 10 „
Final temperature is 52° C.

Thus 150 grams of water at 16° C. have risen in temperature to 52° C.—that is, through 36° C.—and must therefore have received $150 \times 36 = 5400$ calories.

This heat is given out by 10 grams of steam at 100° C. in

condensing to 10 grams of water at 100° C., and falling to 52° C.—that is, in falling through 48° C.

10 grams of water in falling through 48° C. give out $48 \times 10 = 480$ calories.

The remainder of the heat received by the cold water, namely, $5400 - 480 = 4920$ calories, must have been given out by 10 grams of steam in condensing.

Therefore every gram of steam in condensing gave out $\frac{4920}{10} = 492$ calories before its temperature fell at all.

This result is not quite correct, because some of the heat of the condensing steam passed into the beaker, and some of it no doubt passed into the air. There are methods of allowing for this heat which is lost, but they need not be explained here. When the most careful and accurate experiments are made the latent heat of vaporization of water is found to be nearly 540—that is to say, it takes as much heat to convert 1 gram of boiling water into steam without raising its temperature as it does to raise the temperature of 1 gram of water 540° C., or 540 grams of water 1° C.

Or, putting it in another way, 1 gram of steam at 100° C. in condensing to water at the same temperature gives out as much heat as would raise the temperature of 540 grams 1° C.

This is a very large amount of heat. It explains why it is that a person scalded with steam is injured much more seriously than if scalded with water equally hot. It also explains why it is that one may pass steam into cold water and raise its temperature so much that at last it may be brought to the boiling-point. It is quite common in some industries, where large quantities of liquids are boiled at once, to do this by passing steam through a coil of pipes which is placed in the boiling-pan.

To the latent heat of water-vapour our country owes to a great extent its mild and genial climate. For the winds which most prevail here are from the west and south-west. They come, after passing over thousands of miles of ocean, laden with water-vapour, which is largely condensed over our land, and falls as rain. Every gram of vapour in condensing gives out hundreds of units of heat, which go to warm the air and the land beneath

it. The consequence is that our climate is much milder than that of other countries in the same latitude, and not subject to such violent extremes of heat and cold as they are.

The late Professor Tyndall calculated that every pound of water-vapour in our atmosphere which is condensed to a liquid gives out heat enough to raise the temperature of 5 lbs. of cast iron to its melting-point. You have some idea of how hot cast iron must be before it melts.

What, then, would our climate be without those kindly winds laden with heat-giving moisture? They bring us the heat of the tropics as well as the rain which is so necessary for us.

SUMMARY.

When *vapours condense to liquids*, they *give up* just that amount of *heat* which was absorbed in vaporizing them.

A *gram of steam* at 100° C. will *give out* 540 *calories* in condensing to *water* at 100° C.

LESSON XV.

THE RUSTING OF METALS.

Apparatus needed:—Bell-jar or pickle-jar, small dishes, sulphuric acid, and pieces of bright iron and steel; small bottle and vaseline; knitting-needle; pieces of copper plate and lead; iron spoon and Bunsen burner; magnesium ribbon, small crucible with lid, balance and weights.

WHEN a piece of brightly polished iron or steel—which is a kind of iron—is left exposed to the air, sooner or later it becomes, first dull, and then covered over with a fine red dust which we call iron rust. The action does not cease with the formation of a single layer of rust, but the rust falls off in layers as it is formed, until all the iron is converted into red rust. There is no longer any of the metal left, and we say that it has all rusted away. What was bright hard metal is now represented only by a pile of dull red powder.

It is a matter of everyday knowledge that this takes place more quickly out-of-doors than indoors, more quickly in damp air than in dry air, and often more quickly buried a short distance below the surface of the ground than on the surface.

If it is desired to preserve iron from this rusting process, it is necessary to protect its surface from contact with the atmosphere. Small bright iron and steel things must be frequently covered over with a thin layer of oil or grease of some kind to keep the air from them, and this layer must be frequently renewed.

Iron and steel are so largely used nowadays for constructing those parts of large buildings, bridges, and ships which have to withstand great forces and carry heavy loads, that every

precaution must be taken that they do not slowly rust away to useless and valueless powder. The usual way of protecting these iron and steel structures from rusting is to encase them in cement or some other kind of plaster, or to cover them and keep them covered with oil paint. By these means the atmosphere is kept from the metal, and it does not rust. When the great Forth railway bridge—which is built of steel, and cost more than a million pounds to build—was opened, the

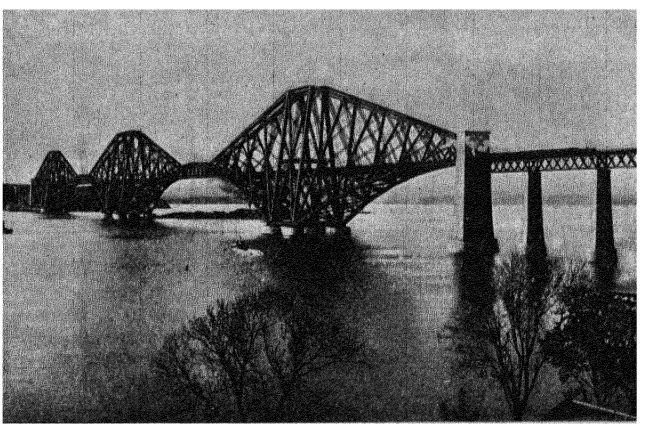


Photo.] [*Valentine & Sons.*

FIG. 69.—The Forth Bridge.

builder said that if the owners of this costly structure did not want their property to fall to powder into the waters of the Forth in a few years, they must keep it well painted. Accordingly, men are kept, nearly all the year round, painting and repainting the parts of this bridge in order to keep the air from the metal. Now, paint is made of oil which has had stirred into it some colouring matter. The oil dries in a thin layer which sticks closely to the surface of the painted object, and for a time forms a barrier between the object and the air. But only for a time, because the layer of oil perishes, and the object must be again painted.

It is, then, well known that iron and steel rust when exposed to the air, and that if covered over with a layer of something

which will keep the air from them no rusting will take place. It is certain, therefore, that the atmosphere is concerned in the rusting of iron.

Iron does not rust in dry air. Take a large bell-jar, or a glass preserve-jar, and invert it over a plate of glass. See that the jar and plate are quite dry. Place on the plate and underneath the jar two small evaporating or other dishes, and in one place a little strong sulphuric acid, and in the other some pieces of iron wire which have been rubbed quite bright, and some steel needles. If the bottom edge of the jar, and the plate

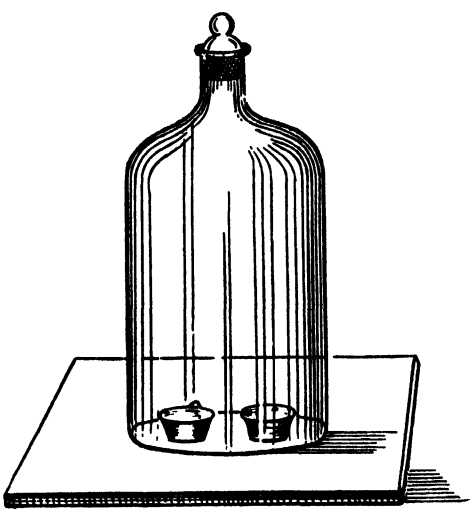


FIG. 70.

where it is in contact with the jar, are smeared over with some thick grease such as vaseline, no outside air can get in, and the air inside will be kept dry, because the sulphuric acid will absorb whatever moisture there may be in it. At the same time expose to the atmosphere out-of-doors some similar pieces of bright iron wire and steel needles. In the course of a few days these will be rusty, whilst those inside the jar will be bright after many weeks.

Water free from air will not rust iron. Boil some water in

a vessel for a quarter of an hour, until all the air which it contains in solution has been driven out of it. Fill a small bottle quite full of the boiled water, and drop into it pieces of bright iron wire and a few steel needles. Cork the bottle tightly, and melt paraffin-wax, or sealing-wax, or wax candle on and around the cork to keep out air, or cover the cork with vaseline. Lay the bottle aside. At the same time place similar pieces of metal into unboiled water contained in a bottle, and after corking it tightly, protect it from air in the same manner as before. In a few days the iron wire and needles in the unboiled water will begin to rust. At the end of months those in the boiled water will still be bright.

Therefore, from these experiments we see that iron and steel rust only when exposed to both air and water. Air alone will not rust them, nor will water alone, but the help of both is needed.

Take a bright knitting-needle, place one end of it in the fire, and keep it red-hot for a few minutes, or make and keep it red-hot in the flame of a Bunsen burner or spirit-lamp. When cool it will have lost its brightness and be covered over with a blackish layer. If it be heated over and over again layers of this black rust will be formed, and may be knocked or scraped off, and this may go on until all the iron has been converted into rust. You may see plenty of this black rust in the blacksmith's workshop. The sparks which fly off when the red-hot iron is hammered are composed of it. A poker which is much used shows this kind of rusting away very well. But there are methods of making a piece of bright iron red-hot in a vessel which is quite deprived of air. When this is done the metal retains its brightness; no black rust is formed upon it.

Iron is not the only metal which rusts in air. Almost all of them do so to a greater or less extent, especially if they are heated somewhat. Brass and copper and lead require to be frequently rubbed in order to remove the thin layer of rust which forms upon them. When this is done the metal has its brightness renewed. If this is not done they remain covered over with a dull layer of rust, although, unlike iron, this layer does not continually fall off, thus exposing fresh surfaces of

metal to the air to be rusted in turn until all the metal is converted into rust. They will, however, rust more quickly if they are much heated.

Those metals which were known as the noble metals do not rust at all readily even if heated strongly. These are gold, platinum, silver, with some other rarer ores. Gold is quite unchanged in pure air; platinum and silver dull very slightly in pure air.

Take a small piece of copper plate or a penny, clean it until it is brightly polished, and then heat it over the flame. It becomes covered over with a close layer of a black substance, which is the rust of copper. If it had been heated in the absence of air no black rust would have formed upon it. Careful weighing would show that the rusted copper weighs more than the bright copper did before heating. So that it appears that the copper attaches to itself some other matter when it rusts.

Place some pieces of bright lead in a small open crucible or

FIG. 71.

in an iron spoon, and heat them over the flame until they melt. At first the surface of the melted metal is bright and silvery, but very soon it becomes clouded over with a dirty-looking

scum. This is the rust of lead. The surface layer may be scraped aside with an iron wire, when the molten metal beneath will be exposed, only, however, to be covered over again with more rust. If the rust is continually scraped aside as fast as it is formed, a large amount of lead may be rusted in the course of a few minutes. It is easy to show that there is an increase of weight when lead is rusted. You have only to weigh the crucible with 3 or 4 grams of lead in it before melting, and afterwards to see that the increase is considerable. This is not the case when lead is melted in the absence of air. Therefore it is clear that something from the air attaches itself to lead when it rusts, and so increases the mass.

Even mercury, which does not change at all in cool pure air, and always appears so bright and silvery, rusts when heated. It is a very slow process; but if mercury be heated in air, a rust, which is a bright red-coloured powder, appears on its surface. This does not take place in the absence of air, and the rust, which is formed in the manner described, weighs more than the mercury out of which it was formed.

Some metals—of which magnesium is one—burn brightly when heated in air, leaving a white ash, which is the rust of magnesium. A piece of magnesium ribbon, held in the tongs, may be lighted by a match, and when it has burnt with a dazzling light, its white powdery ash is evident. It is easy to show that this ash weighs more than the unburnt magnesium. First counterpoise a small porcelain crucible, with its lid, on the balance, and then place about a gram, or less, of magnesium ribbon in the crucible, and weigh the metal. This should be coiled up, not too tightly, but so that it will lie at the bottom of the crucible. Place the crucible, covered with its lid, on a piece of wire gauze, and heat it strongly with the Bunsen or spirit flame. Presently the metal catches fire and blazes. The lid must

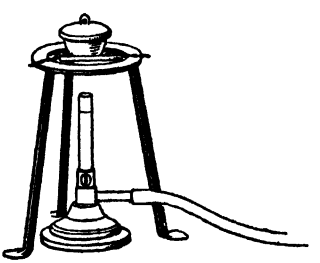


FIG. 52.—(From Tilden's *Hints on the Teaching of Elementary Chemistry.*)

be lifted a little now and then to allow air to get in—for the magnesium cannot burn without air—but not so much as to allow the white fumes which are produced to pass out. When all the metal is burnt weigh the crucible again after it has cooled. It will have increased in mass. The magnesium must have attached to itself some other matter when burning, or when rusting.

We may, then, state as a general thing that when metals rust they must be in contact with air, and that the rust weighs more than the metal which was rusted.

SUMMARY.

In contact with air and water, *iron* is converted into a *red rust.*

To keep iron from rusting, it must be coated with something to protect it from the air.

Iron made *red-hot* becomes covered with a *black rust.*

Heated copper is covered with a *black rust.*

Melted lead becomes covered with a *yellowish rust.*

Mercury heated is converted to a *bright red rust.*

Magnesium burns brightly in air, forming a *white rust.*

The rust is always heavier than the metal from which it is formed.

LESSON XVI.

MORE ABOUT THE RUSTING OF IRON.

Apparatus needed:—Balance, small horse-shoe magnet, reduced iron; muslin and clean iron filings; straight lamp-glass and cork, gas-jar, stout copper wire; deep dish; taper.

THERE is a kind of iron in very fine powder, often called *reduced iron*, which can be got at the dealer's. If this is heated strongly enough in absence of air it melts, and is then like the ordinary metal; but if it be heated in air it takes fire, and smoulders until the whole of it is burnt.

The following very interesting experiment can be performed

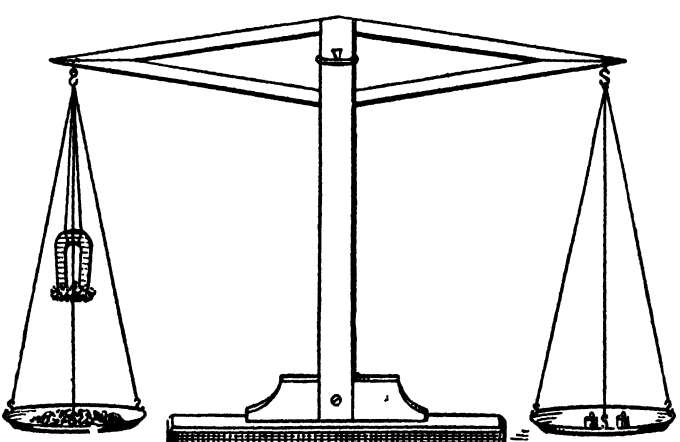


FIG. 73.

with reduced iron. You place a small dish on the pan of a balance, and hang a small horse-shoe magnet from the little

hook, if there is one, over the pan. Into the dish place some of the reduced iron, and counterpoise the whole of it with shot or weights.

Then lift up the dish to the magnet, and some of the powder will adhere to it. Whatever may fall off is caught in the dish. Hold a Bunsen or spirit flame for a moment to the iron on the magnet. It catches fire and smoulders, forming a blackish-brown rust, and that arm of the balance sinks, showing that there is an increase in weight. Reduced iron will not burn, and consequently will not increase in weight, if you heat it in the absence of air.

We must now find out whether there is also an increase in weight when ordinary iron forms its reddish-brown rust at ordinary temperatures. Take some bright, clean, dry iron filings—it is essential that they be not greasy—and tie them up in a thin muslin bag. Weigh the bag and the filings together, and note down the weight. Now place the bag in distilled or rain-water for a few minutes until thoroughly soaked, allow it to drip, and hang it up in the school or in the outer air. In a few days the bag and its contents should be dry again. If they are not so, they should be artificially dried in a moderately warm oven, or should be placed in a draught of air where the water will be quickly evaporated. When sure that the filings are quite dry all through, weigh the bag and its contents again, and an increase in weight will be noted. Something, then, must have become attached to the iron during the process of rusting, and thereby increased the mass of matter.

Where does the iron get this additional matter, and what is it? These are questions which now suggest themselves and require an answer. We can make the iron tell us by making use of suitable experiments.

Get a straight lamp-glass—one of those which are of the same diameter throughout—and get also a cork which will fit one end tightly. Pass a piece of wire through the cork, hook its inner end, and attach to it a little muslin bag of bright, clean, damp iron filings. When the cork is placed in position in the end of the lamp-glass, the bag of filings will hang freely in the air enclosed within the glass (Fig. 74).

Get a rather deep dish, half fill it with water, and place the open end of the lamp-glass in it. You should place melted paraffin-wax, or candle, or sealing-wax over and around the cork so as to make it air-proof.

If, when the glass is in position, the water inside it does not stand at the same height as that in the outside dish, they must be brought to do so. Probably the level of the water inside will be lower than that outside. This can be corrected by slipping

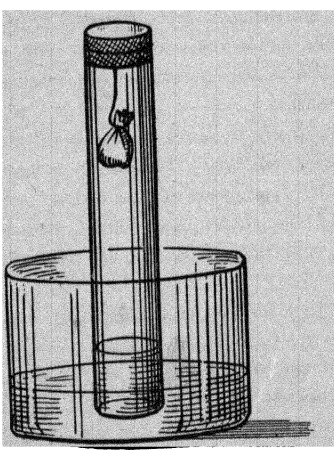


FIG. 74.

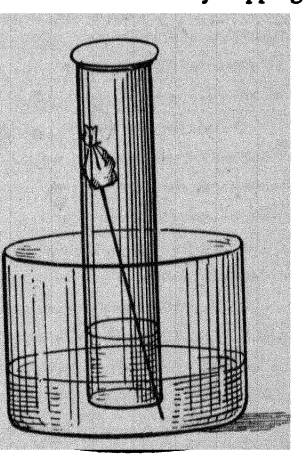


FIG. 75.

one end of a piece of bent glass or indiarubber tubing underneath the edge of the glass into the air-space inside it, when, if the tubing is clear of water, the levels will immediately become equal.

Or a great many difficulties can be avoided, especially that arising from an ill-fitting cork, by inverting an ordinary gas-jar into the dish of water. The little bag of filings can be supported in the air-space inside the jar by attaching it to the end of a piece of stout wire (Fig. 75).

Whichever way it is done the result is the same. If the apparatus be put away in a warm place the water will gradually rise up inside the jar. Day by day it will rise higher and higher, until at last it rises no further. Then more water must

be poured into the dish until the levels are again equal, and then it will be found that the air inside has become smaller in volume by about one-fifth, and that this space has been occupied by water. Indeed, it is the rising of the water inside the jar which shows us that some of the air has disappeared. Where has this portion of the air gone? The iron we know has rusted and increased in weight. It has, therefore, attached to itself some other matter, and in this case we are quite justified in concluding that the portion of the air which is lost has been taken possession of by the iron to form rust with it. If that is the case, then iron rust is the metal plus a portion of the air which the metal has seized.

Why does the iron not absorb more than one-fifth of the volume of the air in which it is confined? No more is ever absorbed, for the water ceases to rise when it has filled one-fifth of what is air-space. Can it be that it has ceased rusting whilst some unrusted iron still remains? Or has all the iron turned to rust?

It is very easy to examine the filings in the muslin bag. They are covered with rust, but the rust is only a thin covering to the particles of metal. For if the rusted filings are rubbed, their outer covering of rust is rubbed off, and unrusted metal is seen beneath. So that the iron is not all rusted away, and the further disappearance of air does not cease on this account. What is it due to?

To get an answer to this question further experiments must be made. In an inverted gas-jar, over water, in the manner already described, we suspend, by means of a stout copper wire, a muslin bag of clean iron filings, and allow them to rust and absorb one-fifth of the air in the jar. However long we may leave the apparatus alone after this has taken place, there will be no more absorption of air.

After equalizing the levels of the water inside and outside the jar, the bag of rusted filings is withdrawn without lifting the mouth of the jar out of the water, and therefore without allowing any fresh air to enter. This is easily done, because the copper wire which holds up the bag of filings is so readily bent in withdrawing it. Another bag of fresh clean

filings is just as easily introduced in the place of that taken away.

The jar is then again put aside into a warm place, but although the filings now in the air-space are fresh and unrusted, there is no more air absorbed; and when, after a while, the bag of filings is withdrawn, they are found not to have rusted.

Certainly, then, air in which iron has rusted as much as it will is not able to bring about rusting in fresh iron. We now see that when our filings have rusted and absorbed one-fifth of the air, they rust no more, and there is no further absorption of air.

The air which is left, constituting four-fifths of the original volume of air, must, therefore, be in some way a different kind of air from ordinary air. This can be easily shown in many ways.

Put a glass plate over the mouth of the jar whilst still under water, lift out the jar, and turn it right end up. Plunge a lighted taper into the *vitiated air*, as we may call it, and as the old chemists called it. The taper is immediately extinguished (Fig. 75). This is not behaving like ordinary air. A favourite experiment used to be to put a live mouse or bird into the vitiated air, and show that in a very little time the animal died. We need not do this, but it is important that we should know that animal life cannot exist in it.

If the iron in rusting merely attaches to itself some of the air, leaving four-fifths of it unchanged, and that portion which is left will not support light nor life nor rust iron, we are bound to conclude that iron absorbs that portion of our air which does support life and light. We must also conclude that air consists of two kinds of gas. One-fifth of it, which we may call *active air*, can support life and light and rust iron, whilst the other four-fifths is quite a different kind of gas, which can do none of these things.

Nothing has been said as to the part played by water when iron rusts. We know the metal will not rust in quite dry air, and that water is necessary for the rusting of iron. The reason why water is necessary, however, must be left until your studies are more advanced.

It has been shown that other metals rust in air, with or

without heat, and that they do not rust in the absence of air. It is not difficult to show that each one, in rusting, absorbs one-fifth of the air, and that the four-fifths which is always left is vitiated air.

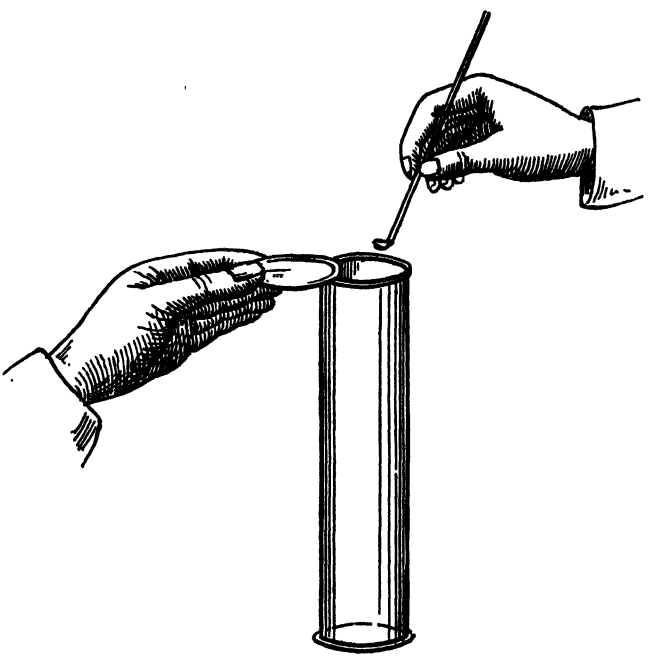


FIG. 76.

In a later lesson you will be shown how active air can be obtained again from the rusts of some metals by using proper means.

Nowadays we call active air *oxygen*, and vitiated air *nitrogen*.

SUMMARY.

When *iron rusts or burns*, there is an *increase in weight.*

If *iron rusts* in a *confined volume of air*, *one-fifth* of the air *disappears.* The remainder is *vitiated*, for—

Iron will not rust in the air which is left.

A *taper will not burn in vitiated air*, nor can an animal live in it.

Air consists of two kinds—*active air* and *vitiated air.*

LESSON XVII.

COMBUSTION OF PHOSPHORUS.

Apparatus needed:—Yellow and red phosphorus, barometer-tube, stout copper wire, small dish; bell-jar or pickle-jar, small dish, sand; flask, cork, tube, and clip; balance and weights; test-tube and cotton-wool; aspirator.

You can get phosphorus in two forms. One of them is a light yellow, half-transparent, waxy-looking, solid substance. This is the ordinary form. You buy it in sticks which are always kept completely immersed in water. *Never touch phosphorus with your fingers.* When it is necessary to take some from the bottle, use a small pair of tongs—sugar-tongs will do. Place the portion you withdraw from the bottle in a saucerful of water, and when you require a small piece, cut it off with a knife under the water.

You must take all these precautions because this yellow phosphorus takes fire so readily at quite a low temperature. Cut off a very small piece, place it on blotting-paper, and fold the paper gently over it for a moment to dry it by absorbing the water. Then lay it on a dry spot, and if the room is warm, the chances are that in a few moments it will burst into flame. You cannot, therefore, be too careful in handling it. Fragments have been known to get under the finger-nails of careless persons and take fire there, causing very painful burns.

The other form of phosphorus is a dark red powder. Placing this beside the yellow waxy kind, one might say that it is impossible for these two substances to be the same kind of matter. They are, however, and for very many purposes the red variety is the better to use, because it is the safer. It does not take fire nearly so readily as the other.

Even when the waxy variety does not burst into flame, it is slowly consumed, as iron is, in air. A stick of phosphorus placed in a bottle which is well stoppered glows for a long time, and the glow is easily seen in the dark. Very many years ago a bottle with a piece of phosphorus in it was much used as a source of feeble light—a kind of night-light, in fact.

Take a barometer-tube—that is, a long tube closed at one end—and stand it with its open end in a vessel of water. Cut off a piece of phosphorus about the size of a very small pea, and place it in a small vessel of lukewarm water. It melts in the water, or if it does not, the water is not quite warm enough. When melted, take a stiff wire, more than half as long as the barometer-tube, place one end of it in the melted phosphorus, and keep it there until the water cools and the phosphorus solidifies again. It is then firmly attached to the wire. Then place the wire with the phosphorus on the top of it into the barometer-tube, and replace the open end of this in the vessel of water.

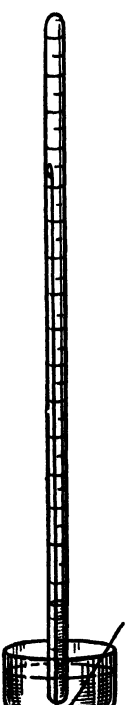


FIG. 77.

Now you have a piece of phosphorus in a confined volume of air. Place the apparatus aside, and examine it from day to day. You see a very slight smoke coming from the phosphorus, and day by day the water gradually mounts up the tube until it fills one-fifth of it. Then it mounts no more, although there may be plenty of phosphorus left on the wire. It is clear that one-fifth of the air has disappeared.

This looks as if the same thing had taken place as occurs when iron is rusted in the air contained in a vessel over water. One-fifth of the air disappears in both cases. We must examine the remainder.

Straighten the wire, place your thumb on the open end of the tube under water, withdraw the tube, and invert it. The wire will slide down to the closed end, and the remainder of

the phosphorus will be immersed in the water contained there. Have a lighted taper ready, take away your thumb, and plunge the light in the air of the tube. The light is extinguished, and so far it appears as if this air is similar to that which is vitiated by iron rusting in it, especially as it forms four-fifths of the atmospheric air.

In this experiment the phosphorus which is unconsumed shows no appearance of any kind of rust. If what went on in the long tube was anything like the rusting of iron, the phosphorus absorbed the active air, or oxygen, and therefore some sort of rust should have been formed. To get further information about this the following experiments should be tried:—

Use a bell-jar as in the figure, resting in a dish of water similar to the one which was used for the rusting of iron over water. A piece of wire may be easily twisted to form a little stand which shall project about one-third up from the surface of the water, and on this stand place a small dish with a little sand at the bottom of it, and a small piece of dry phosphorus on the sand. Or the dish with the phosphorus may be floated on the surface of the water. The level of the water inside the jar and outside will be the same. Make a wire hot at one end, touch the phosphorus through the opening at the top of the jar, immediately replace the cork or stopper, and press down the jar. The phosphorus burns with a very brilliant white light, whilst dense clouds of a white substance fill the jar. After the phosphorus has ceased to burn, the white

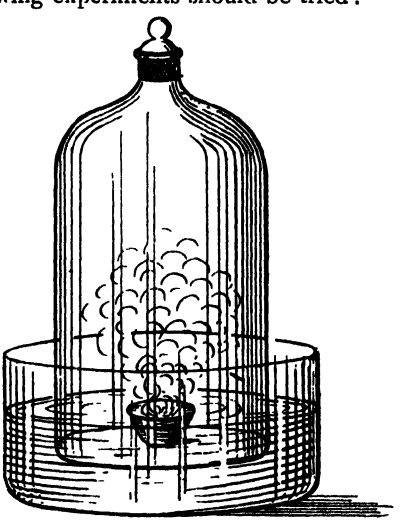


FIG. 78.

clouds gradually disappear, and when the jar is clear you can see that the water has risen some distance inside it. When the jar is cold, and you have poured more water into the dish, so as to make the levels equal again, it will be evident that again about one-fifth of the air in the jar has disappeared.

Take out the stopper, and test the air left in the jar with a lighted taper. You will find that it is vitiated air, for the taper will not burn in it. If you tried to burn more phosphorus in it you would not succeed. Dip your finger in the water inside the jar, and transfer a drop of it to your tongue, and you will find that it tastes sour or acid. Dip a piece of blue litmus-paper into it, and the paper becomes red. Blue litmus-paper always turns red in acid liquids.

What has become of the white fumes which were formed when the phosphorus burnt? Is it possible that they have dissolved in the water of the jar, thus souring it? It is difficult to see what has become of them if they are not dissolved in the water. We will examine these fumes more closely by producing them where there is no water.

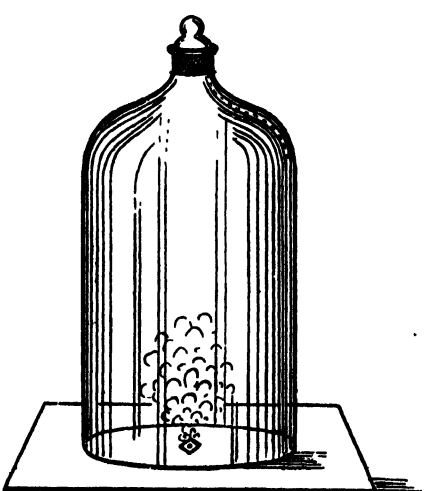


FIG. 79.

Take the bell-jar again, dry it, and cork it well. Place a small piece of phosphorus in a small dish, and stand this beneath the bell-jar on a slate or piece of glass. Lift off the jar, ignite the phosphorus with a touch of a hot wire, and immediately it begins to blaze replace the bell-jar over it. White fumes are plentifully produced; and after the phosphorus

has become extinguished, the clouds slowly subside, and the slate or glass is found to be covered over with a beautiful white powder like fine snow. This is clearly the *rust of phosphorus.*

As you watch it the white rust becomes wetter and wetter and slowly disappears, for it rapidly takes moisture from the air and dissolves in it. If you keep it perfectly dry it does not change at all. Scrape off a little of it, place it in water, and taste a drop. It is sour. Test it with blue litmus-paper, and it proves to be acid. You now see why the water in the previous experiment was sour.

If the white powder is the rust of phosphorus formed by the active air, or oxygen, uniting with the phosphorus, it should weigh more than the phosphorus out of which it was formed. We can easily show that it does.

You take a small dry flask, and fit it with an indiarubber stopper, or a good cork which has been rubbed over with wax. The cork or stopper has a hole through it, and a glass tube passes through the hole. Over the end of the tube is slipped an inch or two of rubber tubing closed by a clip. The whole must be air-tight.

Into the flask drop a small piece of dry phosphorus. Replace the stopper firmly, and counterpoise it on the balance. Warm the flask very gently over the flame, on every or any part except where the phosphorus is, until the phosphorus takes fire. It may be that the burning phosphorus breaks the flask; but if not, when the flask is replaced on the balance, you will find that its weight is unaltered.

You may say that this does not show that the rust of phosphorus weighs more than the phosphorus did.. But it does prove it. You know that the burning phosphorus has absorbed one-fifth of the air. Therefore the remaining four-fifths of the air now fills the jar, and its density and pressure is, therefore, only four-fifths of ordinary air. Open the clip, and air immediately rushes in until the density and pressure inside the flask is the same as at first. Now the flask is no longer counterpoised, but it is heavier than before.

You may show this in another manner. Take a test-tube,

and fit it with a cork through which passes a piece of glass tubing. Make a small hole through the closed end of the test-tube by a sharp blow with the needle-end of a pair of compasses, or something similar. You can do this more easily if you make the end of the tube red-hot first in a Bunsen flame. Hold the tube horizontal, place a loose plug of cotton-wool

FIG. 80.

over the hole, then a small piece of phosphorus about halfway down the tube, and finally another loose plug of cotton-wool, which must not be too near the phosphorus. After this replace the cork. Then, after counterpoising the test-tube, attach it to an *aspirator*, or apparatus for drawing a slow current of air through the test-tube. Fig. 80 will show the apparatus, and the aspirator may be similar to that figured in Lesson II.

Whilst a current of air is passing through, the test-tube is gently heated with a Bunsen or spirit flame until the phosphorus ignites. Clouds of phosphorus rust are produced, but they are caught by the cotton-wool, and cannot therefore escape from the tube. After the experiment is over the tube is weighed again, and found to weigh more than before. Phosphorus, therefore, combines with the active portion, or oxygen, of the air, forming a white powder with it. This oxygen constitutes about one-fifth of the air by volume, the remaining four-fifths being vitiated air, or nitrogen.

SUMMARY.

Phosphorus slowly *combines* with *one-fifth of the air* in a confined space. The *remainder* is *vitiated air.*

The *rust of phosphorus*, which is a white powder, *dissolves* in water, and gives it a *sour or acid taste.*

The *white rust of phosphorus weighs more* than the phosphorus from which it was formed.

LESSON XVIII.

THE COMBUSTION OF A CANDLE.

Apparatus needed:—Candle, test-tube with cork and small glass tube through it; Bunsen burner or spirit-lamp; lamp-glass, U-tube, aspirator, connecting tubes, and caustic soda; beaker or tumbler; bell-jar, dish, twisted wire; flask, delivery-tube, gas-jars, washing-soda, hydrochloric acid, lime-water; piece of charcoal; deflagrating spoon.

Look at the candle before you. It is made of a white fat or grease, hard or soft according to the kind of fat which is used, whilst through it runs a piece of twisted cotton cord, which we call the wick. The cheapest candles are made by dipping the wicks in melted tallow; when they are withdrawn some tallow adheres to the wick. They are dipped over and over again, until there is enough fat about the wicks to make them the required thickness. The better kinds of candles are made by pouring the melted fat into moulds, the middle of each mould being occupied by the wick.

A piece of dry cotton cord burns very poorly, and gives out very little light, whilst a piece of candle-fat thrown upon the fire blazes up quite fiercely. There is, therefore, little doubt that it is the fat of the candle which burns so brightly and gives out so good a light.

Fig. 81.

If you watch a candle burning you will see that at the top the fat is hollowed out into a little cup, which is full of melted fat. You will see quite a current of melted fat moving towards the wick, and if you sprinkle a little dust over this melted fat, you will observe it mounting up the wick towards the flame. The heat of the burning candle melts the

fat near the flame, and the liquid fat rises up the wick by capillary attraction, in the manner you have already learnt.

Whilst the candle is hot and burning well, blow it out. You will now see that a smoky vapour rises from the wick whilst it is still very hot. Whilst this smoke is ascending, approach a lighted taper or match to the wick. Before the light reaches the wick, the vapours will catch fire, and the flame will strike down to the wick, thus relighting the candle (Fig. 82). It appears from this that the heat of the burning candle vaporizes the fat, and that it is really a gas which is burning when the candle is alight.

Fit up a test-tube with a cork, through which passes a short piece of straight glass tubing. Place a little candle fat in the test-tube, and heat it strongly over the flame of a Bunsen burner or spirit-lamp until the fat melts and vapours are seen leaving the tube. Then

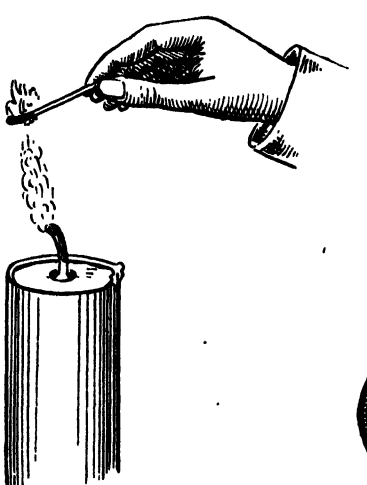


FIG. 82

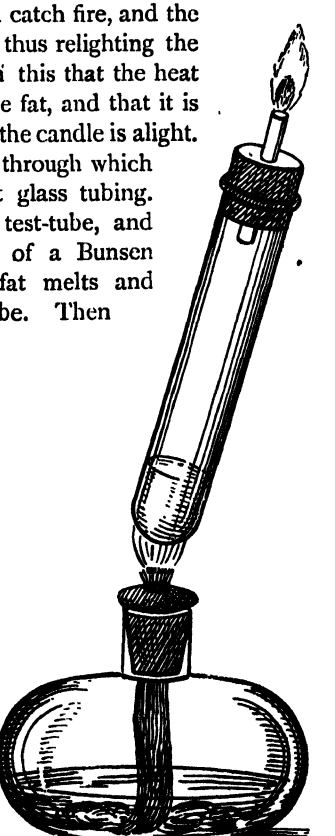


FIG 83.

apply a light to the vapours, and they will burn as they issue forth from the test-tube, just as a candle burns.

It is, then, the fat which by burning provides the light in a burning candle. The wick is useful only to draw up the melted fat to the flame, where it is vaporized by the heat, and then burns. It is really, therefore, a gas or gases which burn in the candle.

You already know that, although there is no visible rust when the candle burns, as there is when iron, phosphorus, and other substances attach to themselves the active portion of the air, yet there are invisible products which, when you catch and weigh them, are found to weigh more than the candle which is burnt. It will be as well to do this experiment over again.

Fig. 84 shows how the apparatus is arranged. The candle

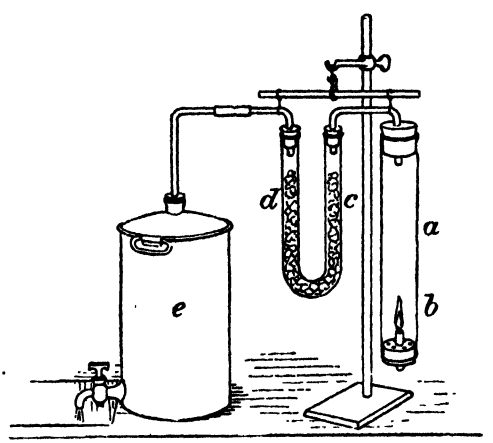


FIG. 84.—(From Jago's *Inorganic Chemistry.*)

b burns in the lamp-glass *a*. The products of its combustion are sucked away in a current of air through the U-tube *cd*, which contains lumps of caustic soda. The current of air is produced by running water out of the large vessel *e*. If the lamp-glass, U-tube, and their contents are weighed before and after the candle is burnt, there is found to be an increase in weight, although some of the candle has disappeared. This

agrees in that respect with the rusting of iron and the combustion of phosphorus.

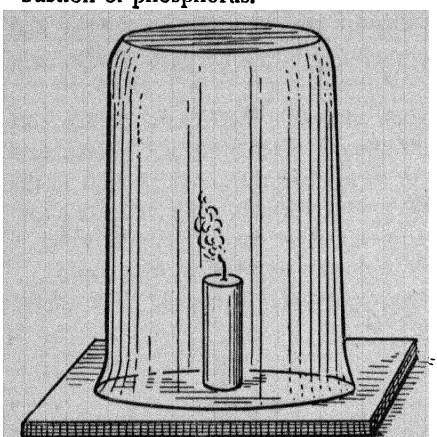


FIG. 85.

The candle will not burn at all without a proper supply of air. Take a short length of candle, light it, and place it on the table. Cover it over with an inverted beaker or tumbler. It continues burning for a few moments, then flickers, and quickly goes out. Can it be possible that it uses up the active part of the air, and is then extinguished because there is no further supply?

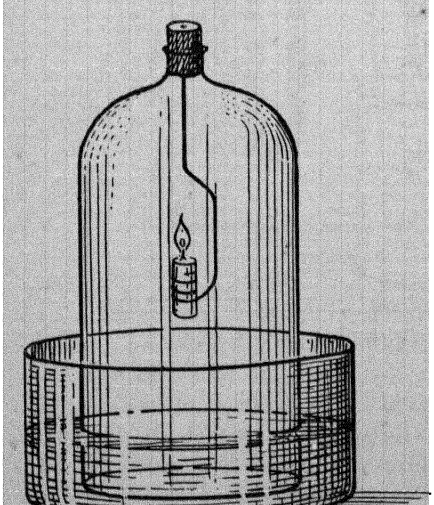


FIG. 86.

Take the bell-jar in which phosphorus was burnt over water, and prepare the dish in which it is to stand by placing water in it. Through the cork, which tightly fits the neck of the jar, pass the end of a piece of wire, the other end being twisted so as to make a holder for about an inch of candle. Place the jar in the dish, light the candle, and put it in position with the cork well pressed down.

The candle burns for a time, which depends upon the amount of air in the jar, then goes out. Allow the jar to cool down, make the levels of the water inside and outside the jar equal, and you will then see that about one-fifth of the air has disappeared. Test the remaining air with a lighted taper. The taper is extinguished. Phosphorus will not burn in it, nor will iron rust in it. It is very likely, then, to be vitiated air —that is, nitrogen.

But what is produced when the burning candle takes possession of the active air? Whatever is produced must be gaseous, for it is quite invisible.

In Lesson X. an experiment is described which it is well to do over again. You take a flask and fit it with a delivery-tube and acid-funnel, as in Fig. 87. Into the flask is placed some common washing-soda, and hydrochloric acid is poured down the funnel. A very heavy gas is given off; this has no smell, will not burn, and will not allow a taper to burn in it. It has, however, one property which distinguishes it from other gases.

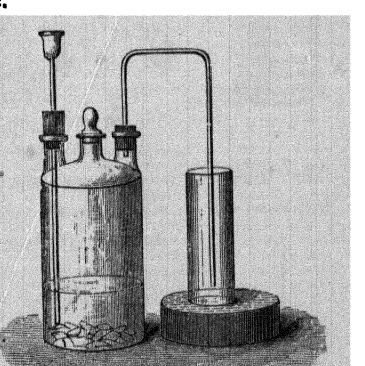


FIG. 87.—From Miller's *Chemistry*.)

You must place in a bottle a little quick-lime, and fill up the bottle with distilled water or clean rain-water. Shake it up well, put it aside, and in a few days there will be a layer of lime at the bottom of the bottle, whilst the water above it will be quite clear. The water has, however, dissolved a little of the lime. We call it *lime-water*.

Now collect a jar of the gas from washing-soda and hydrochloric acid, pour into the jar a little clear lime-water, and shake it up with the gas. The lime-water becomes very milky, and if it is allowed to stand, a fine white powder will settle at the bottom of the vessel. The powder is *chalk*. We may call the gas *chalk-gas*.

Now take another gas-jar, pour some clear lime-water into it, and shake it up. There is no milkiness. Take a piece of charcoal about the size of a nut in a pair of tongs, or twist the end of a wire around it. Make this red-hot in the flame, and then hold it in the air of the gas-jar for about a minute. Withdraw the charcoal, and again shake up the lime-water in the jar. This time it becomes milky, showing that when charcoal burns in air it produces chalk-gas.

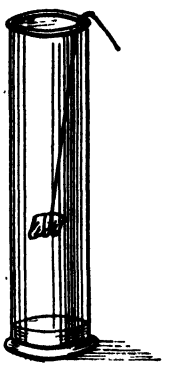


FIG 88.

Twist a piece of wire into such a shape that it will hold a piece of candle at one end. Light the candle, and lower it into a gas-jar of air. Keep it there for a minute, if it will burn so long. Then withdraw the candle, pour in clear lime-water, and shake it up. The lime-water becomes milky, showing that when a candle burns in air chalk-gas is produced, just as it is when charcoal burns in the air.

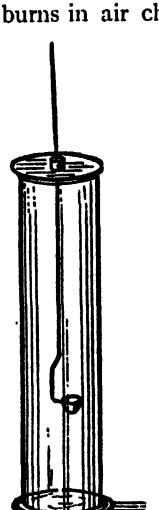


FIG. 89.

Light a candle, and hold in its flame a piece of cold porcelain or glass. In a few minutes this is covered with a thick coating of black soot. The soot must clearly have come from the candle. Scrape off the soot, place it in a *deflagrating-spoon*, make it red-hot in the flame, and lower it into a jar of air. After a minute draw it out, pour in lime-water, and shake up. The lime-water becomes milky. Therefore this black soot is the substance which, existing in the candle, produces chalk-gas when it burns. You remember that charcoal does likewise. Therefore candle soot and charcoal are similar substances.

Light a candle, and hold over it an inverted gas-jar or beaker, which must be perfectly dry and cold (Fig. 90). In a few moments the inside of the vessel is covered over with a coating of moisture, just as if it had been breathed into. This moisture is water

which is condensed by the cold vessel, and it is formed by the burning of the candle.

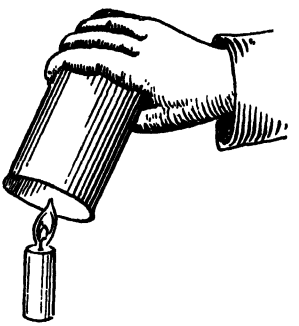


FIG. 90.

It appears, then, that the substance of the candle when burning attaches to itself the active portion of the air—that is to say, the oxygen—and forms with it chalk-gas and water.

In our first experiment the chalk-gas and water produced by the candle burning in the lamp-glass were caught and held by the lumps of caustic soda in the U-tube, and were shown to be heavier than that portion of the candle which was burnt. A more elaborate experiment would show that the mass of the candle which disappears in burning, together with the oxygen it combines with, are exactly equal to the water and chalk-gas which are formed and retained in the U-tube.

SUMMARY.

It is the *vapour of candle-fat which burns* in a candle.

The *products of combustion* of a candle are *heavier than* the *candle* itself.

A *candle will not burn without* a supply of *air*.

A *burning candle consumes one-fifth* of a confined volume of *air*. The *remainder* is *vitiated air*.

When a candle burns, *one product of combustion* is a *gas* which turns *clear lime-water milky*.

Charcoal burnt in air produces this *same gas*.

A *burning candle* also *produces water*.

LESSON XIX.

COMBUSTION OF OIL—OIL-LAMPS.

Apparatus needed:—Test-tube, cork with small tube through it, Bunsen burner or spirit-lamp, various oils; test-tube lamp; various oil-lamps.

FATS which are liquid at the ordinary temperature of the air are called oils. There are many kinds of oils and of different origin. Thus some oils are obtained from animal bodies, together with solid fats; many others are obtained from plants, principally from their seeds and fruits; whilst still others are obtained from the earth, or by the dry distillation of solid bodies obtained from the earth.

The oils obtained from animal bodies are known as *animal oils*. Whale oil and neat's-foot oil, got by boiling the feet of cattle in water and skimming off the oil which rises to the surface, are common examples of animal oils.

Linseed oil from the seed of the flax plant, cotton-seed oil from the seed of the cotton plant, olive oil from the fruit of the olive tree, palm oil from the fruit of some palms, and many others, are *vegetable oils* in common use.

Petroleum is a mixture of *mineral oils* found in many parts of the world springing out of the ground, or obtained by sinking wells; whilst paraffin is also a mixture of mineral oils, which is got by strongly heating some kinds of coal and shale in closed vessels into which no air can enter, and condensing the vapours which are driven off by the heat. The mixture of mineral oils which we call petroleum is very similar to that mixture which is called paraffin oil.

All these oils—animal, vegetable, and mineral—are alike in one particular. Place a small quantity in a test-tube provided with a cork, through which passes a short glass tube. Heat the oil, and vapours will be driven off which readily ignite at the mouth of the tube when a lighted taper is held there. In this respect they behave just as candle-fat does.

They differ among themselves, however, in another and very important particular. Some give off inflammable vapours at a comparatively low temperature, and therefore require but little heating. The mineral oils are of this class. Others do not give off inflammable vapours until more strongly heated, and the animal and vegetable oils are of this class. The mineral oils are always mixtures of oils, some of which are evaporated at much lower temperatures than others. Indeed, some of their constituents vaporize at so low a temperature that it is very dangerous to handle them near a fire or light. They are very dangerous mixtures indeed unless very carefully handled.

Pour a very few drops of petroleum into a test-tube. Cover the mouth of the tube lightly with the thumb, and warm the bottom of the tube to vaporize the oil. Then hold a lighted match or taper to the mouth of the tube, and you will get quite a little explosion. This should show you how dangerous a large amount of the vapour would be.

Since they all give off inflammable vapours when heated, it is easy to make use of them for giving light. We cannot make them into candles as we do the solid fats, but we can place them in vessels, provide a wick to pass through an orifice in the vessel, and rely upon the capillary action of the wick to draw up the oil to the orifice, where it can be heated and vaporized by the flame. We call all such vessels lamps.

A simple lamp can be made of a test-tube. Fit it with a cork, along the side of which you have cut a little nick so that air may pass into or out of the tube. Bore the cork through the middle, and pass through the hole a short straight metal tube, and through the tube a wick composed of some loose soft cotton threads twisted together so as to fairly but not tightly fill the tube. Place some colza or olive oil in the tube, and allow the wick to dip in the oil when the cork is placed in

position. You had better not use mineral oil in such a lamp as this: it would be too dangerous, especially if you dropped the test-tube and broke it. When the oil has mounted to the top of the wick, apply a light.

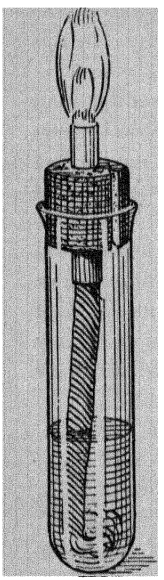


Fig. 91.

This little lamp will burn steadily so long as the wick does not protrude too much and the flame is small. If the wick is allowed to protrude too much, more oil is vaporized than can be supplied with active air, and in consequence a portion of the vapours pass away without burning. This causes the flame to smoke.

With it you can experiment just as you did with the candle, and discover that when it burns, whatever oil you use, it combines with the active part of the air, leaving the vitiated portion unaffected. Also, that it produces chalk-gas and water, and that these are heavier than the oil which is burnt. It will not burn without a constant supply of air, and a cool, bright body held in its flame will be covered over with a layer of black soot, which you can easily discover is a kind of charcoal, because it produces chalk-gas when it burns. In short, you will soon find out that the burning of oil in a lamp and the burning of a candle are quite the same kind of thing, requiring the same conditions, and leading to the same results.

The simplest kind of lamp in common use consists of three portions. There is first a vessel, which should be of metal, to hold the oil. Secondly, the burner, through which is a metal tube to carry the wick. This must have a little hole through it leading from the outside air to the oil vessel. The burner is removable, and usually the vessel is filled by removing it, although sometimes there is a separate opening, which may be closed by a screw-plug. Thirdly, there is the screw or wheel for raising or lowering the wick. The first lamps for burning mineral oils were of this kind (Fig. 92), but they are not good lamps. The flame is naked, so that every breath of air

drives it to one side or another, thus causing an unsteady light and a smoky flame. Moreover, large flames and much light cannot be obtained with these lamps. They are nothing but liquid candles.

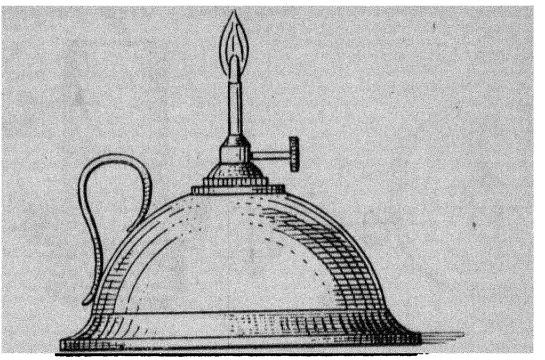


FIG. 92.

The better kinds of lamps are much more complicated. In the first place, they are provided with glass chimneys to protect the flame and increase the draught so as to get a larger air-supply to the flame. In this case the air is allowed to enter

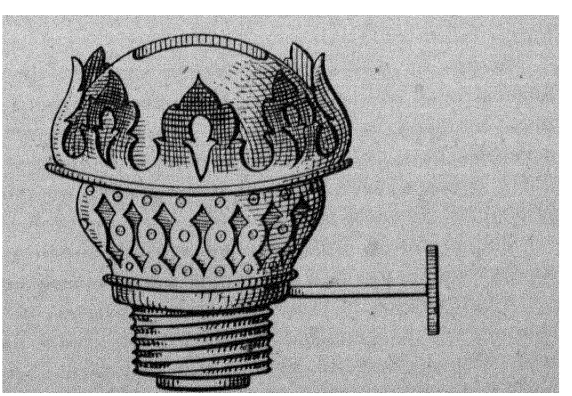


FIG. 93 —Paraffin-lamp burner.

through a number of holes in the burner, below the level of the flame. The wick is larger than in the naked flame lamps, and

it therefore burns more oil and gives out more light, and it is either flat or bent round in a circle. When the wick is flat a metal cone or dome with a slit in it is placed over the top of the wick, and the air-current is so directed that it fully feeds the flame and flattens it out, thus giving a larger and brighter flame

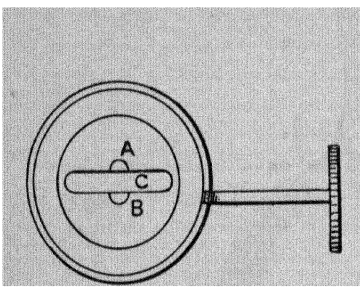


FIG. 94.—Base of burner. A, B, Air-holes; C, Wick-tube.

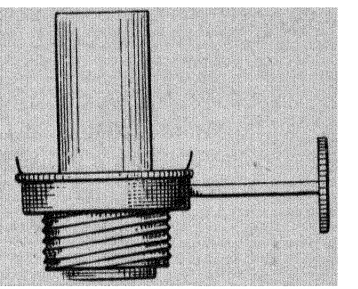


FIG. 95.—Lamp burner with dome removed.

than could otherwise be obtained. When the wick is bent in a circle, air is admitted to the interior of the flame as well as to the exterior, so that the flame is hollow. The best lamps have an extinguisher, which is an arrangement for shutting off the supply of air from the flame, when, of course, the oil-vapours can no longer burn.

Lamps of this kind are chiefly used for the burning of mineral oils, and, for the reasons already given, special care must be taken in using them. A great many fires and fatal accidents have been caused through children, and sometimes older persons, handling lamps carelessly, or through using broken or defective lamps.

The Board of Education has recently issued a number of suggestions for the care and proper use of petroleum lamps.

A lamp which is broken or in any way out of order, or a chimney which is cracked, should not be used. The chimney must fit properly, and when a new chimney is required, the burner should be taken to the shop in order that it may be fitted on.

The wick must quite fill the wick-tube, but it must not be too tight, or it cannot properly draw up the oil for burning. If

the wick is damp, it should be dried in front of the fire before it is put into the lamp. A damp wick cannot act properly, and it causes a spluttering flame. The wick must reach the bottom of the oil-chamber, and there should be a little to spare, but not too much.

Even the best mineral oils contain impurities, which after a time clog up the wick. It should, therefore, be thrown away after two or three months' use. Before lighting the lamp, remove the hard crust which always forms on the top of the wick with a dirty duster. Wicks should never be cut.

The lamp must never be filled when it is alight, nor when it is near a fire or other light. Many dangerous accidents have happened through the vapour of the oil taking fire when a lamp is being filled. Therefore it is better never to handle the oil nor fill a lamp after dark.

After the lamp is filled, see that the burner is properly screwed on, or, if there is a side filling hole, be careful to screw in the plug. If any oil is spilled on the lamp whilst filling, be careful to wipe it all off. In any case it would cause a bad smell in the room, and it may take fire when the lamp is lighted.

Also, on account of the offensive smell otherwise caused, and for other reasons, the burner should be thoroughly cleaned at frequent intervals. Burnt pieces of wick, dead flies, and dirt of all kinds collect in and upon the burner, and become saturated with oil. In addition to the bad smell this causes, there is the danger that rubbish of this kind may stop up the ventilating-holes of the oil-chamber and the air-holes of the burner.

Before lighting the lamp, you should see that the slit in the cone of the burner is exactly over the wick-tube, otherwise, when the lamp is lighted, the flame will not be of a good shape and will come in contact with the metal, making the burner too hot. There is danger in this, because the oil in the chamber will be heated and vapour will be given off too readily; and if the oil is low, and there is consequently a large space in the vessel full of oil-vapour and air, this mixture may explode, and burst the lamp.

For this reason also the lamp should never be allowed to

burn after much of the oil is consumed. It should always be kept well filled. Nor should the lamp ever be allowed to remain burning with the wick turned low, for the small flame thus produced underneath the cone heats the burner greatly. When first lighted the wick should be low. It should then be gradually raised to produce a larger flame, but must not, of course, be allowed to smoke, which it will do if turned up too high.

If the lamp has an extinguisher, it is easily and safely put out. If it has not, the wick should be turned down until there is a very small flame. Then a piece of cardboard may be placed on the top of the chimney, or you may blow across the top of the chimney, to extinguish the flame. You should never blow down the chimney. In turning down the lamp be careful not to lower the wick so much that it falls into the oil-chamber. That would probably cause a serious accident.

The oil-chamber should not be of glass or any other easily broken substance. If such a lamp should fall and the oil-chamber break, the whole place, and probably the persons near at hand, would be set on fire. Even lamps with metal oil-chambers may cause serious accidents by falling when lighted. If the lamp is a hanging lamp, the nail or hook which supports it should be very firm; if it is a table-lamp, it should stand on a firm, level surface. Lamps should be carried about as little as possible when lighted. If it is necessary to move a lighted lamp, it should be done with the greatest care.

Some people make it a practice to pour mineral oil on a fire in order to light it easily, or to make it burn better. This is an exceedingly dangerous practice, which ought never to be done.

In case a person's clothes should be set on fire, the flames should be extinguished by wrapping a hearthrug, blanket, great coat, or anything of that kind closely around them. If you shut out the air from burning clothes, they can no longer burn. Send immediately for the doctor, and in the mean time cut off the clothing and anoint the burns freely with sweet oil, linseed oil, or, better still, carron oil, which can be got at the chemist's.

SUMMARY.

Oils may be of *animal*, *vegetable*, or *mineral* origin.

Heated oils give rise to *vapours* which *burn*.

Mineral oils are generally *vaporized* at a *low temperature*.

When *oils burn* their *products of combustion* are *chalk-gas and water*. They *use up one-fifth of the air*. The *remainder* is *vitiated*.

The *greatest possible care* must be taken in *using oil-lamps*, *especially* those which burn *mineral oils*.

If a *person's clothes take fire*, immediately *shut off the air from them* by *wrapping something* around them.

LESSON XX.

COMBUSTION OF COAL-GAS.

Apparatus needed:—Long clay pipe, small coal, piece of damp clay, Bunsen burner; beaker, bent glass tube, rubber tube, and gas-jar; lime-water; lamp-glass, corks, and tubes; test-tube and dish; wire gauze.

You were told in the last lesson that when coal is distilled dry—that is to say, when it is heated in closed vessels into which no air is allowed to enter—paraffin is driven off. A great many other things besides this are expelled from coal by heat, but amongst them are many substances which are gases at the ordinary temperature of the air. This mixture of gases, when purified by proper means, constitute what is known as coal-gas.

You may easily make coal-gas on a small scale. A long

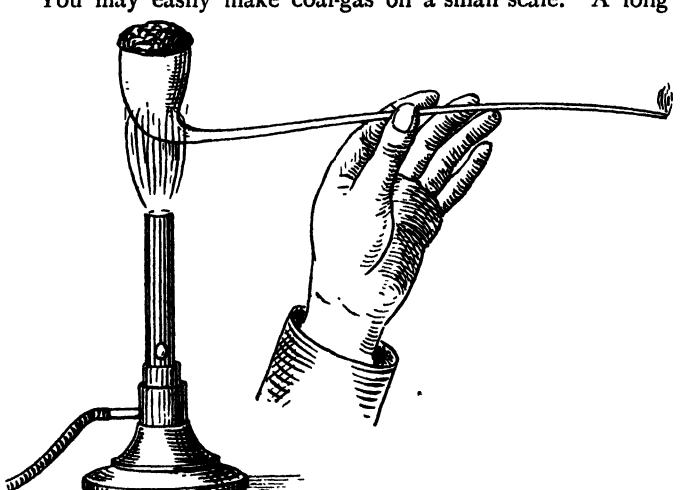


FIG. 96.

clay pipe has its bowl filled nearly full of small pieces of coal —not dust—and then a plug of damp clay is placed over the coal so as to close up the bowl. When the bowl is placed in the fire between the bars of the grate, or when it is strongly heated in the Bunsen flame, vapours issue from the hole at the end of the stem, and these will burn very well if a light be applied to them. This illustrates the dry distillation of coal on a small scale.

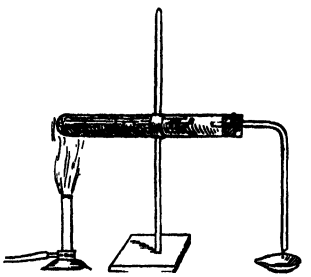


FIG. 97.

Or you may take a hard glass tube closed at one end—a test-tube will do if no other tube is at hand—and fit the open end with a bored cork. A glass tube passes through this, and the gas may be burnt at the end of the tube. Besides the inflammable gas, tar is driven off, and drops from the end of the tube, and you get an unmistakable smell of ammonia.

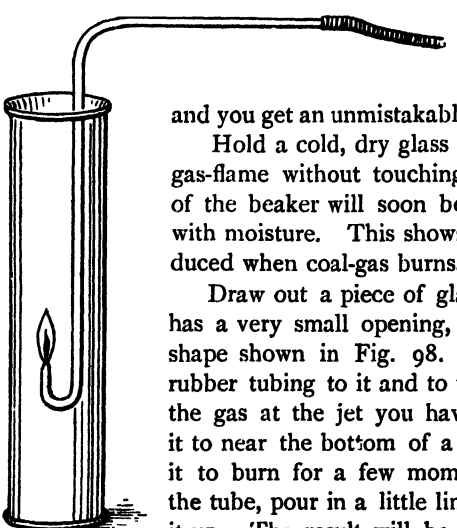


FIG. 98.

Hold a cold, dry glass beaker over a small gas-flame without touching it, and the inside of the beaker will soon become covered over with moisture. This shows that water is produced when coal-gas burns.

Draw out a piece of glass tubing so that it has a very small opening, and bend it to the shape shown in Fig. 98. Attach a piece of rubber tubing to it and to the gas-supply, light the gas at the jet you have made, and lower it to near the bottom of a jar of air. Allow it to burn for a few moments, then withdraw the tube, pour in a little lime-water, and shake it up. The result will be that the lime-water is rendered milky, showing that chalk-gas is produced when coal-gas burns.

It is not difficult to arrange an experiment to show that the gas-flame will use up the active portion of the air (see Fig. 99), but it is more difficult to show—although it can be shown with proper apparatus—that the chalk-gas and water which are produced when coal gas burns in air weigh more than the gas which is consumed. You therefore see that the combustion of

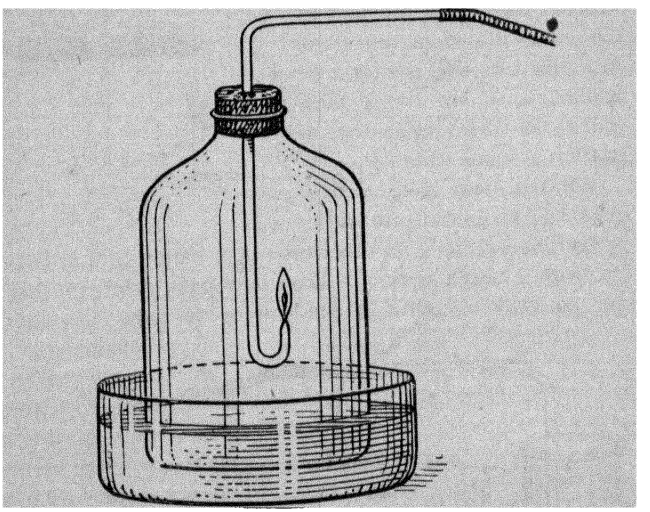


FIG. 99.

coal-gas in air is of a similar kind to the combustion of oils and candles in air. But with coal-gas we can find out another thing about combustion which is not so easy to discover in the combustion of oil and candles.

You take an ordinary lamp-glass, and fit each end of it with a cork. Through the upper one bore a hole, and pass through it a short piece of glass tubing more than half an inch in diameter. Through the lower one bore a central hole for a piece of rather wide glass tubing bent at right angles, and another hole through which is passed another piece of narrower tubing bent also at right angles. Attach a piece of rubber tubing to the smaller tube, and connect it with the gas-supply.

Turn on the gas, and wait a few moments, until all the air

in the lamp glass is expelled; then, placing a finger on the end of the upper tube, apply a lighted taper to the open end of the lower central tube. The gas will ignite there, and you will have a flame of coal-gas burning in air. Now swiftly take away your finger from the upper tube, and the flame will recede up the tube until it appears as a little faintly luminous flame of air burning in an atmosphere of coal-gas. The active portion of the air will there be combining with the coal-gas, but there will be a considerable amount of coal-gas escaping from the upper tube uncombined, and this may be, and should be lighted, to prevent its escaping into the air.

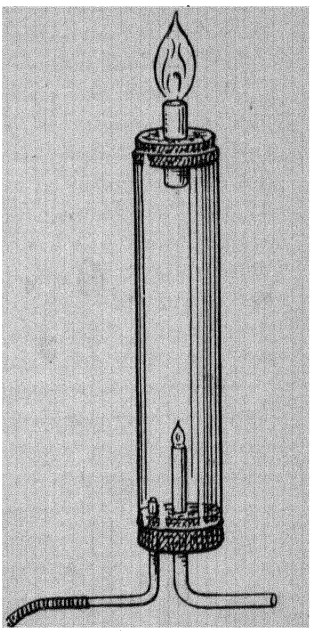


FIG. 100.

We see from this that it does not matter whether the gas and the active air unite in an atmosphere of coal-gas or in an atmosphere of air. The union of the two is all that is needed to show the phenomenon of combustion.

If a mixture of coal-gas and air is made in the right proportion—namely, one part of coal-gas to about twelve parts of air—the two will be in the best possible condition for union, and the combination will be so rapid that an explosion will result. Invert a test-tube full of water in a vessel of water, and pass coal-gas into it from a rubber tube, one end of which is slipped over the gas supply, and the other held under the open end of the test-tube. When full, withdraw the test-tube, keeping it inverted, and apply a light to its mouth. The gas burns quietly at the mouth. Whilst burning, erect the tube, and the flame runs down to the bottom of it. Now fill the test-tube again with water, invert it in the vessel of water, and then, lifting its edge, allow about $\frac{1}{13}\frac{2}{3}$ of the water to run out and air to enter.

Into this pass coal-gas to fill the tube. You have now a mixture of one part of coal-gas to about twelve parts of air. Give the gases a few moments to thoroughly mix, withdraw the tube from the water, and immediately apply a light to its mouth. Quite a little explosion results.

You must not experiment with large quantities of this mixture, because it is so dangerous. Many serious accidents and fires have resulted from coal-gas leaking into a room, until the room is full of an explosive mixture of the gas and air. Thoughtless persons have then sought for the leak with a light in their hands, with the result that an explosion has occurred. When a leak in the gas-pipe is suspected, turn off the gas at the meter, and open every door and window in the room until clear of gas. Then turn on the gas again, and look for the leak with all the doors and windows wide open.

We cannot conveniently use an ordinary gas-flame for the purpose of heating vessels, because of the coating of black soot which it so easily deposits, but it is possible to mix just so much air with gas that not enough is added to make the mixture explosive, but enough to make the gas burn with a flame which gives out but very little light, much heat, and which does not deposit soot. This is done in the gas-burner called the Bunsen burner.

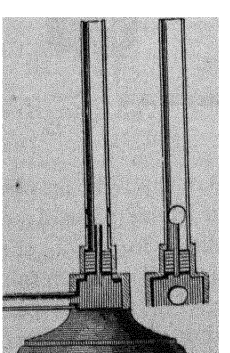


FIG. 101.—(From Miller's *Inorganic Chemistry.*)

This is made of metal. The tube may be unscrewed, and then the small hole through which gas is admitted to the tube may be seen. At the base of the tube are holes which admit air, and it is furnished with a collar which can be turned round so as to cover or uncover the air-holes. When the holes are covered so that no air can enter and mix with the gas, the flame is luminous; but when they are uncovered, air enters and mixes with the gas, and the mixture of gas and air which burns at the top gives a hot, non-luminous flame which does not deposit soot.

You can get just such another flame in this way. Procure

a piece of closely woven wire gauze, fix it about 3 inches over an ordinary gas-burner, turn on the gas, and light it above the gauze. Air mixes with the gas before it passes through the gauze, and the mixture burns with a non-luminous flame. The flame does not strike through the gauze to the gas underneath. This is because the metal gauze carries away the heat so rapidly that it does not get hot enough on its under side to ignite the gas there. An arrangement of this kind may be made to serve instead of a Bunsen burner. The vessel to be heated must be supported a short distance above the gauze.

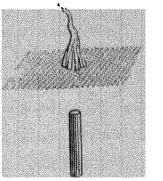


FIG. 102.—(From Newth's *Inorganic Lecture Experiments.*)

Use is made of this property of wire gauze to make a lamp for miners to use in an atmosphere containing sometimes much coal-gas, for in coal-mines gas frequently escapes from the coal and mixes with the air in the mine, forming an explosive mixture. The lamp is a small oil-lamp, which is completely enclosed in impervious metal and wire gauze. The explosive mixture of gas and air—called by the miners fire-damp—enters through the gauze, and may then be seen burning around the oil-flame. But the gauze prevents the flame getting outside, and so exploding the mixture in the mine. Of course, when gas is seen burning in the miners' lamps means are immediately taken by increased ventilation to drive the fire-damp out of the mine. When an explosion occurs many lives are lost through the explosion, and many others after the explosion is over through the suffocating chalk-gas, or after-damp, as it is often called, which is produced by the union of the coal-gas and the oxygen of the air.

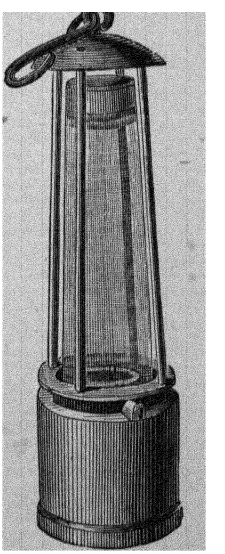


FIG. 103.—(From Tyndall's *Heat.*)

SUMMARY.

Coal-gas is *driven* out of coal by *heat.*

When *coal-gas burns* it *produces chalk-gas and water.* It *uses up one-fifth* of the *air,* and the *products of combustion* are *heavier* than the gas which is burnt.

Its combustion is a *union* of *active air* with the *gas.*

A *mixture* of *one part of coal-gas* to *twelve parts of air* is *very explosive.*

A *mixture of coal-gas and air* in the right proportions will *burn* in air with a *non-luminous flame.*

A *flame passes* with *great difficulty* through *wire gauze.*

LESSON XXI.

FLAMES.

Apparatus needed:—Candle, oil-lamp, gas-burner; piece of drawn-out glass tubing; wire gauze; thick copper wire; Bunsen burner.

THE hot gases which are produced from the wax or grease of a candle by heat, those which are produced from various oils by heat, as well as coal, and other inflammable gases, which are gases at the ordinary temperature, produce so much heat when they combine with the oxygen of the air that a flame is the result.

Look closely at a candle-flame. It is not difficult to distinguish three distinct portions of it. First of all, there is a space, 2, around the wick which is darker than any other portion of the flame. Immediately outside this dark zone, as it is called, is the luminous portion, 3, or zone of the flame. This is brightest about the middle of the flame, but not so bright towards the pointed tip. Beneath the luminous zone, and surrounding the lower half of it, is a zone of flame, 1, of a clear blue colour, but not giving out much light. This is the non-luminous outer zone.

FIG. 104.—(From Furneaux's *Elementary Chemistry.*)

The candle-flame, then, consists of three clearly distinct portions—

(*a*) The dark central zone.

(*b*) The luminous zone.

(*c*) The non-luminous outer zone.

The flame of an oil or gas-lamp shows the same three portions.

Let us first examine more closely the dark central zone. When the flame of a candle is burning very steadily and erect, thrust into the middle of this zone the head of a common lucifer-match, and hold it there. If this is done deftly it will be some seconds before the phosphorus on the match ignites, often not until the wood in the luminous zone has caught fire. Evidently, then, the middle of the flame is not so hot as the luminous zone.

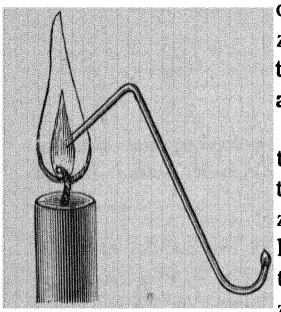


FIG. 105.—(From Furneaux's *Elementary Chemistry.*)

Draw out a piece of glass tubing so that one end of it forms a jet. Hold the other end in the dark central zone of a candle-flame, and apply a light to the jet. A flame appears there, showing that the dark central zone contains unburnt combustible gases which may be drawn off by a tube.

Take a sheet of stout white paper, and, holding it horizontally, lower it down on the candle-flame, until it reaches the top of the wick. Remove it quickly, and you will find a ring of black soot upon it, the inside of the ring being clear of soot, or nearly so. From this we learn that it is the luminous part of the flame which deposits soot on a cold body—the dark zone not depositing any soot If the paper is held in the flame a little longer, it will be charred in a ring by the luminous zone of the flame.

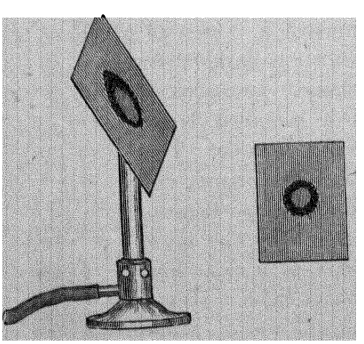


FIG. 106.—(From Newth's *Chemical Lecture Experiments.*)

Take a piece of fine wire gauze, and lower it horizontally into the flame. The gauze becomes red-hot in a ring, whilst the middle of the ring is dark (Fig. 107).

Hold a piece of thin wire horizontally in the flame, and the portion of the wire in the dark zone remains dark, whilst on each side of it that portion in the luminous zone is red-hot.

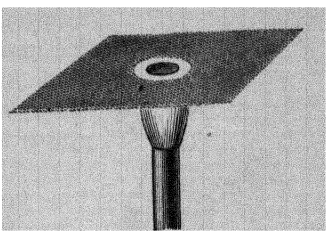


FIG. 107.—(From Furneaux's *Elementary Chemistry*.)

All these experiments go to show that the dark zone consists of unburnt, comparatively cool gases, whilst the luminous zone is the place where the gases are combining with the oxygen of the air, and evolving a large amount of heat.

The beautiful pale blue, non-luminous zone is chiefly below and around the lowest portion of the luminous zone. A piece of wire or wood held in it for a moment will show that this zone is very hot. No combustible gases can be drawn from it by a tube, and it deposits no soot on a cold body held in it. The most volatile of the candle gases are burning here, and they are supplied with an abundance of air for combustion.

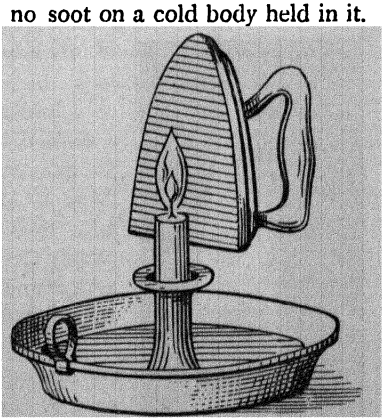


FIG. 108.

If you hold a large cold metal surface, *e.g.* the face of an ordinary flat-iron, close to the side of a candle-flame, the luminosity of the flame is immediately lessened. The flame is cooled by the cold flat-iron, and does not, therefore, give out as much light as before.

You may even extinguish a candle-flame by holding a cold metal, which carries off heat very quickly, very close to it. Wind a few turns of thick copper wire closely upon a lead pencil until the spiral thus made is about the length of the candle-wick. Slide the wire spiral over the wick of a lighted

candle, but without touching the wick. The flame is immediately extinguished. This is not because the spiral keeps air from the flame, as you might think, for if the spiral is first made red-hot in another flame, and then slipped over the candle-wick, the flame is not extinguished. The reason why the flame is extinguished by a cold spiral is that it takes so much heat from the flame that the wax or fat ceases to be vaporized.

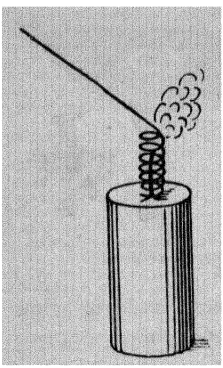


FIG. 109.

There is no soot to be obtained from the dark zone, and none from the non-luminous zone of a candle, oil, or gas-flame, whilst plenty is deposited on a cold body held in the luminous zone. Now, soot is a kind of charcoal, and charcoal consists chiefly of a substance called carbon, and it is very likely that the luminous zone owes its brightness to particles of white-hot solid carbon in it.

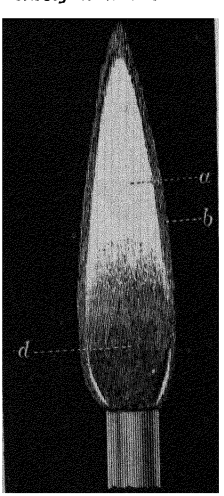


FIG. 110.—(From Newth's *Text-book of Chemistry*.)

Close the air-holes of a Bunsen burner, turn on the gas, and light it. The flame is luminous. The same three regions are as plainly visible in this flame as in a candle-flame. There are the dark central zone, *d*; the luminous zone, *a*; and the non-luminous outer zone, *b*. The central zone can be shown to consist of unburnt and comparatively cool gas. It can be drawn off by a tube and lighted at a jet, and if you thrust a pin through a match so that it can be hung in the tube of a burner with

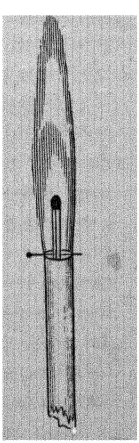


FIG. 111.—(From Newth's *Chemical Lecture Experiments*.)

the phosphorus end uppermost, the gas can be ignited without the match taking fire so long as the flame burns steadily. The luminous zone will deposit soot abundantly, and may be shown to be hotter than the dark zone by the same methods as were used in experimenting with the candle.

Suppose now you open the air-holes of the burner. The character of the flame is immediately altered. It is no longer luminous, it does not deposit soot, and it is a great deal hotter than the luminous flame. This is because what is now burning is not coal-gas surrounded by air, but a mixture of coal-gas and air surrounded by air. As you already know, air passes in at the air-holes, and becomes mixed with coal-gas issuing from the small orifice at the bottom of the tube. By the time they reach the top of the tube the gas and air are thoroughly mixed.

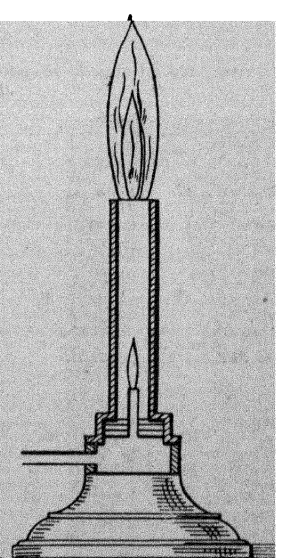


FIG. 112.

Whilst the gas is burning with a non-luminous flame very slowly turn down the tap. The flame becomes smaller and smaller, until at last it disappears. Look down the tube, or through the air-holes, and you will find a small luminous flame at the gas orifice near the bottom of the tube. Clearly, when the flame disappeared from the top, it travelled down the tube until it settled there. It is now a luminous coal-gas flame burning in air. Turn up the gas again, and hold a lighted match to the top of the tube. Here will appear a large but sickly looking, slightly luminous flame. It is evident that this is not the flame of coal-gas in air, for that is burning at the bottom of the tube.

To find out why a flame travels down the tube of the Bunsen burner when the gas-supply is carefully reduced, we

must first remember that a mixture of coal-gas and air will explode only when there is about twelve times as much air in the mixture as coal-gas.

Turn on the gas when the burner is arranged to give a non-luminous flame, but instead of igniting it, hold a large empty test-tube, or a gas-jar inverted, over the top of the tube. In a few moments this will be full of the mixture of gas and air issuing from the burner. Remove it gently, keep it inverted, and apply a light to the mouth of the tube. The mixture burns quietly at the mouth of the tube. It is clear, then, that there is not enough air here to make an explosive mixture. This is the same thing as saying there is too much gas.

When the mixture is burning at the top of the burner-tube, and the gas-supply is slowly and cautiously lessened, at last there comes a time when the mixed gas and air are in the right proportion to explode. Then the flame swiftly travels through the mixture in the tube, and settles on the gas orifice.

SUMMARY.

A *candle, oil, or gas flame* consists of *three portions*—

(*a*) Dark central zone where there are unconsumed gases.

(*b*) Luminous zone where the gases are partly consumed.

(*c*) Outer non-luminous zone where consumption is complete.

A *cooled flame* is *not so luminous* as before. *If cooled too much it is even extinguished.*

LESSON XXII.

DISCOVERY OF THE ACTIVE CONSTITUENT OF AIR.

Apparatus needed:—Test-tubes or closed hard glass tube, cork, delivery-tube, dish, Bunsen burner, and red oxide of mercury ; red lead ; lime-water and glass tube.

On August 1, 1774, Dr. Priestly, a Birmingham minister, who was in the habit of making many chemical experiments, discovered that by heating a certain powder, called by chemists red precipitate, he could obtain a gas from it quite different in its properties from any gas which any one had previously any experience of. He also obtained this gas, in which he said that a candle burnt much brighter than in common air, by heating another powder called red lead.

Priestly did not name the gas, but in the same year he told the celebrated French chemist, Lavoisier, of his discovery, and the latter immediately set to work, and ultimately showed that the gas given off from red precipitate was one of the constituents of the air—that constituent which up to now we have often called the active portion of the air. Until this time everybody thought that air was a simple gas. There was no suspicion that it was a mixture of gases.

It was known to these chemists that red precipitate contained mercury, so Lavoisier placed some mercury in a glass vessel, *a*, called a retort, the other end, *d*, of which was bent so that it dipped through a dish of liquid, *e*, into a vessel of air, *m*, standing inverted in it. The mercury in the retort was heated continuously for many days and nights, and a dark-coloured rust gathered on its surface, whilst the air in the jar

became less and less. This was shown by the liquid rising up the jar, and it proved that a portion of the air was being

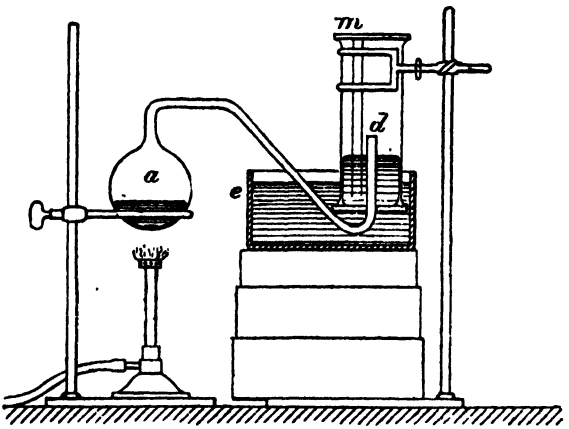


FIG. 113.—(From Jago's *Inorganic Chemistry.*)

absorbed by the mercury. When about one-fifth of the air had been absorbed, the liquid rose no further.

The experiment was then stopped, and the rust formed on the surface of the mercury carefully removed, put into a tube closed at one end, heated strongly, and the gas given off was collected in a jar. The amount of gas given off was just equal to that which the mercury had absorbed, and when it was added to the vitiated air left in the jar, it formed a mixture which was just like ordinary air.

To this gas, absorbed from air by heated mercury, and given off again when the red rust is more strongly heated, Lavoisier gave the name *oxygen*. We have seen that oxygen forms one-fifth of the volume of the air; to the remaining four-fifths, which up to now we have generally called vitiated air, the name nitrogen has been given.

To form this red rust of mercury by heating it in air is a long and tedious process. It is very easy, however, to get oxygen from this red powder, which we call nowadays *red oxide of mercury*.

Take a piece of hard glass tubing closed at one end. Fit the other end with a cork and a glass delivery-tube bent to the shape shown in the figure. Place a small quantity of red oxide of mercury in the bottom of the tube, and replace the cork. Allow the end of the delivery-tube to dip in the water of a small dish in which is an inverted test-tube full of water. Heat the red powder strongly with a Bunsen burner or spirit-lamp, and gas will be given off, which will bubble into the test-tube and fill it. At the same time you will notice that in the tube which contains the red precipitate a coating of small drops of mercury has formed on the sides of the tube over the powder. Of course, when the heat has driven off the oxygen, the mercury which is left is vaporized, and condenses again on the tube in the manner you observe.

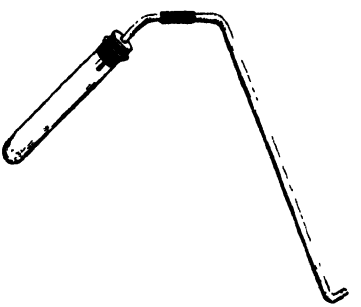


FIG. 114.—(From Newth's *Elementary Practical Chemistry*.)

Take the test-tube full of oxygen, put your thumb over its mouth, and withdraw it from the dish. Take a splinter of wood with a spark on the end of it, and insert it into the tube. The spark immediately develops into a flame. This is how you will, in future, generally recognize oxygen gas.

The rust which forms on lead when it is kept melted in air is yellow when cold. This rust is commonly called *litharge*, and it is formed, as you already know, from lead and oxygen. In other words, it is a *compound* of lead and oxygen. We call lead and oxygen *elements*, because we do not know any two or more substances which unite to form lead or to form oxygen.

You may heat litharge, but you will not be able to drive the oxygen from it and leave the lead behind. This is what you might have expected, because, since lead combines so readily with oxygen, it is because the metal has a strong liking for the gas; and therefore, being once united, their separation

will be difficult. This is the case also with the compound formed when phosphorus unites with oxygen, and also with chalk-gas and water, which are formed when the substance of a candle, or oil, or coal-gas unite with oxygen. If two boys like each other so much that they are always together, it is not easy to separate them.

But if we put litharge into a little crucible or iron spoon, and heat it strongly for a long time, stirring it frequently with an iron rod so that the air can get at all of it, the litharge unites with more oxygen, and forms a red powder, which is called red lead. This additional amount of oxygen, got so reluctantly, can be readily driven off. Put some of it into a closed hard glass tube, heat it strongly in the flame of the Bunsen burner, and push a splinter of wood with a little spark on it down the tube. The spark becomes a flame, showing that oxygen has been driven off from the red lead. After a good heating the yellow litharge is left in the tube.

You have seen through numerous experiments that the air consists of four-fifths of its volume vitiated air or nitrogen, and one-fifth active air or oxygen. This proportion is nearly constant, wherever we take the air we experiment with. In this country, and in all other countries, in deep valleys, on high mountains, or still higher in the atmosphere, over land and over the ocean, in the coldest as well as in the hottest regions of the earth, the proportion of oxygen to nitrogen is almost exactly the same.

It is the oxygen which is the necessary portion of the air, for no animal and no plant could exist without it. When the oxygen is taken from the air no life is possible in it. We breathe air into our lungs, and in all parts of our bodies the oxygen unites with carbon, giving us heat, and forming chalk-gas, which we breathe out again. If you place some clear lime-water in a glass, and breathe through it by means of a tube, the lime-water becomes milky, which is a clear proof that you are breathing out chalk-gas.

By means of the apparatus illustrated in Fig. 115, this can easily be seen. A and B are bottles, each containing lime-water. They are fitted up with tubes as shown in the figure,

and if the mouth be applied to E, all the air entering the lungs must bubble through the lime-water in B, whilst all the air leaving them bubbles through A. After a few breaths the liquid in A is milky, whilst that in B has remained clear.

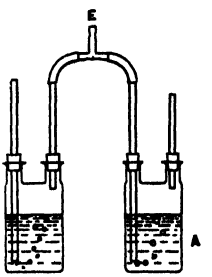


FIG. 115.—(From Newth's *Chemical Lecture Experiments.*)

You remember that by boiling ordinary water we were able to drive out from it a certain amount of air which it had dissolved. It is this small amount of air in water which supplies fish and water-plants with the oxygen which is necessary for their life. Neither fish nor water-plants can live in boiled water. The fish would be suffocated at once, and the plants also would quickly die.

More than two hundred years ago, Robert Boyle, an English chemist and man of science, experimented with the air-pump which had just then been invented. He describes how a lark and a sparrow were placed under the glass of an air-pump by him, and how both birds died within a few minutes of the air being pumped out, being suffocated for want of air. It seems a cruel experiment, but it proved for the first time that animals cannot live in a space which contains no air. And we know, although Robert Boyle never knew, that it was not want of nitrogen which killed them, but want of oxygen.

SUMMARY.

Dr. Priestly *first prepared oxygen* (active air) from red precipitate *in* 1774.

Red precipitate, or *red oxide of mercury*, is the *red rust* formed when *mercury* is kept *heated in air.*

Oxygen is driven off from it by heat.

Things burn very *brightly* in *oxygen.*

Oxygen may be *driven off* from the *red rust of lead* by heat.

All air consists of about *one-fifth* of its volume *oxygen* and *four-fifths nitrogen.*

LESSON XXIII

MORE ABOUT OXYGEN.

Apparatus needed:—Test-tube, cork, and delivery-tube, dish, Bunsen burner, retort-stand, chlorate of potash, black oxide of manganese, gas-jars, glass plates; candle, wire, charcoal; deflagrating-spoon, sulphur, phosphorus, sodium, iron wire, magnesium ribbon.

THERE is no substance, so far as we know, which is so abundant in the world as oxygen. Quite one half of that portion of the surface of the solid earth which is within our reach consists of oxygen. Eight parts by weight in every nine of water is oxygen; whilst, as you know, one-fifth of the total volume of the air is this same gas.

Since oxygen is so abundant, it ought to be very easy to obtain large quantities of it. In the air it is not united with any other substance, but only mixed with nitrogen and other gases in very small quantities, and yet we do not know of any practicable way of taking the nitrogen out of air, leaving us the oxygen unmixed. And oxygen has such a strong desire to unite with other substances that its separation from them after union is, as we have seen, not always an easy matter.

When we desire to obtain a fairly large supply of this gas to experiment with, the most convenient way is to heat a white crystalline substance which is called chlorate of potash. When this is strongly heated, the oxygen is driven off from it, and may be collected in inverted jars filled with water. It has been found, however, that if powdered chlorate of potash is first mixed with a small quantity of a certain black powder called manganese dioxide, or even with a little fine sand, the oxygen is given off at a much lower temperature.

Therefore powder up a little chlorate of potash, and mix

well with it about one-sixth of its weight of manganese dioxide. Fit a long test-tube with a cork and a delivery-tube of the shape shown in the figure. The cork must be a sound one,

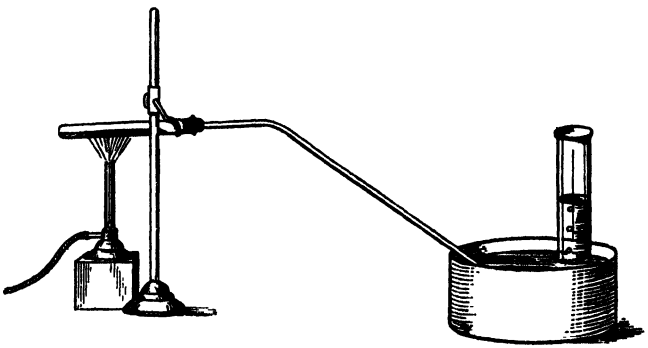


FIG. 116.

and must fit well. Fill the test-tube about one-third full of the oxygen mixture, replace the cork, and fix it in a nearly horizontal position, the mixture being spread along the tube nearly to the cork. The delivery-tube must dip into a basin or dish of water, and in the dish must stand the gas-jar full of water and inverted.

Then heat the tube of oxygen mixture, cautiously approaching the flame to it, and moving the flame to and fro from end to end of the tube. In a minute or so bubbles of gas will leave the delivery-tube and rise up through the water in the dish. Then if the open end of the gas-jar is held over the bubbles, they will rise up into the jar, displacing the water.

As each jar is filled with gas, slide a greased glass plate under its open end, lift it out, and place it on one side. Or, instead of a glass plate, you may use a saucer, and lift out each jar with its mouth in a saucer of water. The water in the saucer will seal the jar, and prevent the outer air entering it, or the oxygen leaving it.

You will note that the gas is colourless and transparent. If you smell at one of the jars you will find that it has no odour and no taste.

Lower a lighted candle into a jar of oxygen. The candle

burns very brightly indeed, but the gas does not burn at all. Throw a little lime-water into the jar when the candle is withdrawn, and shake it up. The lime-water becomes milky, showing that chalk-gas was produced when the candle burnt in it. You remember that the same gas was produced when a candle burnt in air. The burning of a candle in air and in oxygen is therefore the same process as you would expect, but the burning is much more vigorous in pure oxygen than in a mixture of oxygen and nitrogen.

Bend the end of a stout wire around a piece of charcoal, or place it in a deflagrating-spoon, hold the charcoal in a flame until a portion of it is red-hot, and then lower it into a jar of oxygen. The charcoal burns very brilliantly indeed, especially if it is a piece made from bark of a tree. When it has burnt for a while, test the gas in the jar with lime-water. The lime-water becomes milky with chalk, and the gas is therefore chalk-gas. Charcoal consists very largely of a substance called carbon. Chalk-gas is formed by the union of carbon and oxygen, and it is on that account called by chemists an *oxide of carbon.*

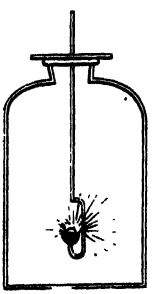


FIG. 117.—(From Furneaux's *Elementary Chemistry.*)

For the next experiment you will also require a deflagrating-spoon. You can easily make one, if necessary, with a piece of wire and a small piece of tin plate. Take some sulphur, about the size of a pea, place it in the spoon, hold it in a flame until the sulphur begins to burn, and lower it into a jar of oxygen. The sulphur continues to burn with a very brilliant violet flame. After withdrawing the spoon, throw in a little water, and shake it up. Then taste a drop of the water, and you will find it is sour.

The sulphur in burning has combined with the oxygen, forming an oxide of sulphur. This is a colourless gas, with a very disagreeable smell, as you well know. It easily dissolves in water, and forms the sour solution which we call an *acid.*

Take a very small piece of phosphorus, dry it, and place it in the deflagrating-spoon. Touch it with a hot wire, and immediately it begins to burn place it in a jar of oxygen gas.

It burns vigorously, with a very dazzling light, throwing out the white fumes formed by the union of phosphorus and oxygen, which we may now call *oxide of phosphorus*. This is very soluble in water, and forms an acid solution.

Take a small piece of the soft metal sodium, place it in the deflagrating-spoon, and heat it strongly in the Bunsen or spirit flame until it begins to burn. Then introduce it into a jar of oxygen, where it will continue to burn. White fumes will be produced, which show that the oxide of sodium is a white solid. After the spoon is withdrawn, throw in a little water, and shake it up. The fumes dissolve in the water. Taste the solution. It is not sour, but has a nasty soapy taste, and if you dip your fingers into it, it feels soapy. The solution is therefore *not an acid*. We call it an *alkali*.

Papers may be bought very cheaply stained with a substance called *litmus*. Some of these are red and some are blue. Take one of the blue litmus-papers, and place it in the acid solution of oxide of sulphur or oxide of phosphorus, and the blue litmus on the paper instantly becomes red. Take the reddened litmus-paper, and place it in the alkaline solution of oxide of sodium. It is again turned blue. These litmus-papers are very convenient for testing whether a liquid is an acid or an alkali.

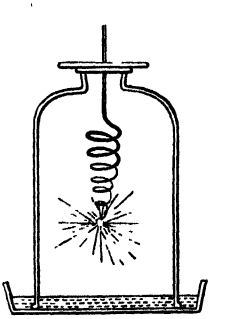


FIG. 128.—(From Furneaux's *Elementary Chemistry*.)

Take a piece of clean thin iron wire, and wind the end of it into a little loose spiral around a lead pencil. Make the end of the spiral red-hot, and dip it into a little powdered sulphur. Some of the sulphur attaches itself to the iron wire, and burns, or can be set on fire. Whilst the sulphur is burning, introduce the spiral into a jar of oxygen, which should have a little water at the bottom. The wire catches fire, burns brightly, and throws off drops of melted oxide of iron, which fall into the water at the bottom of the jar. This oxide is not soluble in water, and it has no effect on either red or blue litmus-paper. It is therefore neither acid nor alkaline,

but *neutral.* It is not the familiar red rust which is found when iron slowly combines with oxygen in moist air, but the black oxide produced when iron is heated in air. We see, then, that there is more than one oxide of iron. They differ from each other in the proportions of oxygen they contain.

The oxides formed when substances burn in oxygen are the same as those produced when these substances are burnt in air. Thus bodies such as the candle, oil, and coal-gas always produce oxide of carbon and water whether they are burnt in pure oxygen or in air. A piece of wood also forms oxide of carbon whether burnt in air or oxygen, showing that wood contains carbon. The oxides of sulphur, phosphorus, or sodium which are formed when these substances are burnt in air, are similar to those formed when they burn in oxygen. In all cases of burning in air, the combination is with oxygen, not with nitrogen.

In Lesson XV. an experiment was described in which magnesium ribbon was burnt in air, producing an oxide which weighed more than the magnesium. The amount of increase in weight was, of course, the weight of the oxygen which united with the magnesium whilst burning. If this experiment is performed in pure oxygen, there is exactly the same increase in weight for the same mass of magnesium used.

SUMMARY.

Oxygen is usually *prepared* by *heating chlorate of potash.*

A *candle burns brightly in oxygen, producing chalk-gas and water*, just as when burning in air.

A *piece of charcoal burns in oxygen, producing chalk-gas*, just as when burning in air.

Sulphur burns in oxygen, producing a suffocating gas, just as when burning in air.

Phosphorus burns in oxygen, producing a white powdery oxide, just as when burning in air.

Sodium burns in oxygen, producing a white oxide.

Iron burns in oxygen, producing a black oxide.

In all cases, burning in oxygen is more vigorous than in air, but the same substances are produced. Burning in air is combining with the oxygen of the air.

APPENDIX

THE BALANCE.

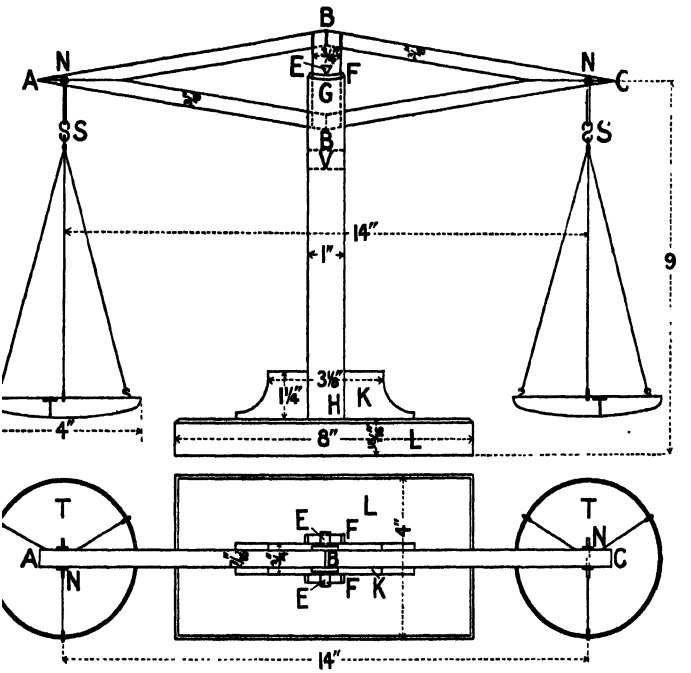
A BALANCE is necessary in any school taking a Natural Science Course. If bought from the dealers it involves a considerable expense, and although it is almost universally thought that its construction is beyond the powers of the ordinary amateur and even craftsman, this is a mistake.

On pp. 8–10 of his "Chemical Lecture Experiments" (Longmans and Co.), Mr. G. S. Newth describes the construction of a home-made balance, mainly of wood, which is very simple in construction, efficient in use, and quite accurate enough for elementary experiments. One has been made in the workshop of this school, which differs from Mr. Newth's in but one or two small particulars. The following description, together with the figures, should be a sufficient guide to any person who can use a few simple tools in making a similar one.

The instrument may be said to consist of four parts, viz. the beam, the pillar, the base, and the pans. Any straight-grained hard wood is suitable to use, although mahogany is perhaps the best. Canary pine (American bass wood, *Tilia Americana*) is another wood which will do very well, and there is no other reason why the softer red, white, or yellow pines should not be used, except that they *are* soft and easily dented and broken.

Fig. 1, which is an elevation of the balance, gives the necessary details for the construction of the beam. This is a parallelogram made of four pieces of wood, joined by glue or seccotine at the angles. A drawing of this should first be made, full size, on paper, making the outside measurements of the parallelogram $15\frac{1}{2}'' \times 2\frac{1}{2}''$. From this drawing the exact angles at which the pieces must be cut can be obtained. A solid piece, X (Fig. 3), is placed across the shortened diagonal, and two thin pieces, W, W,

are placed outside it. The whole are then firmly cemented together. When firm, two holes must be bored horizontally through the beam at NN, 14 inches apart, and ¾ inch from each end.



FIGS. 1, 2.

Through each hole is firmly wedged about an inch of wire cut off from a stout knitting-needle.

At E a piece of an old triangular file, about 1¼ inches long, is passed horizontally through the beam. To prepare the file, which should be a small one, place it in the fire until it is red-hot. Then take it out, and allow it to cool slowly. This will make it soft so that it can be readily cut with a hack-saw or cold chisel. Before cutting it, however, grind out all the file-marks on the portion you intend to use. When cut to the proper length and its ends squared up, it may be made red-hot again, and then dropped into cold water. This will re-harden it.

Now place one end of the piece of file on a piece of wood, mark out exactly the size of its triangle, and select a bit which will bore a round hole which lies entirely in the triangle. Prove this by actual boring through a waste piece of wood. The round hole can be converted into a triangular one by means of a narrow chisel.

Then, laying a straight-edge across the beam from A to C, draw a line across the middle of the shorter diagonal BB, and bore an exactly horizontal hole in such a position that when cut into a triangular hole, the lowest angle of the triangle shall be about $\frac{1}{8}$ inch above the middle line. Then wedge the piece of file through the hole with its sharpest angle downward.

The base consists of two pieces —the one, L, is 8″ × 4″ × 1″, and screwed from underneath to the middle of this is the other, K, 4″ × $1\frac{1}{4}$″ × $\frac{3}{4}$″. The ends of this may or may not be cut away as shown in the figure.

The pillar is composed of two pieces, G, H, each $8\frac{1}{2}$″ × 1″ × $\frac{3}{16}$″. They are screwed to the block K, and have a solid piece, V, glued between them near the top to give rigidity. Their tops are slightly hollowed, and to these hollow parts are cemented with shellac or seccotine two pieces of glass tubing, F, F. The tubing should be softened in the flame, and then bent partly around a large bottle in order to get an even curve. Pieces of the right length can then be cut off with a file, and the pillar pieces cut to the same curve.

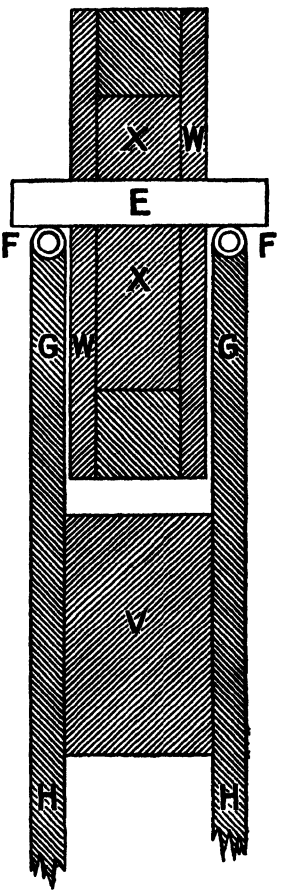


FIG. 3.—Showing section through BB (Fig. 1). Full size.

When the balance has been thus far put together, the beam should be suspended in position on the glass tops of the pillar

pieces for trial. It should balance quite horizontally when at rest, and when set in motion it should oscillate "lively." If it does not answer to these tests, it will be as well to attach to the outside of the bottom of the beam a strip of sheet lead, the width of the wood, and about 4 inches long. Then portions of this can be cut or scraped away until it balances perfectly.

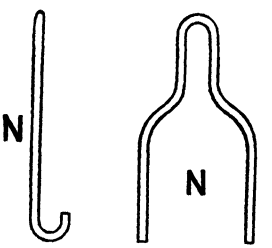


FIG. 4.—Suspension hook.

Take a piece of tolerably stout brass or iron wire, and bend two hooks of the shape and size given in Fig. 4. These are then hung from the pieces of knitting-needle at NN (Fig. 1).

The pans are each suspended by three silk threads attached to S-shaped hooks. These little hooks can be easily bent out of large pins. Brass balance-pans, of equal weight, can be bought from dealers in scientific apparatus for about 2*s.* per pair. Instead of brass pans, tin sand-baths, costing about 1*d.* each, may be made to serve. Holes must be bored near the rims, and care must be taken that they are equidistant. After the pans are suspended the balance may need further slight adjustment before it is ready for use.

The sizes here given are suitable for a balance which may be used for demonstrating before an ordinary-sized class. But for larger numbers the balance may be increased in size so long as the proportions are kept.

A set of gram weights, suitable for such a balance, may be bought for about 3*s.* 6*d.*

THE END.

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