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MODERN SCIENCE

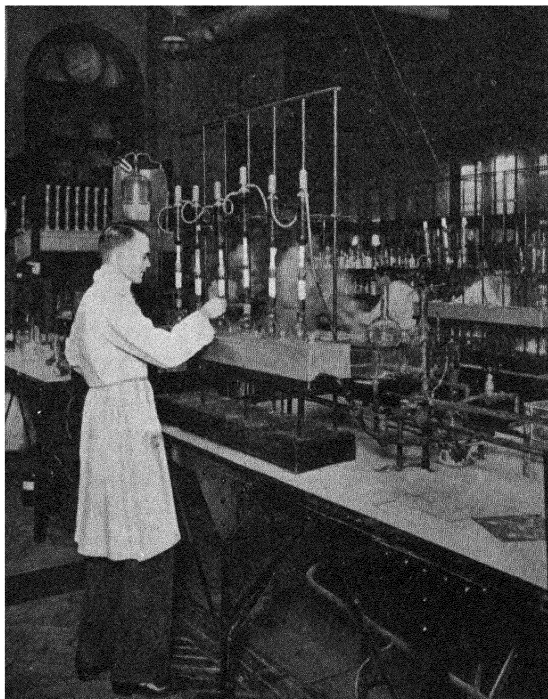
BOOK II. CHEMISTRY



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A CHEMIST AT WORK
(Courtesy, Lever Bros.)

MODERN SCIENCE

BOOK II
CHEMISTRY

BY

G. W. MANFIELD, B.Sc. (LOND.)

PARK SENIOR MIXED SCHOOL, DAGENHAM

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PREFACE

THIS second volume of the series is devoted to Chemistry, a subject which is bristling with difficulties. Once again the academic side has been omitted and Chemistry has been presented in its simplest form, devoid of symbols and formulae and without reference to atoms and molecules, or to those theories on which the science is built.

Here, an attempt has been made to deal with those substances and their reactions which have the greatest bearing on everyday life.

To encourage reading—and no science teaching can be complete without private study and research—a selection of suitable reference and interest books is given. This list is merely suggestive, and many excellent volumes have been omitted through pressure of space. Special emphasis may well be placed on the use of the descriptive booklets issued by some of the leading manufacturers. These provide up-to-date material and excellent diagrams.

At the end of each chapter the briefest of summaries is given, not in a form suitable for copying into a note-book, but rather such that the scholar is guided as to the notes he should make for his own use.

My thanks are due to those who have made this work possible ; to the publishers for their careful production,

to the many firms who have loaned me photographs and diagrams, and to Mr. Donald Smith, F.R.G.S., F.R.Hist.S., F.Coll.H., for his great help and advice in the preparation.

G. W. M.

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BOOKS TO READ AND STUDY

No single science text-book will teach you all there is to learn, and from one of this size you will only get the first ideas of the subject.

Perhaps you will become eager for more knowledge, and to help you in your choice of books a list of those most suitable is given.

Cultivate the habit of reading about science and about the men who are regarded as great scientists. You will find that these men started when they were young, and that they persevered in spite of many disappointments and have discovered many new laws, many more substances, and have made possible many of the amenities of our modern lives.

Master Minds of Modern Science, by T. C. Bridges and H. H. Tiltman. (Harrap.) 7s. 6d.

Heroes of Civilization, by J. Cottler and H. Taffe. (Harrap.) 7s. 6d.

The Master Thinkers, by R. J. Harvey-Gibson. (Nelson.) 2s. 6d.

Pioneers of Science and Invention, by C. W. Long. (Macmillan.) 7d.

The Book of Chemical Discovery, by L. A. Coles. (Harrap.) 7s. 6d.

Discoveries in Chemistry. (Blackie.) 1s. 3d.

Achievements of Chemical Science, by J. C. Phillip. (Macmillan.) 2s. 6d.

Modern Chemistry : the Romance of Modern Chemical Discoveries, by F. Prescott. (Samson Low). 12s. 6d.

- Chemistry in the Service of Man*, by A. Findlay. (Longmans, Green.) 6s.
- Coal and the Coal Miner*. (Blackie.) 1s. 3d.
- About Coal and Oil*. (Blackie.) 1s. 3d.
- Coal Tar*, by A. R. Warnes. (Pitman.) 3s.
- Gas and Gas Making*, by W. H. Y. Webber. (Pitman.) 3s.
- Domestic Utilization of Gas*, by N. S. Smith and R. N. LeFevre. (Walter King.) 5s. 6d.
- Iron and its Workers*. (Blackie.) 1s. 3d.
- Iron and Steel*, by C. Hood. (Pitman.) 3s.
- Chemistry in the World's Work*, by H. E. Howe. (Chapman & Hall.) 15s.
- A First Book of Chemistry*, by W. A. Whitton. (Macmillan.) 2s. 6d.
- The Chemistry of the Garden*, by H. H. Cousins. (Macmillan.) 2s.
- Lime and Cement*. (Blackie.) 1s. 3d.
- Soap*, by W. A. Simmons. (Pitman.) 3s.
- How Photography came about*. (Blackie.) 1s. 3d.
- Practical Mechanics*. (George Newnes.) 6d. monthly.
- Chemistry in Commerce*. (George Newnes.) 1s. weekly.
- Pamphlets issued by the leading Manufacturers and Industrial Undertakings.

CHAPTER I

AIR

IN the previous book you have studied how man uses the wonderful forces of Nature and the laws which govern them. Here you will study *substances* or *materials* and discover how they are made and how they can be used.

Let us take the most common of all substances, *air*, and by experiment, find out what it is made of.

Burning in Air

First of all we will investigate what happens when substances burn in air. If you have tried to start a bonfire with old magazines you will have discovered that your efforts meet with very little success if the books are placed in a heap. The books burn only at the edges while the centre remains untouched, and yet, if the pages are torn out and lightly screwed up, the fire blazes away quite merrily.

Experiment I

Place a lighted candle on the bench and cover it with a jar (Fig. 1).

Observe the effect of the jar on the flame.

Why does the candle go out? There is still some left that could burn if it would. Now repeat your experiment, and just as the flame is about to go out, raise the jar an inch or so. The flame

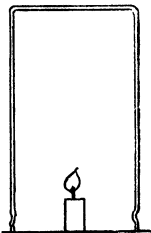


FIG. 1.

brightens and the candle continues to burn away. Replace the jar and it will go out as before.

These simple experiments show you that air is necessary for substances to burn and that burning ceases within a short time after the supply has been cut off.

For this reason we throw a rug over a person whose clothes are on fire, and use the bellows to pump air into a fire which is not burning well. In factories, where large furnaces are used, a constant supply of air is forced in by pumps or fans and the fire burns fiercely, giving a great heat.

Experiment 2

Take a piece of magnesium wire or ribbon and hold one end with a pair of tongs. Heat the other end until it burns and notice carefully what is formed.

Experiment 3

Take a tiny piece of phosphorus (do not hold it with your fingers) and place it in a tin lid. On a warm day you will find that it begins to burn without heating, but, if it is cold, warm the tin gently.

Observe what happens.

In Experiment 2, you will have found that the shiny wire or ribbon changes to a white ash (*magnesia*) when it burns. Some of this floats away in the air, but the remainder will stay in the form of a wire if your hand is steady.

The phosphorus, of which you will learn more later, easily takes fire and produces clouds of white fumes.

These are poisonous and for this reason only a tiny piece of substance is used. These fumes float away and in a short while are lost to view, and the phosphorus itself has disappeared.

Now these two experiments show you that when substances burn in air, they change considerably. The magnesium changes to a white ash or powder, while the phosphorus becomes white fumes.

With *burning* we usually associate light and heat, thus the flame from burning oil or from coal or from magnesium gives both, but a piece of smouldering brown paper is also burning, and while heat is produced, there is very little light. On the other hand, the wire or filament in a modern electric fire gives both light and heat, and yet we should not say that it was burning.

Burning or *combustion*, as it is called, implies a change in the form of the substance. The magnesium turns to a white ash, coal gives a grey ash with a certain amount of smoke and soot, but the filament in an electric fire or lamp does not change sufficiently for us to notice.

To explain this change, a German scientist, Stahl, who lived in the eighteenth century, considered that all combustible materials contained a substance which he called *phlogiston* (from the Greek word meaning—I set on fire). He thought that when they burnt, this substance was released, leaving behind a different substance, which he called a *calx*.

i.e. Substance \rightarrow phlogiston + calx.

Bear this explanation in mind when you work the following experiments.

Experiment 4

Weigh a crucible with its lid. Then insert a small piece of magnesium wire loosely coiled. Again weigh and by subtraction find the weight of magnesium used.

Place the crucible on a tripod, with the lid only partly covering the top, and heat it for some minutes.

When the magnesium appears to have completely changed to ash, remove the crucible and, when cold, weigh again.

Heat again for some minutes and reweigh. Enter your results as follows :

(a) Before Heating

Weight of empty crucible and lid - — grms.

Weight of crucible, lid and magnesium — grms.

∴ Weight of magnesium - - - — grms.

(b) After Heating

First Time

Weight of crucible and contents - - — grms.

∴ Weight of contents - - - — grms.

Second Time

Weight of crucible and contents - - — grms.

∴ Weight of contents - - - — grms.

If the results of the second weighing are not the same as those of the first, reheat and weigh again until they agree.

(a) Has the magnesium gained or lost weight during this experiment ?

(b) Does your result agree with Stahl's theory ?

Experiment 5

Repeat Experiment 4 using small pieces of zinc. You will have to heat the zinc a little longer.

Does your result here agree with that obtained in Experiment 4?

By using a crucible with a lid the ash is prevented from escaping into the air, so that what remains in the crucible is the complete product of combustion.

If Stahl's theory is correct, then you would expect the weight to be less after the substance has been burnt, but your results will show you that it is more. The magnesia weighs more than the original magnesium, and it would appear that something is added to the substance rather than taken from it.

A French scientist, Lavoisier, discovered this in 1774. He found that when tin was heated, it gained weight, but he went a step further and proved that the air lost just as much weight as that gained by the tin. Thus he showed that Stahl's theory was wrong and that instead of the substance losing phlogiston, a part of the air became joined to it, and the statement to represent this change will be :

Substance + part of the air \rightarrow calx.

Rusting

You will be familiar with the state of a piece of iron after it has been left in the damp air for a day or two. It quickly loses its brightness and becomes covered with a reddish coloured *rust*. Perhaps you will wonder what this has to do with burning, but you must have patience for a while and you will find that burning and rusting are very closely connected.

Experiment 6

Make the inside of a test-tube damp and sprinkle some iron filings in the closed end of the tube.

Invert the tube and place it with its open end in water (Fig. 2). Mark the level of the water in the tube and allow it to stand for a few days.

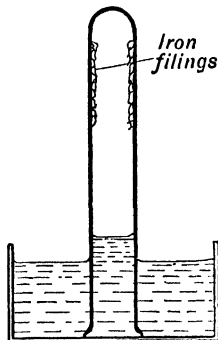


FIG. 2.

After a day or two or perhaps a week, you will find that your iron filings are rusty, and on measuring the amount of air in the tube, you will find that some of it has been used—evidently during the process of rusting.

Writing this in the form of a simple statement, we shall have :

Iron + part of moist air \rightarrow iron + iron rust.

This change in the iron is very similar to the change in the magnesium wire, but it is not so complete since some of the iron is left behind. Rusting is sometimes known as *slow burning*, and if iron were left for a long time in contact with an unlimited amount of moist air, it would change completely to the red rust, just as the magnesium changes to magnesia.

Iron is used extensively in buildings ; there are iron girders and pillars, iron windows, iron railings and sheet iron for the roofs of sheds. All these would rust and crumble to powder if they were not protected in some way.

Just as a piece of canvas is coated with rubber to make

a mackintosh (Book I, p. 35), so the simplest way of protecting iron is to cover it with some substance which will exclude the air.

Sheet iron is often covered with zinc—a metal which does not rust. This used to be done by the aid of electricity, and the iron was said to be *galvanised*, but in the modern process, sheets of iron are immersed in large tanks containing molten zinc. The zinc adheres to the iron, giving it a protecting coat through which water and air cannot pass. In the course of time, however, the zinc wears off, leaving the iron exposed, and rusting begins, with the result that the sheet soon perishes.

By a similar process, iron is often coated with tin—another metal which does not rust. The resulting sheets are made into the cans which hold our tinned foods, and into kettles and saucepans. Iron, when treated in this way, is usually spoken of as *tin plate* and when we speak of cocoa *tins*, *tin* kettles and so on, we must remember that they are not made from pure tin, which is very expensive, but from iron coated with a thin layer of tin for protecting purposes. Tin is more satisfactory than zinc since it is not affected by weak acids, and while it is still used for those cans which contain fruit and foods generally, its high price makes it unsuitable for other articles. As a result, other cheaper metals, chiefly zinc, are mixed with it and used for this purpose.

Another common method of protection is by the use of paint or varnish. Oil paint is made by mixing a coloured substance, often white lead coloured by a pigment, with linseed oil (made by crushing flax seed). This mixture dries with a hard, waterproof surface and is an ideal

coating for iron girders, window frames and railings. Varnish, too, may be used for this purpose, but it differs from paint in having practically no colour. It dries with a very hard, waterproof surface and is made from special gums and waxes.

Most metals change slightly when left in contact with air—copper and brass soon become dull on the surface, and even zinc, used to protect iron, changes. With these metals, however, the change is so slow that it would take very many years for them to crumble. *Tarnishing* is the name given to this change and those of you who have brass ornaments or copper kettles, know that they have to be cleaned regularly in order to keep them bright. By cleaning, the thin film of calx, formed on the surface when the metal uses part of the air, is removed.

What Air is made of

So far we have considered air as being just one substance, but now the time comes to investigate this to find out if this is correct.

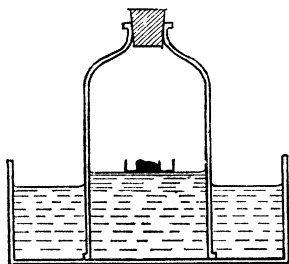


FIG. 3.

Experiment 7

Place a piece of phosphorus about the size of a pea in a tin lid and float this on the water in a trough.

Cover this with a bell-jar (with the cork out) and mark the level of the water in the jar by means of a piece of gummed paper.

Ignite the phosphorus by touching it with a red-hot skewer and quickly cork the jar (Fig. 3).

Allow the fumes to subside and observe the level of the water in the bell-jar.

Experiment 8

Remove the cork from the bell-jar and plunge in a lighted taper.

(a) Does the taper burn when it is in the jar?

(b) Is there some phosphorus still unburnt? You can test this by warming the tin lid.

You have already seen that part of the air is used up during burning, but in Expt. 7 you have found that after a time the phosphorus stops burning, but not because there is none left, for there is still some left in the dish. When the fumes dissolve in the water, the water rises in the jar, but it does not fill it. Evidently then all of the air has not been used, and as the phosphorus goes out and the lighted taper is extinguished, it cannot be air that is left in the jar. Thus, by burning, something must be extracted from the air, leaving behind something that will not allow things to burn in it.

That part of the air used up is called *oxygen*, and left behind is another gas called *nitrogen*, together with small quantities of other gases. Roughly, the atmosphere at earth level is made up of a mixture of 21 per cent. oxygen, 78 per cent. nitrogen, and 1 per cent. of other gases. This belt of atmosphere extends for about $6\frac{1}{4}$ miles and then gradually changes to almost pure nitrogen. Scientists who have recently explored the upper atmosphere or *stratosphere* have been forced to use sealed gondolas, and to provide themselves with supplies of artificial air or oxygen in order to keep alive.

One other substance is present in the air, and this you can prove by the following experiment is water vapour. The amount varies from day to day, and is more in some places than in others.

Experiment 9

Bring a bottle of cold water into a warm room and observe the thin film or mist which immediately forms on the outside of the bottle.

Why is it that this thin mist soon disappears ?

The presence of the water vapour on the outside of the bottle explains why iron will rust when left in contact with the atmosphere, and also accounts for the formation of mists, fog and dew.

In the next chapter you will study what really happens when things are burnt in air and learn something about that part of the air which we call oxygen. For the time being, we must mention just one or two more facts about *nitrogen*.

Nitrogen loves its freedom; it is what the chemist calls *inert*, that is *lazy*, as it will not easily join up with other substances. Yet nitrogen is not useless—very far from it, since no plant or animal life would be possible without it.

Animals and human beings obtain their necessary supplies of nitrogen from the foods they eat, while the majority of plants, even though surrounded by it in abundance, have to obtain it from the soil. This is arranged by the use of special manures and fertilisers containing nitrogen, or by growing crops, such as peas, beans or clover, which contain in their roots wonderful bacteria. These tiny living bodies have the power of

extracting nitrogen from the air and of passing it over to the plant. By the decay of such plants the soil becomes rich in nitrogen (see Chapter 11).

SUMMARY OF THE CHAPTER

Things to Note

1. What happens when substances burn or rust in air.
2. The substances contained in air.
3. The methods of preventing rusting.
4. The uses of tin plate and galvanised iron.

CHAPTER 2

MORE ABOUT OXYGEN

How Oxygen is made

THE name *oxygen* has already been given to a part of the atmosphere. You have learnt that it is this part of the air which is used up when substances burn, and here we are going to study the gas a little more fully and discover some of its characteristics and uses.

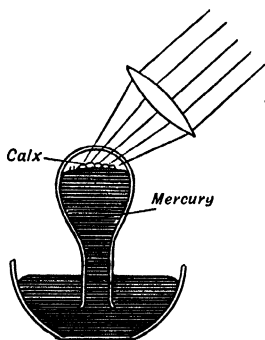


FIG. 4.

Just before Lavoisier's experiments on combustion, an English scientist, Priestley, discovered that gases were given off when certain substances were heated. He made a study of these gases or *airs*, as he called them, using very rough apparatus and heating the substances by means of a burning glass.

Among others, he discovered *oxygen* and this he obtained by heating a red substance called *mercurius calcinatus* (i.e. mercury calx). In Fig. 4 is shown apparatus similar to that which he used. The bottle was filled with quicksilver and inverted with its mouth under the surface of some more quicksilver in a basin. The red powder was introduced into the neck of the bottle and floated to the top, where it was heated. A gas was given off which pushed out the mercury and gradually filled the bottle.

Priestley found that this gas had rather remarkable characteristics; a candle burnt in it with increased brilliance, while a mouse lived in it far longer than it would live in the same quantity of air.

You can prepare oxygen in the science room using materials similar to those used by Priestley, but there is a better and more satisfactory method which is also described.

Experiment 1

Place some red mercury oxide in a small ignition tube. Hold the tube in a pair of tongs and heat the red powder strongly for several minutes.

Hold a glowing splint at the mouth of the tube.

Examine the sides of the tube and note the formation of tiny beads of mercury.

Experiment 2

Mix together powdered potassium chlorate and manganese dioxide in the ratio 3 : 1 and place the mixture in

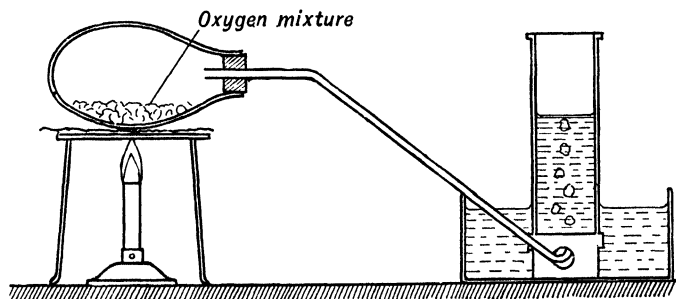


FIG. 5.

a hard glass flask. Arrange the apparatus as in Fig. 5 and heat the flask strongly. Collect the gas in gas jars.

Experiment 3

Use one jar of the gas to test its smell.

Into another jar plunge a glowing splint.

Heat a piece of carbon to red heat in a spoon and place this in the gas. Repeat, using sulphur and red-hot iron filings.

Describe carefully what happens in each case and make a note of all the properties of the gas as shown by these experiments.

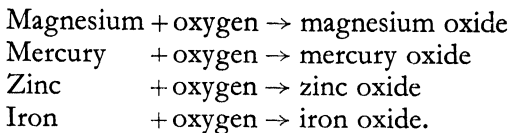
When Substances burn—Oxides

In the last chapter you learnt that substances use up part of the oxygen in the air when they burn. Magnesium changed to a white powder and showed an increase in weight ; iron, if heated in air for some time, will become covered with a thin scale (called smithy scale) and also shows an increase in weight ; mercury gives a red powder (mercurius calcinatus) and zinc a white powder.

Now the burning or heating of these substances gives us entirely new substances ; there is no similarity between the magnesium ribbon and magnesia, or between zinc and the white zinc powder. These new substances we call *oxides*, because they are formed when oxygen is used up. Thus magnesia is magnesium oxide, smithy scale is iron oxide, white zinc powder is zinc oxide, and so on.

This process, by which one substance changes to another which is quite different, is called *combination*, and the new substance formed is called a *compound*, since it has been made or compounded of two or more separate substances.

Writing these changes in the form of simple statements or equations we have :



You must be very careful over your use of the term *compound*. For instance, if you mix together two such substances as sand and sugar, you will not obtain a new substance, but one which has the properties or characteristics of both sand and sugar. This *mixture*, as it is called, is not a compound, for there has been no change in either of the substances—it is possible to separate them by mixing with water and filtering the liquid obtained.

When combination takes the form of a substance uniting with oxygen, the special name of *oxidation* is given to the change. Thus, substances burning in air are said to be *oxidised*.

Oxidation is also responsible for the formation of iron rust. The iron, left in contact with moist air, changes to a red powder, which we call rust, but which is another form of iron oxide.

Certain conditions are always necessary before one substance will unite or combine with another, and these vary considerably for different substances. Magnesium, for example, will not combine with oxygen until it is heated ; iron will not form iron rust unless water is present ; phosphorus requires only a little heat, while copper and lead have to be heated considerably.

The temperature to which a substance must be heated in order to make it take fire, *i.e.* combine with oxygen, is

called the *ignition temperature*. For phosphorus this temperature is very low, but for iron, copper and lead, it is very high. When once the substance is burning, the combination produces heat, and if this is sufficient to keep the temperature above the ignition point, then burning will continue. This explains why it is possible to blow out a candle flame—the breath carries away some of the heat of burning and reduces the temperature below the ignition point. Water, too, prevents burning in the same way by extracting heat.

A Box of Matches

In primitive times, fire was produced by vigorously rubbing together two pieces of wood. By so doing, the savage raised the temperature of the wood above the ignition point and produced fire. Unfortunately, wood has a high ignition temperature and the sticks had to be rubbed for a long time. By the invention of matches, where the tip is coated with a substance of low ignition temperature, the necessity for this continuous rubbing disappears.

Matches were first introduced in 1833. At first, the tips were made of a mixture of phosphorus and sulphur, but these were dangerous as they took fire too easily. Moreover, the phosphorus gave off poisonous fumes which were dangerous to health, causing a terrible disease known as *phossy jaw*.

To-day, another safe form of phosphorus is used—red phosphorus—this does not give off fumes and is not poisonous. When rubbed on a rough surface the friction produces sufficient heat to cause ignition. Yet another form of match is in very common use, one we call the *safety match*, as it will ignite only when rubbed on a

pecially prepared surface. The tips of these matches are made chiefly from the substance potassium chlorate, which you have seen readily gives up its oxygen when heated. The striking surface contains red phosphorus and powdered glass, and when a match is rubbed on it, sufficient heat is released to produce oxygen from the match head. This oxygen makes it easy for the match head to ignite. The stems or sticks are often impregnated with wax to make them burn more readily.

When Substances split up

In Expt. 1 on p. 13, mercury oxide was heated and oxygen was given off. Tiny beads of mercury were formed on the sides of the tube, and so by this experiment the oxide was changed back into mercury,

i.e. mercury oxide \rightarrow mercury + oxygen.

This reverse process, the splitting up of a substance into two or more separate substances, is called *decomposition*, but while it is possible to split mercury oxide by heating it, the same is not true for all the oxides. Some of them are very difficult to split up, but there are other methods besides heating.

Just as oxidation is the special name given to the process of adding on oxygen, so the decomposition of an oxide, either partly or wholly, is given the special name of *reduction*. Later, you will learn that some substances possess great reducing powers, that is, they can reduce oxides to simpler substances with ease.

How we use Oxygen

Oxygen is essential for all forms of animal life, and here lies the greatest of its uses. The human body, like the

steam engine, requires fuel and this is obtained in the form of food. Just as oxygen is necessary for the fire in the furnace of the engine, so it is necessary to enable the food taken into the body to be consumed. The air we breathe in fills the many air sacs in the lungs, and from these the oxygen is extracted and carried by the blood stream to the tiniest tissues of the body. Here it is consumed and the waste products are carried back by the blood to the lungs and breathed out of our bodies.

Just as the fire burns slowly when the air is shut off, so our bodies become tired and we feel drowsy when living in air which is deficient in oxygen.

All hospitals have a ready supply of oxygen, usually contained in stout iron cylinders. This they give to patients who are suffering from certain diseases of the lungs, or to those who are unconscious through some form of gas poisoning. The pure oxygen purifies the blood, driving out the poisonous gases and restoring the circulation.

This use of oxygen for living is by far the greatest, but there are others which, though less important, must be considered.

Many of you will have met a gas called *acetylene*; it is used in many cycle lamps and can be prepared by allowing water to drip on to a substance called calcium carbide (a compound of a metal calcium with carbon). The gas burns with a bluish flame which is very hot. When oxygen is mixed with acetylene the flame produced is extremely hot, about sixteen times as hot as boiling water. This is hot enough to cut iron as easily as scissors will cut paper, and the *oxy-acetylene* flame is used for cutting thick sheets of metal or girders. In addition to

its use for cutting, it can be used for joining metals, since it is hot enough to fuse or melt them, and this allows them to run together and join up. Oxy-acetylene torches produce a very brilliant light in addition to great heat, and the user has to protect his eyes by wearing coloured glasses.

There is another flame which is even hotter, and this, too, is obtained by mixing two gases—oxygen and another called hydrogen, which you will study later. This *oxy-hydrogen* flame has a temperature of about 5000° F., or about twenty-four times that of boiling water.

Another use for oxygen, a use which can be of great destructive power, is in the manufacture of explosives. You have already learnt that when a substance burns it combines with oxygen. Now, if the burning can be made extremely rapid, an explosion will occur, and a large number of quite ordinary substances can be made explosive if the correct quantity of oxygen is mixed with them. Thus petrol vapour, coal gas, alcohol can all be made to explode by the addition of the correct amount of oxygen.

The gas in the burner often goes out with a bang and the petrol vapour and air in the engine of a motor-car explodes when ignited—both due to this mixing with oxygen and very rapid burning.

Not all explosives, however, depend upon the oxygen in the air. Many contain substances rich in oxygen, substances which are prepared to give it up easily. Thus, in gunpowder, a mixture of charcoal, sulphur and saltpetre, the last named supplies the oxygen necessary for very rapid burning.

There are other explosives which do not depend upon

oxygen at all. They are usually classed as *detonating explosives* and only explode when struck.

Now you have met some compounds such as magnesium oxide, zinc oxide, iron oxide—all substances which are stable—and with these the substances of which they are made are very friendly towards each other. These substances are difficult to split up, but there are others which are very easily separated and many of these contain nitrogen, that very lazy gas which you have already met.

Nitrogen does not like joining up with any other substance, and at the slightest opportunity it will break away. Mercury fulminate, made up of mercury, carbon, oxygen and nitrogen, is so unstable that if it is lightly tapped, it breaks up instantly, and the gases released by expanding so rapidly cause an explosion. This substance is one of the most violent explosives, and on being struck, exerts a pressure of over 200 tons per square inch.

Another high explosive is one called nitrogen iodide, a compound of two substances, nitrogen and iodine. This will break up and explode at the lightest touch with a feather.

There are many other explosives in addition to these, and almost all of them contain nitrogen—the substance which does not like combining, and which breaks away from its compounds very readily.

Noteworthy Oxides

Scientists tell us that the world was once a fiery mass, so hot that all the substances were gases. Through many thousands of years these gases cooled and became first liquid and then, still later, combined to form solids.

This great heat made it easy for substances to combine, and the metals took full advantage of the opportunity. Iron joined up with oxygen to form some of our iron ores ; copper and tin did likewise, and these metals are rarely found in the free state, but almost always joined up into compounds. Some metals, however, will not combine easily, the more important being gold and platinum, and these are found in the earth in the free state.

The oxides of some other substances are also very useful to mankind. Lead oxide finds a great use in the making of paint, zinc oxide when mixed with some form of grease, usually Vaseline, is a valuable ointment ; magnesium oxide, too, is used as a medicine ; calcium oxide is quicklime, used extensively by builders.

There are two oxides which are very important, and both are oxides of a substance we call carbon. Carbon itself is a wonderful substance. To begin with, it exists in several different forms. Diamond, anthracite coal, the lead in your pencil, soot and wood ash are all the same substance—carbon—in different disguise. Carbon, too, combines with many other substances—in fact, there are over 250,000 carbon compounds, and a special branch of science called *organic chemistry* is devoted to their study. Carbon is present in oils and fats, in our clothes, in all foods, and even in the tissues of plants and animals.

The most common oxide of carbon is a gas produced when carbon is allowed to burn in a plentiful supply of air or oxygen. To do this, however, means first preparing a jar of oxygen, and it will be more convenient to adopt another method using other substances.

Experiment 4

Place some small lumps of marble (or chalk) in a flask and to it add dilute hydrochloric acid. Collect the gas by upward displacement of air (see Fig. 6).

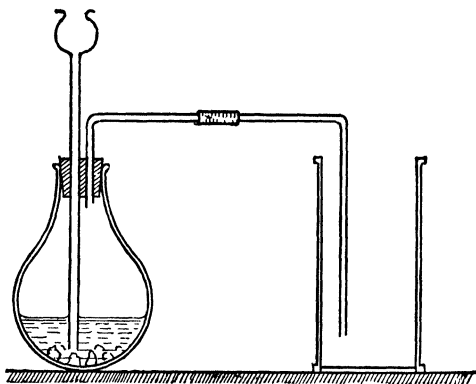


FIG. 6.

Experiment 5

Into one jar of the gas plunge a lighted match.

Use another jar to test its taste and smell.

Arrange the delivery tube to pass through lime water.

Describe what happens to the lime water.

Continue passing the gas for some time.

Heat the lime water and describe what happens.

Experiment 6

Blow through a tube leading into a jar of lime water. Describe what happens.

What does this show you regarding the gases breathed out of the human body?

From your results you have learnt that carbon dioxide is heavier than air, that it will not permit substances to burn in it, and that it produces a milkiess with lime water. This milkiess is due to the combination of the carbon dioxide with the lime to form calcium carbonate and this is insoluble in water.

Carbon dioxide + lime water \rightarrow water + chalk.

The milkiess disappears after a while due to the formation of a new substance—calcium bicarbonate—and this does dissolve in water. The milkiess reappears when the liquid is heated, for then the bicarbonate changes back to the carbonate ;

i.e. water + chalk + carbon dioxide \rightarrow water
+ calcium bicarbonate,

and on heating,

water + calcium bicarbonate \rightarrow water
+ chalk + carbon dioxide.

From Expt. 6 you have seen that carbon dioxide is contained in the gases we breathe out, and if you refer back to page 18, you will be able to understand what happens.

The blood carries with it food partly in the form of carbon and oxygen. When this is consumed or burnt, carbon dioxide is produced, and this is carried back by the blood to the lungs and there given up and breathed out. How is it, then, that the supply of oxygen in the air is not used up and replaced by this carbon dioxide? Nature provides for this by the plants. During the day-light the plants, instead of using oxygen, breathe in

carbon dioxide, extract the carbon to build their bodies and release pure oxygen, thus preserving the balance of oxygen in the air. When it is dark, the reverse process takes place, and oxygen is used by plants as well as by animals. This breathing of plants is done chiefly by the leaves and the splitting up of the carbon dioxide by the green colouring matter in the leaves and stems (called chlorophyll) and the sunlight.

Carbon dioxide, since it will not permit burning, is often used in chemical fire extinguishers (see Fig. 7). By

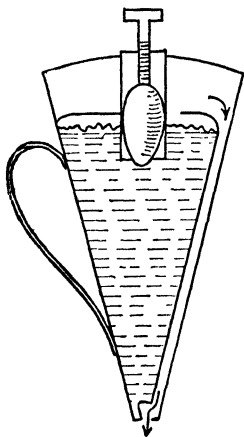


FIG. 7.

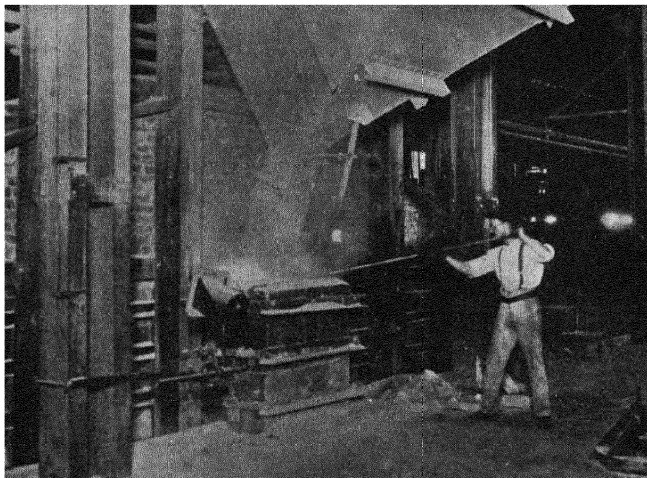
depressing the plunger a glass bottle of hydrochloric acid is broken, and this, when released, reacts with the chalk, producing large quantities of carbon dioxide which issues from the jet as a stream of gas. When directed on the fire, this prevents burning.

Carbon dioxide, too, is used extensively in the preparation of mineral waters. More than will dissolve is forced into the liquid—either water or fruit juice—under pressure and the bottle is then corked. On opening the stopper, the gas escapes and the liquid bubbles.

Another important use for carbon dioxide is as a refrigerant. The gas solidifies at a very low temperature and it is in this solid form that it is sold for freezing and preserving. Solid carbon dioxide has many advantages over ice, the greatest being the fact that when it vaporises no water or liquid is produced, the solid evaporates to

gaseous carbon dioxide which mixes with the surrounding air.

The other oxide of carbon is far more dangerous as it is a deadly poison. It is formed when carbon is burnt in a limited supply of air, and is called *carbon monoxide*. This



CHARGING A FURNACE WITH "BARCH"

(Courtesy, United Glass Bottle Mfg. Co)

gas is contained in the exhaust fumes from motor cars and in coal gas. It burns with a bluish flame and changes to carbon dioxide,

i.e. carbon monoxide + oxygen \rightarrow carbon dioxide,

and this will show you that this oxide does not contain as much oxygen as the more common carbon dioxide.

At least one other oxide must be mentioned before leaving this chapter, and this is the oxide of a substance called silicon. Silicon oxide will be more easily recognised under the name of *sand*, although the chemical name for it is *silica*. Its great use is in the manufacture of glass, an industry which is at least 4000 years old. The story goes that some Phoenician sailors made their camp fire on a sandy beach, and used lumps of crude soda to support their cooking vessels. After the fire had died down they found rough lumps of glass formed by the melting together of the sea sand and the soda. The materials in use to-day are still the same, although science has enabled the glass manufacturer to make his glass of all colours and shapes by the addition of other substances, and by the use of machinery.

The materials in most common use to-day are silica sand, soda and limestone, and these, together with some rough crude glass, are heated to a great temperature in tanks, lined with clay and silica. These tanks are divided into two compartments by a wall with a hole in it. Through this hole passes the molten glass, and this is drawn off and blown by compressed air machines into bottles and vases of various sizes and shapes, or rolled into sheets.

In the next chapter another oxide will be encountered—water—but, before entering on the study of this, it will be advisable to pause and note down all the important facts that have been presented in this chapter.

Remember that, so far, the only compounds studied have been oxides, but there are very many more formed by the combination of other substances, and some of these will be met as the work proceeds.

SUMMARY OF THE CHAPTER

Things to Note

1. The methods of making oxygen.
2. How oxides are formed.
3. The meaning of ignition temperature.
4. The meaning of the terms: combination, decomposition, compounds, oxidation and reduction.
5. The necessity for nitrogen in explosives.
6. The many uses for oxygen.
7. The oxides of carbon and silicon and their uses.

CHAPTER 3

WATER AND HYDROGEN

How Water can be split up

WATER, like air, is a very common substance ; it exists almost everywhere—in our bodies, in all plants and foods, in the earth and even in the air. Here we are going to

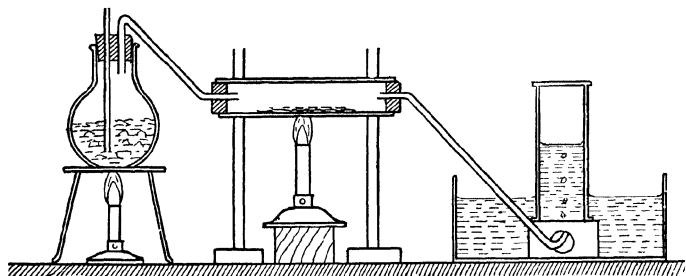


FIG. 8

discover what it is made of, that is, what substances combine to form water.

Experiment 1. (Demonstration)

Fit up the apparatus as shown in Fig. 8. *A* is a hard glass tube containing magnesium wire or ribbon, *B* is a flask containing water for the production of steam, and *C* is a gas jar for the collection of gas.

Remove the gas jar and pass steam through the tube until all the air is expelled—until bubbles cease to come

from the delivery tube. Then place the jar in position and heat the magnesium wire.

Observe the changes which take place in the form of the magnesium and note whether any gas is produced.

Experiment 2. (Demonstration)

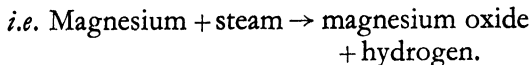
Take a full gas jar and test the gas to see if it will burn. (Stand clear of the jar when you do this.)

Experiment 3

Remove the powder from the tube when it is cool and examine it. Compare it with the powder produced when magnesium burns in air.

You will have noticed that the magnesium glows brightly in the steam and changes to a white powder very similar in appearance to the white powder produced when magnesium is burnt in air. In fact, the powders are exactly the same—*magnesia* or magnesium oxide. The gas issuing from the tube is inflammable and burns with a blue flame. This gas Priestley called *inflammable air*, but to-day we know it as *hydrogen*.

Now what has happened in the tube? The steam has been changed into hydrogen, while the magnesium has changed to magnesium oxide.



As there was no air in the tube, both the hydrogen and the oxygen must have been provided from the steam. Thus, water (or steam) must contain both hydrogen and oxygen, and later you will be able to show that this is all it does contain.

Water is a compound of these two gases and this will serve to show you the great difference between the properties or characteristics of a compound, and those of the substances which make it. Hydrogen is inflammable, oxygen makes things burn fiercely, but water, a compound of both, extinguishes fire. Thus water, or any compound, is an entirely different substance and has none of the

characteristics of the substances which combine to make it.

Here are two more experiments which show how water can be split up.

Experiment 4

Fill the trough of the electrolysis apparatus with water and add one or two drops of sulphuric acid. Connect the terminals to a battery of accumulators and adjust the resistance until bubbles of gas begin to form in each tube (Fig. 9).

(a) Is the rate of production of gas the same in each tube?

(b) From which pole does most gas appear?

Test the gas which is formed at the positive pole to see if it will relight a glowing splint.

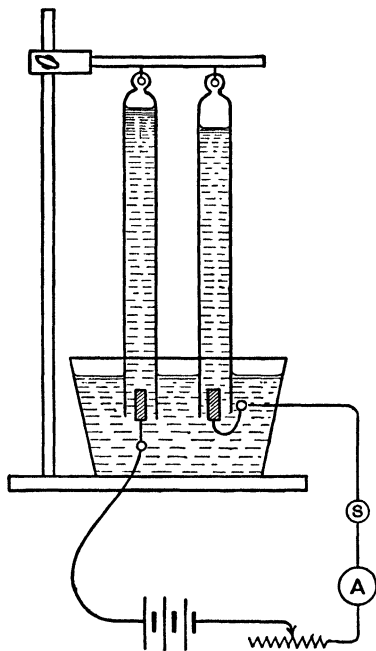


FIG. 9.

Test the gas formed at the negative pole to see if it will burn.

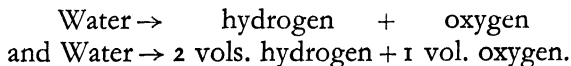
Experiment 5

Take a small piece of sodium metal (do not handle with your fingers) and place in water contained in a trough.

Observe the action between the sodium and the water. Prevent the movement of the sodium by holding it against the side of the trough with a piece of stout wire.

Describe what happens when sodium is placed in water.

In Expt. 4, water containing a tiny quantity of sulphuric acid has been split up by electricity into two gases, one of which burns, while the other relights a glowing splint. In other words, the water has been decomposed into hydrogen and oxygen. You will have noticed that the volume of hydrogen produced is twice that of the oxygen, and to represent these changes by equations, we must write :



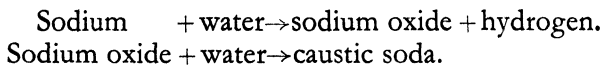
You will be wondering why sulphuric acid has not been included in these equations, and you might imagine that the gases have been partly produced from this substance. However, it has been found that the quantity of sulphuric acid present never changes, it is the same after the experiment as before, and cannot have been decomposed by the electric current.

The reverse of this experiment was performed as early as 1781 by an English chemist named Cavendish. He

mixed two volumes of hydrogen with one of oxygen in a stout glass globe and exploded them by electricity. Water was obtained and both gases were completely used up. By experimenting with other proportions, say four volumes of hydrogen with one of oxygen, he proved that while water was still produced, two volumes of hydrogen were left over, unconsumed.

Sodium, used in Expt. 5, is a metal which has a great liking for air and water, and has to be stored in oil. When placed in water, it jumps about and produces a fizzing noise and sometimes takes fire. Hydrogen gas is produced and may be collected by holding a test-tube over the sodium.

Here is yet another example of the way in which substances combine. The sodium, owing to its great liking or *affinity* for oxygen, is able to extract it from the water, and in so doing releases hydrogen. The sodium oxide formed also combines with the water, forming a solution of a new substance or compound, called *caustic soda*. You can show that some substance is present in the water by evaporating it over a bunsen burner. Caustic soda is a very important chemical, and in Chapter 9 you will study its properties and uses in greater detail, but for the present we will leave it with the equations for the action :



Hydrogen

Several methods for the preparation of hydrogen have already been given, but these are not very convenient if it is wanted in any quantity.

Hydrogen was discovered in the sixteenth century by a scientist, Paracelsus, who found that a gas was produced when iron was allowed to combine with sulphuric acid. The gas made in this way is not very pure, since the iron usually contains other substances, and in the science room it is more satisfactory to use zinc.

Experiment 6. (Demonstration)

Fit up the apparatus as shown in Fig. 10, placing a few pieces of scrap zinc in the bottom of the flask.

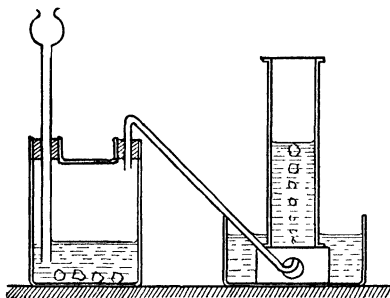


FIG. 10.

Pour a little dilute sulphuric acid into the flask through the funnel.

Allow the first bubbles of gas to go free and then collect two or three tubes or jars of the gas for use in the next experiment.

Experiment 7. (Demonstration)

Take the first tube of gas and notice its colour.

Smell the gas.

Hold the second tube of gas upside down and hold a lighted match at the end. Notice the flame.

Open the third jar and allow it to stand on the bench with the mouth upwards. After ten minutes test the gas in the jar to see if it will burn.

Make a list of the properties of hydrogen from your results.

Experiment 8

Fit up the apparatus as shown in Fig. 11. The hydrogen flask is fitted with a delivery tube ending in a jet, and above this jet is placed a flask of cold water.

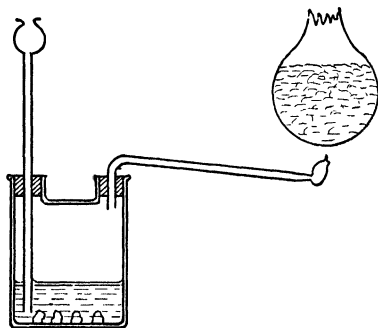


FIG. 11.

Allow the gas to escape for a few minutes and then light the jet and use the cold water flask to condense the products of combustion.

Notice what is formed on the bottom of this cold flask. Taste the liquid formed.

The results of your experiments have shown that hydrogen gas has neither colour nor smell, that it burns with a blue flame, and that it is lighter than air. When it burns in air, it combines with oxygen to form water and, in Experiment 8, this was condensed on a flask of cold water.

Owing to its lightness it was, until recently, used in airships and balloons. It has a considerable lifting power, being $14\frac{1}{2}$ times lighter than air, but its inflammable nature makes it dangerous. Its place has been taken by another very light gas called *helium*. This is much safer, as helium is classed as one of the *inert* gases, that is, one of those substances which do not combine with others. Thus helium will not burn and its use in aircraft makes them safe from fire and explosion.

The coal gas supplied to your houses by the gas company contains a large proportion of hydrogen. With it is carbon monoxide, mentioned in the last chapter, and the following table shows the general composition of domestic coal gas.

Substance.	Per cent. present.
Hydrogen - - - -	47%
Carbon Monoxide - -	14%
Oxygen - - - -	1%
Other compounds of carbon, hydrogen and oxygen -	38%

SUMMARY OF THE CHAPTER

Things to Note

1. How water can be made and the substances contained in it.
2. How water can be split up.
3. How hydrogen is made in large quantities.
4. The characteristics of hydrogen.

CHAPTER 4

MORE ABOUT WATER AND OTHER LIQUIDS

When Substances dissolve

THE term *dissolve* is in very common use ; we say that sugar dissolves in our tea or coffee, making it sweet. The sugar seems to disappear, but the fact that our tea tastes sweet proves that it has not been lost, but is actually present in the tea although invisible.

A substance which disappears from view when placed in a liquid is said to *dissolve*, and the new liquid formed is called a *solution*.

Experiment 1

Take a test-tube and partly fill it with water. Add a small quantity of salt and shake the tube for a second or two.

Taste the liquid.

Here you have dissolved some salt in water, and by tasting the liquid proved that the salt is still present. Actually the salt has broken up into millions of very tiny parts, and each tiny drop of water has one or more of these salt particles attached to it. Not all substances dissolve so readily as salt and sugar ; in fact, some do not dissolve at all and are said to be *insoluble*, while others dissolve in such small quantities that they appear to be insoluble.

Experiment 2

Partly fill a number of test-tubes with water and then test the following substances to find if they are soluble. Use only very small quantities, not more than could be placed on a threepenny piece: soda, sand, saltpetre, salammoniac, chalk, sulphur, mercury.

Substances which dissolve.	Substances which do not dissolve.

Experiment 3

Take two tins and into each pour the same amount of water. At the same time, drop a lump of sugar into each tin and place one to one side. Stir the liquid in the other tin until the sugar dissolves, then inspect the other tin.

Complete this statement: "Substances dissolve more readily if the liquid is —."

Experiment 4

Place equal quantities of water in each of two tins. Boil the water in one of the tins and then, at the same instant, drop a lump of sugar into each.

- (a) Which lump is the first to dissolve?
- (b) Complete this statement: "Substances dissolve — in hot water than in cold."

Experiment 5

Take two tins each containing an equal quantity of water. Powder one lump of sugar in a pestle and mortar and, at the same instant, drop a lump into one tin and the powder into the other.

Observe which is the first to dissolve.

In these experiments you have seen that sugar dissolves more readily in warm water than in cold, and that solution is helped by stirring (or shaking) and by powdering. The same is true for most substances, although there are one or two which are more soluble in cold water than in hot.

Some common Solvents

Water allows a greater number of substances to dissolve in it than any other liquid. Not only solids will dissolve, but liquids and gases as well. Thus alcohol, carbon dioxide, ammonia, with very many others, will dissolve in water. The solubility of solids and liquids, with one or two exceptions, increases with rise in temperature, but for gases the reverse is true. The solubility of gases decreases as the temperature is raised, and generally the gas is forced out of solution on boiling.

There are, of course, other solvents besides water. Thus mercury will dissolve gold, silver, tin, copper, zinc and potassium, and the solutions produced are given the special name of *amalgams*. These amalgams are used to a great extent: thus tin amalgam is used for backing mirrors; zinc amalgam is used on the plates in some electric cells to prevent the wasting away of the zinc

when the battery is left standing. Mercury, too, is still used a great deal in the extraction of gold and silver from the ores. The gold ores are first crushed to powder by stamping mills and then passed over plates covered with mercury. The gold is dissolved and later separated by boiling the amalgam to drive off the mercury. A more modern process employs a dilute water solution of a substance called potassium cyanide. When the crushed ore is washed by this liquid in contact with air, the gold dissolves and is separated from the solution, either by the addition of zinc, which replaces the gold, or by the use of an electric current.

More common solvents include petrol which will dissolve grease, methylated spirit which dissolves varnish, and turpentine which dissolves tar, paint and sulphur. Benzine, like petrol, dissolves grease and is frequently used for cleaning clothes. More examples of solution will occur as you work through this book: make a note of all that you meet.

More about Solutions—Crystals

Experiment 6

Take a test-tube and partly fill it with water.

Add a little powdered alum and shake until it dissolves.

Repeat using more alum and continue until there is a little left at the bottom of the tube which will not dissolve after repeated shaking.

- (a) Is the alum left in the tube insoluble in water?
- (b) How could you prove your answer to be correct?

Experiment 7

Using the same solution as in Experiment 6, warm the tube and add alum until no more will dissolve and there is a little still left at the bottom. Carefully pour off the liquid into another tube and allow it to cool.

(a) What appears at the bottom of the tube when it is cold?

If you have followed the instructions carefully, you will find that when the hot solution is cooled a small quantity of alum appears at the bottom of the tube. There is a limit to the amount of a solid which will dissolve in water, and, since alum is more soluble in hot water than in cold, the surplus is rejected and reappears as solid. The cold solution obtained in this way contains as much substance as it can hold and is said to be *saturated*.

Experiment 8

Prepare a strong solution of alum by dissolving the powder in cold water. Shake it well until no more will dissolve. Pour off the solution into a shallow basin and set it aside to evaporate.

(a) After a few days examine the basin and describe what has happened.

(b) How does the solid produced by evaporation differ from that originally dissolved in the water?

N.B.—Solids produced in this way are called *crystals*.

Experiment 9

Repeat Experiment 8, using solutions of: sal-ammoniac, copper sulphate, and salt.

Examine the shape of the crystals formed.

Experiment 10

Select a good crystal of copper sulphate and let it hang by a piece of cotton in a solution of copper sulphate.

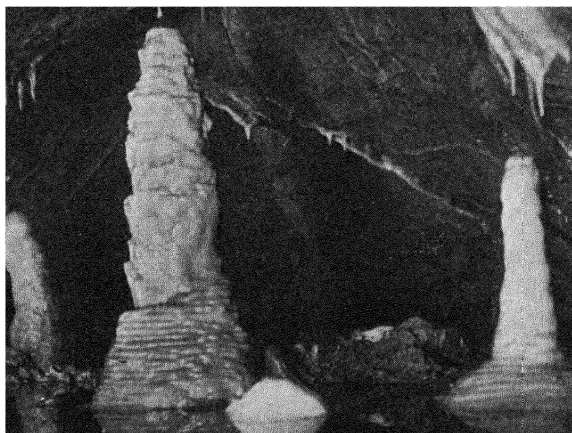
Examine its size at intervals.

Among other facts, your experiments have shown that when a solution is allowed to evaporate, the solid substance is left behind in the form of crystals. Exactly how these are formed no one knows, but there is some force present in the solution which has the power of joining up the millions of tiny particles of substance into these regular shapes. Moreover, crystals of any substance prepared in the same way are always of the same shape, although they may differ in size.

Experiment 10 provides an illustration of the way in which the size of crystals may be increased. Nature provides a more striking example in the formation of those long rocky pillars, somewhat like icicles, which are to be found suspended from the roofs or standing on the floors of many caves. Those hanging from the roof are called *stalactites*, and those standing on the floor, *stalagmites*, and striking examples of these are to be found in the caves at Cheddar, a small town near Bristol. Their formation depends upon the evaporation of the water from the solution of calcium bicarbonate which drips from the roof. When the dripping is very slow each drop leaves behind a tiny deposit of limestone, and through countless years the stalactite is built. Should some of the solution drip on to the floor of the cave and evaporate, a stalagmite is gradually built up.

The formation of crystals from solutions is not the only method. Wire, when bent sharply to and fro, becomes

brittle and finally breaks. The wire has become crystalline in form, and metals subject to continual vibration often behave in this way. The frame of a bicycle or the chassis of a motor-car, being constantly vibrated by the uneven nature of the road, sometimes snaps for this reason.



STALAGMITES IN CHEDDAR CAVES
(Courtesy, Gough Caves)

Again, some substances can form crystals without being first dissolved in a liquid. Thus salammoniac crystals can be obtained by heating the solid in a tube. It will be seen to turn to vapour and change again to crystals on the side of the tube. Solid iodine behaves in a similar manner, and this process is given the special name of *sublimation* and the substances are said to be *sublime*.

Many crystals contain water, as the following experiments will show.

Experiment 11

Heat some crystals of washing soda in a test-tube.

Hold the open end of the tube near the bottom of a flask containing cold water.

- (a) What changes take place in the form of the soda ?
- (b) What condenses on the bottom of the flask ?

Experiment 12

Repeat Experiment 11, using blue copper sulphate.

- (a) What changes take place in the copper sulphate ?

Add one or two drops of water to the solid left in the tube.

- (b) Describe the changes which take place.

Other crystals, among them Epsom-salt, alum and gypsum, also contain water, while salt, saltpetre, sal-ammoniac and iodine do not. The water joined up with the substance in crystals is known as *water of crystallisation*, and it is the loss of this that makes the soda turn to powder when heated or when left in contact with air. In copper sulphate the water of crystallisation causes the blue colour and, when it is driven off, the colour disappears. If this white copper sulphate is left in a damp place, or if a drop of water is added, the blue coloration returns, and here is a very simple test for water, for no other liquid will cause such a change.

Natural and Artificial Waters

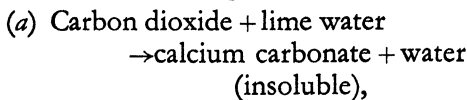
The water found in the earth, in springs, rivers, lakes and the sea, together with rain, is called *natural water*. Tap water or the water prepared by burning hydrogen is *artificial water*, as it has been either purified or made.

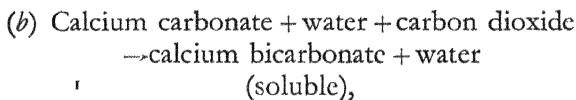
Spring and rain water are the purest forms of natural water and, although rain collects particles of soot and dust from the air and in towns may dissolve some of the gases from factory chimneys, it is usually fit for drinking purposes. The purity of river and lake water depends upon the district in which they are situated. Rivers, particularly, often contain dissolved substances, as they often flow through limestone and chalk. In the vicinity of large towns, rivers are often used to carry sewerage and refuse to the sea and the water becomes contaminated. Many rivers flow into the sea and this water will contain a large proportion of dissolved solids. The Irish Sea, for example, has been shown to contain 40 grammes of solid to 1000 grammes of water and of this, 25 grammes is dissolved salt. Some inland seas, the Dead Sea for instance, contain a much greater proportion of solids in solution.

Hard and Soft Waters

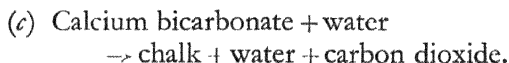
In many districts the water is said to be *soft*, meaning that it easily forms a lather with soap, while in other districts the water is *hard* and a lather is difficult to obtain.

Now on page 23 you have seen that carbon dioxide produces a milkiness with lime water. The solid produced is actually calcium carbonate (chalk). After passing the gas for some minutes, the milkiness disappeared, due to the formation of another compound—calcium bicarbonate—which dissolves in water. On heating, the bicarbonate is changed back to the insoluble carbonate and carbon dioxide is given off; *i.e.*,





and on heating,



Tap water containing only bicarbonate in solution is said to have *temporary* hardness. It can be made soft by boiling, when the chalk is thrown out of the liquid and settles on the sides of kettles, boilers, etc., as a hard greyish *fur*.

Yet another way of softening such water is by the addition of lime. This combines with the additional carbon dioxide in the bicarbonate (equation *b*) and forms chalk (equation *a*).

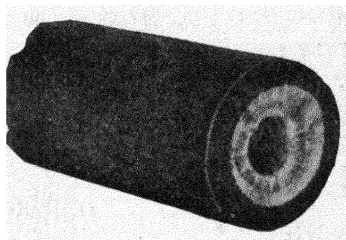


FIG. 12.

Rock-hard deposit like this, in boiler pipes and kettles means loss of heat and waste of fuel.

(Courtesy, Electrolux Ltd)

Water containing dissolved gypsum (calcium sulphate) is said to be *permanently* hard. Gypsum is more soluble in cold water than in hot, but cannot be expelled by boiling. With soap the gypsum forms a curd and, until all the gypsum has been absorbed in this way, no lather can be obtained. When permanently hard water is treated with soda (sodium carbonate), chalk is thrown out of the solution and forms a fur, while soluble sodium sulphate is left in the water. As only very small quantities are dealt with, the resulting water does not taste of

the dissolved sodium sulphate (Glauber's salt). The reaction in equation form is :

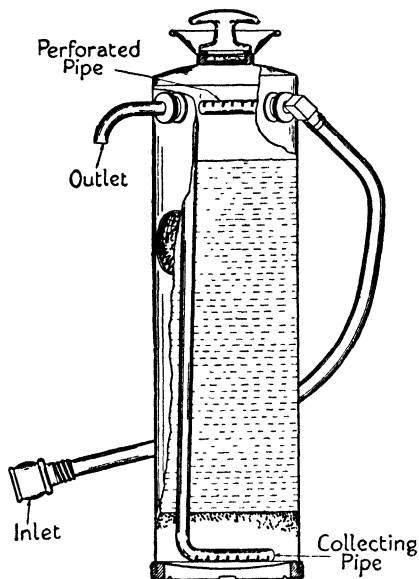
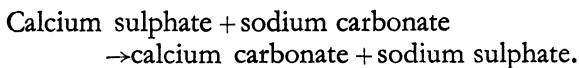


FIG. 13.—A WATER SOFTENER
(By courtesy of Electrolux Ltd).

Many households employ water softeners to remove the hardness from their water, and in Fig. 13 such a softener is shown.

Our Water Supply

In Book I you have seen how we obtain our water supply, and here we shall consider how the water we

drink is made pure. All waterworks employ large filter beds (Book I, p. 29), and through these the impure water from the storage tanks or reservoirs is passed. Filtration is found to extract many of the impurities of reservoir water and is the chief method of purification, but as you have learnt, it does not remove dissolved substances nor does it destroy germs or bacteria. In many towns the water is treated with a gas, *chlorine*, which has the power of destroying bacteria. Peculiarly enough, chlorine is a deadly poison and was employed as one of the poisonous gases during the Great War. However, very tiny quantities of this gas have the power of destroying bacteria without making our water supply poisonous.

Chemists often require absolutely pure water for their tests and experiments, and this they prepare by a process known as *distillation*.

Experiment 13

Set up the apparatus shown in Fig. 14. The flask is partly filled with tap water, while the receiver is cooled by immersing it in cold water and allowing a stream of water to run over it.

Heat the water in the retort until it boils and then regulate the flame until the water continues to boil steadily.

After a while water will be seen to collect in the receiver and this may be used for the following experiments.

Experiment 14

Take a small quantity of the *distilled* water and taste it. Compare its taste with that of ordinary tap water.

Experiment 15

Take equal quantities of tap water and distilled water in test-tubes, and to each add the same amount of a soap solution.

Shake each tube and decide which water is the first to form a lather.

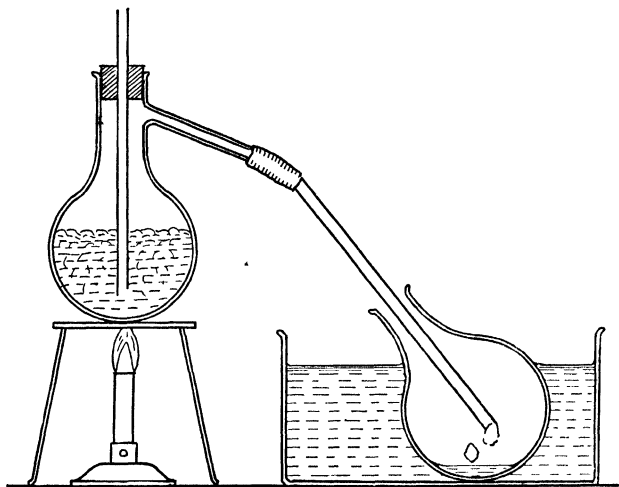


FIG. 14.

Your apparatus is suitable only for the preparation of distilled water in small quantities and, when a regular supply is needed, larger apparatus has to be employed. Distilled water is always used in accumulators, since the presence of dissolved substances would cause the formation of a *fur* on the plates, and in time this would ruin the battery.

The process of distillation is used not only for the puri-

fication of water. Chemists use it repeatedly for purifying other liquids containing dissolved substances. Suppose we have a mixture of two liquids, A and B with boiling points of 60° C. and 100° C. By heating the mixture and by keeping the temperature at 60° , liquid A will vaporise and distil, while liquid B will remain in the retort. Thus it is possible to separate the liquids, and this method is of great importance in the extraction of the by-products of coal gas and crude oil. This process for the separation of liquids from mixtures is given the name of *fractional distillation* and will be met again in later chapters.

SUMMARY OF THE CHAPTER

Things to Note

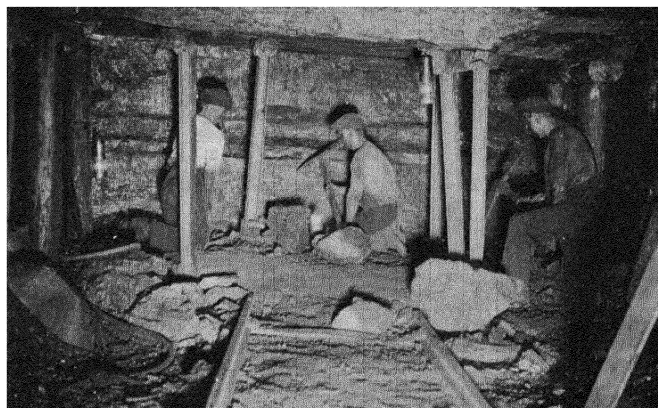
1. The meaning of the terms: solution, dissolve, solvent, amalgam.
2. Solvents in common use.
3. How crystals can be formed.
4. The causes of hardness in water.
5. How the hardness in water can be removed.
6. The processes of filtration and distillation.

CHAPTER 5

COAL, COAL GAS AND PETROLEUM

How Coal is obtained

COAL is one of the most useful and valuable substances obtained from the earth. It warms and lights many of



COAL MINERS AT WORK
(By courtesy of "The Daily Herald")

our homes, cooks our food, drives our railway engines and steamships, and helps to provide the power for working the machines in many factories and workshops.

Millions of years ago coal was alive as some form of vegetation—ferns, grasses or trees. At this time, known

as the *Carboniferous* period in the earth's history, the earth's crust was sinking and in the hollows, marshes and swamps were formed. Here a luxuriant vegetation soon sprang up, similar in some ways to that found in the tropical jungles to-day. When the trees and undergrowth died, they sank into the marshes and became covered by layers of sand and gravel from the sea and from rivers. Other layers of vegetation followed and

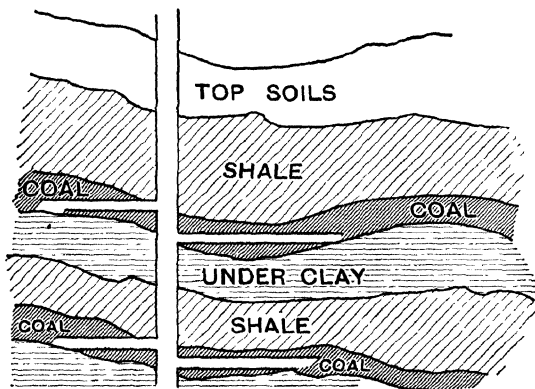


FIG. 15.

gradually disappeared into the marshes, sinking lower and lower as the earth's crust sank. During thousands of years this submerged vegetation decayed in the absence of air and was compressed by the weight above and eventually became what we call coal.

In many mines the coal is found in layers or seams, sometimes only a few feet in thickness, separated one from the other by layers of sand or *shale* (Fig. 15).

Coal was in use in England as early as 1400, but it was not until the eighteenth century that it became important.

With the increasing use of machinery at this time, coal became a necessity and coal mining became one of our great industries.

In the early days coal mining was a very dangerous occupation. Many disastrous explosions occurred, causing great loss of life, and it was the invention of the miner's safety lamp (Book I, p. 75) by Sir Humphry Davy that made the work of the coal miner less hazardous. Even to-day mine disasters are not uncommon, and it will be as well to consider why this should be.

First consider what coal is made of. There are various types, of course, but all contain a very large proportion of *carbon*. The following table will show you how the various types differ :

Type.	Per cent. carbon.	Hydrogen.	Oxygen.	Sulphur.	Ash.	Nitrogen.
Northumberland (soft)	79	6	10	1.5	1.5	1.7
South Wales (soft)	83	6	6	.8	2.5	1.7
South Wales (anthracite)	92.5	3	2.5	—	2	—

In coal mines a gas called *methane*, a compound of carbon and hydrogen, is very common. The miners call it *fire damp* and it is also known as *marsh gas*. Methane is produced when vegetable matter decays and is often found near marshes. It is inflammable, and when mixed with oxygen and ignited becomes explosive, forming carbon dioxide and water.

i.e. Methane + oxygen → carbon dioxide + water.

With the naked flame used by miners in the early days of the industry, explosions were very common. Moreover, methane is poisonous and, as it is both colourless and odourless, it often overcame miners before they were aware of its presence. To-day, with the use of electricity and gas masks and the invention of apparatus to detect the presence of fire damp, our coal mines are far safer, and although accidents still occur, they are far less frequent.

More about Carbon

It has already been mentioned that carbon exists in many different forms. In the earth it is found in a natural state as *graphite*, as *coal*, and again as *diamond*.

Graphite is soft and scaly in appearance, and one of its main uses is in the manufacture of pencils. At one time graphite was thought to contain lead, and this gave rise to the term *lead pencils*. Another use for this important form of carbon is as a lubricant, as it gives a very smooth surface and requires a high temperature to melt it. Probably the most valuable use of graphite is in the manufacture of high temperature crucibles and retorts. A mixture of clay, sand and graphite is first moulded to the required shapes and then baked at red heat. These vessels will withstand very high temperatures without melting.

Diamonds are crystalline carbon, and the fact that they are rare makes them valuable. They are found in the form of roughly rounded pebbles, and are cut to various shapes. By cutting, the many faces reflect the light and cause them to sparkle. Diamond is the hardest substance known and is insoluble in all liquids. When pure it is

colourless, but the presence of tiny quantities of other substances gives it a tinge of colour.

Artificial carbon, too, takes many forms. There is *lampblack*, one of the most pure, containing over 98° carbon. It is prepared on a large scale by burning substances rich in carbon—petroleum or turpentine—in a limited supply of air. The smoke from these burning substances is passed into rooms hung with rough blankets, and to these the soot or lampblack adheres. Its chief uses are for the manufacture of boot and stove polish, for printer's ink, and as a pigment for paint.

Charcoal, prepared by burning wood in a limited supply of air, is important because of its ability to absorb gases and liquids. It is found that 1 c.c. of charcoal can absorb as much as 150 c.cs. of ammonia, and for this reason it is used in the construction of gas masks and at the sewerage farm for the absorption of unpleasant gases. Charcoal is always present in the filtering plant at the waterworks, as it absorbs many of the unwanted substances present in the reservoir water.

Two other forms of carbon are produced at the gas works. Coke is left behind in the retorts, while a hard scaly substance called *gas carbon* gradually forms around the sides. Coke is about 90 per cent. carbon and forms a very useful fuel. In addition, it is extensively used in the preparation of iron and steel (see Chapter 6). Gas carbon is very hard and is almost pure. Its greatest use is for the construction of the carbon rods or plates in electric batteries.

Carbon Compounds

You have realised the great value of carbon when you learnt that it combines with other substances to form over

a quarter of a million compounds. To deal with all these is, of course, out of the question, but there are some which are of great importance and in such common use, and these only will be studied. You have met the oxides in Chapter 2, and Methane or marsh gas on p. 52, and now you will pass on to *coal gas*. In the chapters which follow, still more carbon compounds will be met, and if you reserve a page of your note-book for them, you will realise how very varied and useful they are.

Coal Gas

One of the chief uses of coal is for the manufacture of coal gas. A visit to a gas works will prove most interesting and instructive, and those of you who have the opportunity should take advantage of it.

Now the coal we burn in our fires combines with the oxygen in the air, forming carbon dioxide, and this with the smoke vanishes up the chimney. In the smoke are very many valuable materials and the coal fire, while it is very pleasant and gives out a great deal of heat, is very wasteful. At the gas works coal is not burnt, but roasted in the absence of air in large retorts, and the gases produced are separated and stored.

Experiment 1

Obtain a clay pipe and pack it with coal dust. Seal the bowl with clay and support the pipe so that the bowl can be heated (Fig. 16).

- (a) Observe the gas which comes from the stem.
- (b) Test the gas for smell.
- (c) Test the gas to see if it will burn.

The coal gas, when it issues from the retorts, is very hot and the first stage in the purification is cooling. It is passed through a series of long pipes which are sometimes water-cooled. Here, some of the tar contained in the gas is condensed and this is run off into separate con-



GAS RETORTS AT A MODERN GAS WORKS

(By permission of the Gas Light and Coke Co)

tainers. From the *condenser*, as it is called, the cool gas passes through water to remove more tar and then it enters what are called the *scrubbers*, where it is washed free of still more tar and ammonia. There is still sulphur to be removed, and to do this the gas issuing from the scrubbers is passed over quicklime. The sulphur combines with the lime, but the coal gas

passes on into the familiar *gas holders* where it is stored for our use.

Coal gas, as supplied to our homes, consists chiefly of *hydrogen* and *methane*, but there are other substances present as the following table will show you. This table is an extension of that given in Chapter 3 and shows the composition of an average sample of domestic gas.

Substance	Substances combined in it.	Per cent. in coal gas.
Hydrogen	Hydrogen	47
Methane	Carbon and hydrogen	24
Carbon monoxide	Carbon and oxygen	14
Carbon dioxide	Carbon and oxygen	2
Oxygen	Oxygen	1
Nitrogen	Nitrogen	9
Other gases	Carbon and hydrogen	3

Now, in order to burn this gas, oxygen is necessary and it is important to remember that at least two of the gases present—methane and carbon monoxide—are very poisonous. Thus, if there is insufficient oxygen for complete combustion, some of these will remain unburnt and be present in the air as a great danger to life. For this reason it is very important that all gas appliances should have a sufficient supply of air and some means of ventilation should be provided.

Experiment has shown that 1 cubic foot of coal gas requires .9 cubic feet of oxygen in order to burn it completely, and on burning, only carbon dioxide, water vapour and nitrogen are produced. None of these is

poisonous and hence it is the *unburnt gas* of which we must beware, and not the fumes produced by burning.

Gas for Heating

In the early days, coal gas was burnt at the end of a metal pipe, but, in 1855, Robert Bunsen invented a gas burner which gave a very hot flame. In Fig. 17 you will see the various parts of a modern *bunsen burner*. The gas from the pipe issues from a tiny hole and passes up a straight tube. At the bottom of this tube are two (or sometimes only one) holes which may be opened or closed by a movable metal ring.

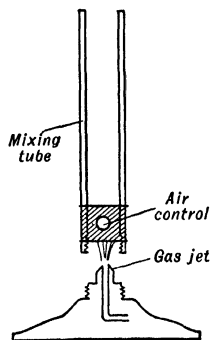


FIG. 17.

Experiment 2

Unscrew the tube of a bunsen burner and light the gas issuing from the jet.

Note the force with which the gas issues from the jet and the type of flame produced.

Experiment 3

Replace the tube in position and close the air holes. Light the gas and notice the type of flame produced.

Gradually open the holes and observe the changes which take place in the flame, and make a drawing of it when the holes are fully open.

You have seen that by burning gas with the air holes closed, a wavering, luminous flame is obtained which is

yellowish in colour. There is very little noise, but, as the air holes are opened and air is drawn into the tube and

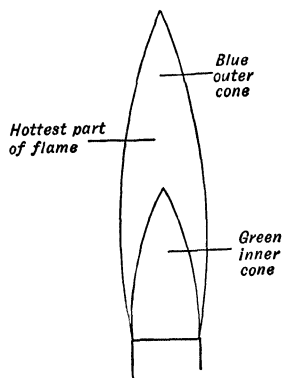


FIG. 18.

mixed with the gas, the character of the flame changes. It becomes noisier and takes a more definite shape. When the holes are fully opened the flame is almost invisible and consists of two distinct parts as shown in Fig. 18. This flame, used on our gas rings and gas stoves, is extremely hot, the greatest temperature being obtained just above the centre cone.

Gas Lighting

When London was first lit by gas the burner was a narrow tube and a flickering yellow flame was obtained. To-day we should consider it but a poor substitute for electric or gas light, but in those early days it was considered marvellous.

The first improvement came with the gas jet, whereby the flame was spread into the shape of a fan by the use of a tube of the shape illustrated in Fig. 19, or by a small burner (Fig. 20).

In 1885, an Austrian—Dr. Carl Welsbach—invented the gas mantle,

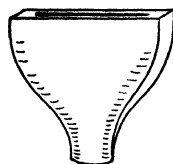


FIG. 19.

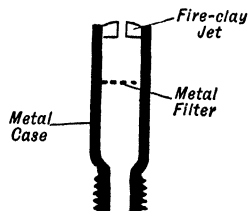


FIG. 20.

which he heated by a bunsen flame. His mantle was woven of a special type of grass—*china grass*—grown chiefly in India and saturated with a mixture of two substances, *thorium* and *cerium*. He found that when burnt a skeleton mantle was left behind, and this, although very fragile, gave a brilliant light when heated. Silk thread now replaces the china grass and modern gas mantles are much stronger and not so likely to break when touched.

Experiment 4

Set up a gas mantle as in Fig. 21, so that a bunsen burner can be placed under it. Light the burner and adjust it for full air and then insert it under the mantle. Gradually reduce the air supply and notice the effect on the light obtained.

The By-products of Coal Gas

You have learnt that during the purification of coal gas, coal tar, ammonia and sulphur compounds are extracted. These, with the coke and carbon left in the retorts, are called the *by-products* and are very valuable. Of these, coal tar seems to be the most important, and the chemist is able to use it to make very many useful materials.

Coal tar itself is a black, sticky substance with a distinctive smell. It is made up of about twelve separate

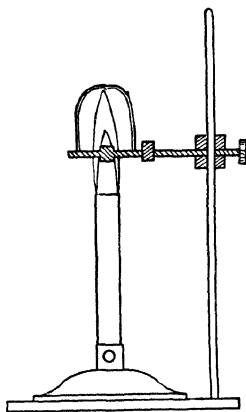


FIG. 21.

substances, most of which can be made to form valuable compounds.

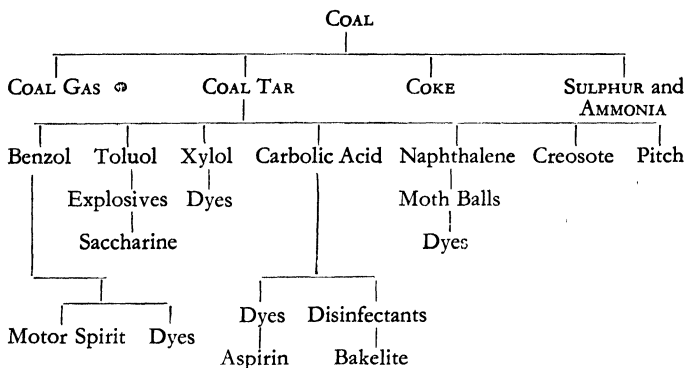
The crude coal tar is placed in large retorts and distilled (see page 49) at a low temperature. Those substances with low boiling points—benzene (or benzol), toluol and xylol (xylene)—distil first and are collected. When separated, these oils can be used to make such substances as *aniline*, which is used for dyes and explosives, and *saccharine*, a substance much sweeter than sugar.

Benzol, too, when blended with petrol, forms a good fuel for motor-cars.

By raising the temperature of distillation, *carbolic acid*, sometimes called *phenol*, is produced. This is used as a disinfectant and is also employed in the manufacture of dyes. Its greatest use, however, is in the manufacture of *bakelite*, named after its discoverer—Professor Baekland. Carbolic acid, mixed with another substance—*formaldehyde*—produces a plastic material which can be moulded to any shape, and when baked becomes very hard. Bakelite is not affected by water, heat or electricity, and is used extensively for electric switches, lamp holders, pocket torches, wireless cabinets, decorative bowls and dishes.

At a still higher temperature, *naphthalene*, used in moth balls, and *creosote*, used for preserving timber and staining floors, is distilled, while *pitch* is left behind in the retort. Even this is useful for paving our roads and for water-proofing roofs.

The great value of coal can be seen from the following summary, and the number of its by-products is steadily increasing.



Petroleum, Petrol and Asphalt

The term *oil* was originally given to the liquid produced by pressing the seeds or fruit of the olive tree, but, to-day, oils are grouped into two main types: *vegetable oils* obtained from the seeds or fruit of trees or plants, and *mineral oils* obtained from the earth. It is the latter type of oil which you will study here.

How Oil is obtained from the earth

Oil from the earth was known to the Greeks and Romans, who found it as a scum on the surface of ponds, but it was not until 1860 that its value was appreciated. To-day it is used, not only as an illuminant, but as a fuel for aircraft, motor-cars, trains and ships. Moreover, more than two hundred by-products can be obtained from it.

Rock oil, as it is called, is found in many countries, and the diagram (Fig. 22) shows our main sources of supply. It occurs in the earth at various depths, sometimes

floating on salt water, and usually with natural gas above it (Fig. 23). No one knows just how it has been formed,

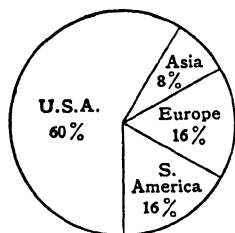


FIG. 22.

but some scientists believe that it has been produced by the slow decay of vegetable or animal matter under the action of heat. In other words, they think that it was formed by a process similar to that used for the manufacture of coal gas, and that the rock oil left in the ground corresponds to

the coal tar left in the retorts. Whether this is correct we do not know, but it seems quite likely.

Sometimes the gas escapes from the earth of its own accord, and the oil itself often finds its way to

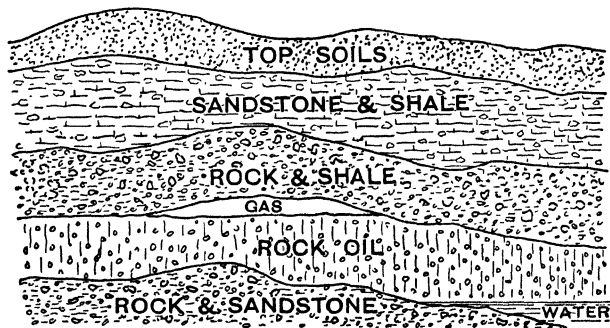


FIG. 23.

the surface as a spring. In order to obtain it in large quantities, deep wells are sunk in the ground and the oil pumped from them. Sometimes there is so much gas present that the oil is forced to the surface as soon as it is free. Such wells are called *gushers* and thousands

of gallons are often forced out before the flow can be controlled.

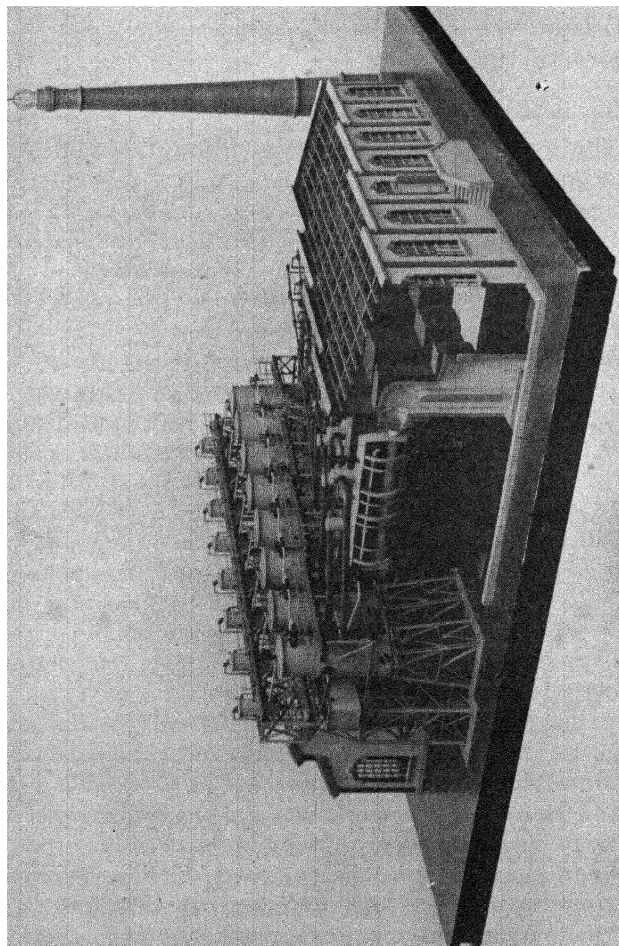
What is Rock Oil ?

The rock oil pumped up from the well is said to be *crude*, meaning that it is impure. It is a very complex substance, being made up of a large number of different compounds of carbon and hydrogen—called *hydrocarbons*—with impurities of sulphur and nitrogen. To separate these various hydrocarbons the method of *fractional distillation* is employed.

In the early method of separation, the crude oil was placed in a steel boiler or *still* and heated by a furnace. As the temperature rose, so the various constituents distilled over. At intervals, the products of distillation were removed and later re-distilled to purify them still further.

To-day, this method has been replaced by what may be called a *continuous* process. The crude oil is allowed to flow through a number of separate stills, maintained at different temperatures. The first still is kept at a relatively low temperature, and from this are obtained *petrol* and *naphtha*. From the next, kept at a higher temperature, comes paraffin, and then follow heavier oils used for preparing lubricants, petroleum jelly and paraffin wax. The latter is used for candles and as an electric insulator.

Each fraction is then refined or purified still more. The petrol fraction is re-distilled and yields first what we know as an aviation petrol, then come petrol for motor-cars and benzine (not benzene), used chiefly as a solvent in cleaning and for the manufacture of paint. The heavier oils are also refined and separated into various grades for lubrication.



CONTINUOUS OIL DISTILLATION PLANT
From the model in the Science Museum (by permission)

Another form of rock oil, known as *asphalt oil*, is treated by a similar process and yields heavy fuel oils, used in many steamships in place of coal, and a substance known as artificial asphalt, used for paving and water-proofing.

Asphalt

Natural pitch or asphalt—another hydrocarbon—is found in many parts of the world, chiefly North and South America. It occurs near the surface and is easily mined without boring. When mixed with sand and lime it forms a useful material for paving streets. The mixture readily becomes liquid on heating, but, when cold, forms a hard durable surface.

Rock oil and natural asphalt, together with the gas found in oil deposits, are classed as a special group of hydrocarbons, and are given the name of *bitumens*.

SUMMARY OF THE CHAPTER

Things to Note

1. The formation of coal.
2. The forms and uses of carbon.
3. The manufacture of coal gas.
4. The necessity for an adequate air supply when using gas.
5. The construction of the bunsen burner and the gas mantle.
6. The value of coal tar.
7. The products obtained from rock oil.
8. The comparative uses of coal and oil.

CHAPTER 6

IRON AND STEEL

YOUR history will have taught you that thousands of years ago man used stones for the heads of his tools and weapons, and that later, when he discovered copper and tin, bronze replaced the stone. When iron was first discovered, it was considered more valuable than gold, owing to the difficulty of separating it from the *ore* in which it was found. To-day, its value lies not in its scarcity, for it is one of the most plentiful of substances, but in its usefulness. It can be made almost as hard as diamond, as soft as copper or as brittle as glass. It can be beaten into sheets or drawn into fine wires. It conducts heat and electricity, and can be made magnetic.

Iron ores

Iron occurs in plentiful quantities in the earth, sometimes very near the surface, yet it is rarely found in the pure state. You have already learnt that it readily combines with oxygen to form iron rust or smithy scale (p. 14), and the following experiment will show you that it combines with sulphur just as easily.

Experiment 1

Test some iron filings with a magnet and then mix them with an equal volume of sulphur in a crucible.

Heat the mixture strongly for some minutes, until any excess sulphur has boiled off. Allow to cool (Fig. 24).

- (a) Describe the form of the substance in the crucible.
- (b) Test it with a magnet.

By heating these two substances together, a compound of iron and sulphur is formed—*iron sulphide* or *iron pyrites*. This happened in the early days of the earth, when it was still in a molten state and the metals present combined with other substances—sulphur or oxygen—to form compounds. For this reason, metals, with the exception of gold, platinum and silver, which do not combine on heating, are rarely found in the free state and are mined as metallic ores.

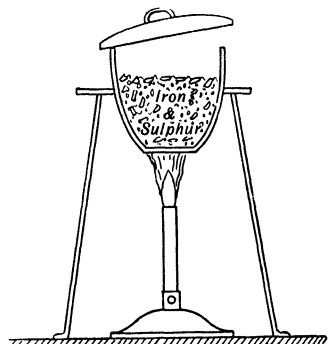


FIG. 24.

Iron occurs in different forms, as red or brown *haematite*, which are oxides; as *magnetite* or magnetic ore, yet another oxide, and as *pyrites* or iron sulphide. All ores contain impurities, often in the form of sulphur, carbon, and sometimes copper.

The Blast Furnace

In order to obtain iron from the ore, three things have to be done. In the first place, the impurities mixed with the ore must be removed. Secondly, the iron must be separated from the oxygen combined with it, and thirdly, it must be made tougher by the addition of carbon.

These three operations are accomplished by what is called a *blast furnace*, and the iron ore tipped in at one

end runs out as molten iron at the bottom.

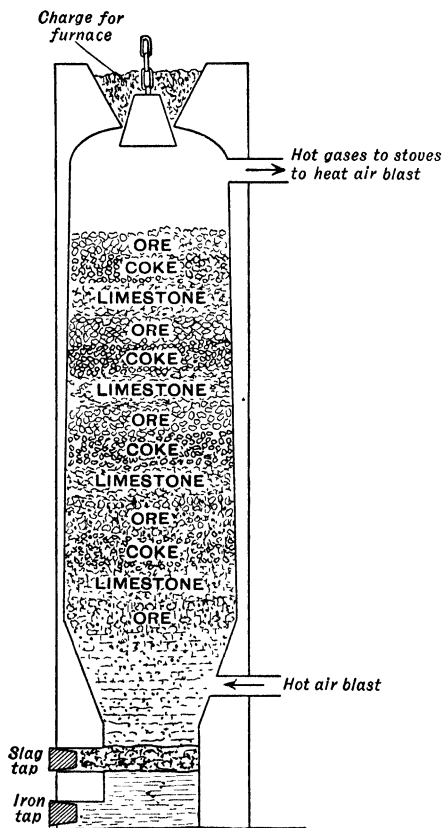


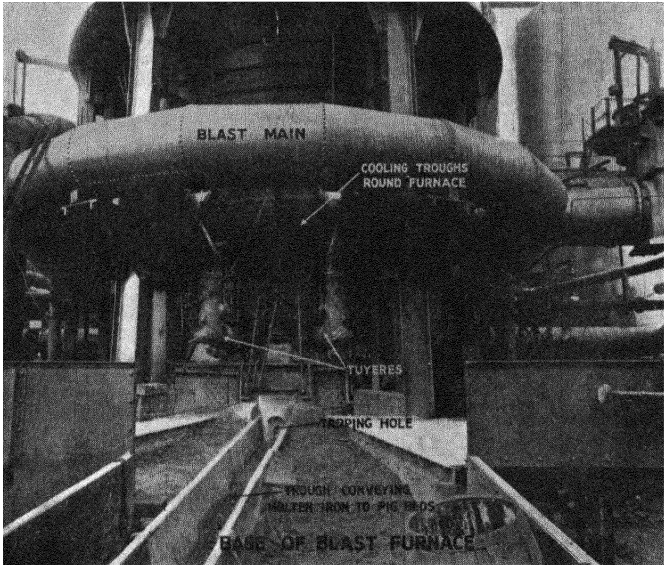
FIG. 25.

Fig. 25 shows a simplified diagram of a blast furnace.

The towers are sometimes over a hundred feet high, and from the outside appear like large chimneys, widening at the bottom for some distance and then gradually tapering towards the top. These furnaces are lined on the inside with a special type of fire brick. The iron ore, mixed with coke and limestone, is fed into the top through the valve *A*. At the base a hot air blast is led into the furnace

by a number of jets, and this raises the temperature of combustion. Once the furnace is in operation, the hot

gases produced in it are used in special stoves to raise the temperature of the air blast. Here is one of the reasons why it is costly to set blast furnaces in operation and explains why they are kept working almost continually.



(Courtesy, The Staley Coal and Iron Co)

Moreover, when a furnace is allowed to stop working, the fire brick lining has to be renewed and this, too, is a very expensive task.

Now what happens to the ore in the furnace, and what is the purpose of the lime and coke? The iron oxide is reduced by the gases produced by the burning fuel. The oxygen in the air blast combines with the coke to

form carbon dioxide, and this in turn is reduced, by the excess of coke, to carbon monoxide.

i.e. Carbon + oxygen \rightarrow carbon dioxide.

Carbon dioxide + carbon (coke) \rightarrow carbon monoxide.

The ascending carbon monoxide reacts with the iron oxide to form metallic iron and carbon dioxide,

Iron oxide + carbon monoxide,
 \longrightarrow Iron + carbon dioxide :

The molten iron, together with the impurities which combine with the limestone to form a slag, flows to the bottom of the furnace, where they collect with the slag floating on the iron. The gases ascending the tower consist chiefly of nitrogen, carbon dioxide, carbon monoxide with small quantities of hydrogen and hydrocarbons. These are led away and used to heat the air blast. Even the slag is used, sometimes as a road material or for railway ballast.

From time to time the molten iron is run off and poured into small sand or fire-clay moulds, where it solidifies as *pig iron*. This name originated from the earlier method, when the iron was run into sand moulds which looked like a mother pig with her family of little pigs around her. To-day, the iron is often converted directly into steel while it is still hot, thus making it unnecessary to re-melt it.

Cast and Wrought Iron

The iron from the blast furnace contains carbon dissolved in it, or combined with it, and the amount of carbon present determines the characteristics of the iron.

Thus cast iron contains 3 to 4 per cent. of carbon and is hard, but brittle when cold. It is used very extensively for objects which do not have to withstand severe strains or shocks—hot water radiators, roof gutters, pipes, stoves, etc.

Wrought iron contains but little carbon and is prepared from pig iron by what is called the *puddling process*. Pig iron and iron ore are mixed and heated together in a furnace. The carbon in the pig iron combines with the oxygen in the ore to form first carbon monoxide, and this burns to give the dioxide. Other impurities still present are also oxidised and the iron drawn off contains very little carbon. *Wrought* iron has quite different properties from those of cast iron. It melts at a higher temperature and is easily beaten into sheets or drawn into wire. It is used for nails, chains, horseshoes, but is rapidly being replaced by a form of steel.

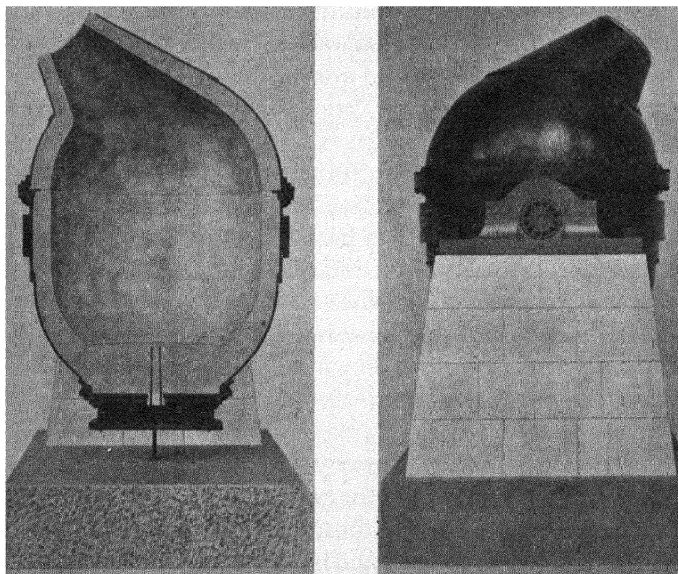
Steel

Less than one hundred years ago steel was known as a form of iron containing less carbon than cast iron but more than wrought iron, but, to-day, there are many forms of iron called steel, and these contain a variety of metals in addition to carbon.

In 1856 an English inventor—Sir Henry Bessemer—following on the work of an American, William Kelly, introduced a very simple method for making carbon steel.

Molten iron from the blast furnace is run into an egg-shaped vessel with an open top, called a *converter*. This is lined with clay, and a blast of air is forced through the iron. During this process the carbon and impurities combine with the oxygen in the air blast and are forced

out as gases, or join up with the clay lining as a slag. This oxidation is accompanied by a great amount of heat, and the iron is kept in its liquid state without the necessity for external heating. In about fifteen minutes



MODEL OF A BESSEMER CONVERTER

From model in Science Museum (by permission)

the process is complete and the steel is ready for moulding or rolling into shape.

Another process, which produces even better steel, is known as the *open hearth method*. Pig iron, wrought iron and iron ore are mixed and placed in a shallow basin in what is called a *regenerative furnace*. This contains valves

on each side (Fig. 26), and at first the gas enters by valve *A* and, when burnt, leaves by valve *B*, heating the exit pipe in its passage. After a while, the gas is forced to enter by way of the hot valve *B* and leave by *A*, and by

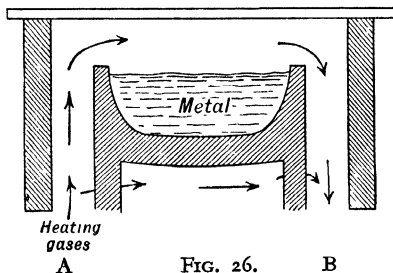


FIG. 26. B

reversing the direction at intervals, the temperature of the incoming gas is raised and greater heat is produced in the furnace.

This process takes much longer than the Bessemer process, but it has the advantage of producing a steel of better quality.

Metals and Alloys

Iron has already been classed as a *metal* and it will be as well to consider just what this term means before proceeding further. All substances may be roughly grouped under one or other of the headings—*metals* and *non-metals*. Those we know and call metals are, with the exception of mercury, solid at ordinary temperatures and melt only at high temperatures. They are good conductors of heat and electricity and may be hammered into sheets and drawn into wires. Moreover, they have a lustre or shine and, if cut and polished, they reflect light.

This is but a very general classification and it is possible to find many non-metallic substances which possess some of the properties of metals. Thus carbon has a very high melting point and may be drawn into wires, but we know it as a non-metal. Later you will learn one other great difference between substances of these two groups, and here we are concerned with those substances which help to make the many types of steel we have to-day.

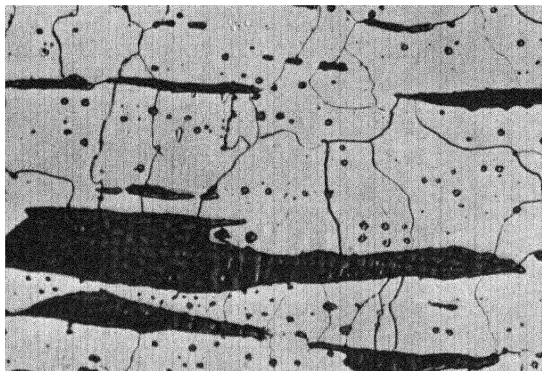
For many years it has been known that a mixture or *alloy* of two metals possessed quite different properties from those of the metals composing it. Thus brass is hard, while copper and zinc from which it is made are much softer. Again, *sterling silver* is hard, but is made from two soft metals—copper and silver.

Those alloys containing mercury are given the special name of *amalgams*, and these have been mentioned in Chapter 4. For the present, only those which form the many types of steel will be considered.

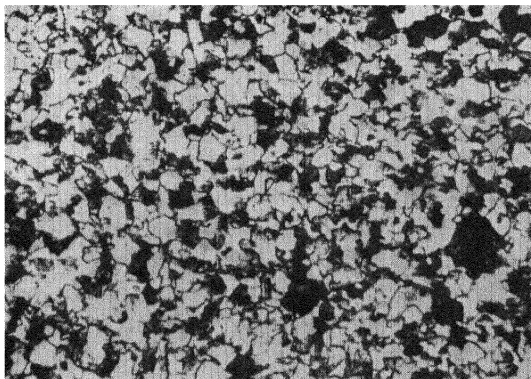
High Speed Steels

The steel tools used on machines and lathes for cutting and drilling metals have to withstand the great heat produced by the friction between the swiftly moving surfaces. For this reason the steel used for these tools is given the name of *high speed steel* to distinguish it from ordinary steel, which would soon lose its edge and would be useless for work of this type.

As early as 1868 an Englishman, Robert Mushet, found that a lathe tool he was using kept its edge remarkably well, although it was nearly red-hot. He analysed the steel and found it to be an alloy of iron and a metal, *tungsten*.



WROUGHT IRON COMPOSITION
(Microphotograph of section)
(Courtesy, The Ford Motor Co)

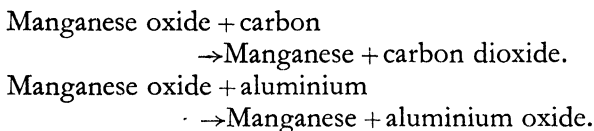


CHROME STEEL COMPOSITION
(Microphotograph of section)
(Courtesy, The Ford Motor Co)

Tungsten is found as an ore, *scheelite*, and, when separated from it, is used considerably for the filaments of electric light bulbs. Among the uses of its compounds is the making flannelette non-inflammable, since after the fabric is soaked in the tungsten solution, it smoulders but does not flame.

Another very hard steel is formed by mixing a definite quantity of the metal *manganese* with pure iron. This steel is so hard that no lathe tool will cut or drill it and its greatest use is for armour plate.

Manganese is found chiefly in the form of oxides, and is prepared from them by heating to a high temperature with either carbon or aluminium.



Other important steels include *chrome steel* (chromium and iron) used for ball bearings and files, *molybdenum steel* used for lathe tools, and *nickel steel* which is very hard and tough, used for safes and for those machine parts which have to withstand great strain.

In the manufacture of these steels great care has to be taken to make sure that the metals are present in correct proportions. The characteristics of the steel depend upon this, and different types of steel can be prepared from the same materials if the proportions are varied. Thus carbon steel with .9 per cent. of carbon is hard, while, when there is only .4 per cent. of carbon present, the steel is soft or *mild*. Mild steel has replaced wrought iron to a very great extent.

SUMMARY OF THE CHAPTER

Things to Note

1. The purpose and use of the blast furnace.
2. The value of carbon in the manufacture of iron.
3. The characteristics of cast and wrought iron.
4. The manufacture of carbon steel.
5. What high speed steels are made of.

CHAPTER 7

MORE USEFUL METALS

IRON and steel are the most useful metals we have, but there are many others which play important parts in our lives. Some of these have been known and used for hundreds of years, while others have only come into use during recent years. For example, aluminium, used today for kitchen ware and for the framework of aircraft, was almost unused twenty-five years ago.

Aluminium and its alloys

Although aluminium is one of the most plentiful of substances in the earth, it is never found in the free state. The chief ore from which the metal is obtained is called *bauxite*, and this is a form of oxide containing water.

Electricity is employed to separate the metal from the ore. A solution of bauxite is poured into an iron tank and a current of electricity passed through it. The tank itself is used as negative, while a number of carbon plates form the positive. The electric current splits up the oxide and aluminium collects at the bottom of the tank while oxygen escapes (Fig. 27).

Aluminium is bluish white in colour and can be very highly polished. It is very light, having a density one third that of iron, and can be rolled into thin sheets and drawn into wire. Its lightness and great strength make it

a very suitable material for use in the construction of the framework of both airships and aeroplanes.

When left in contact with dry air it oxidises very slowly, but in moist air a thin film of greyish oxide soon forms on its surface, and this seems to protect it from further change. This, together with the fact that it is a good conductor of heat, makes it a suitable material for kettles, saucepans and kitchen utensils generally.

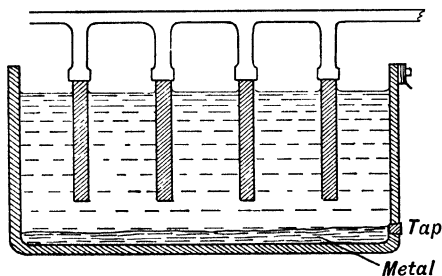


FIG. 27.

Powdered aluminium is a powerful reducing agent, and when mixed with some metallic oxides and ignited, reduces them to metals.

Metallic oxide + aluminium \rightarrow metal + aluminium oxide.

During this action great heat is produced and the temperature is often raised to 3000° . This property of aluminium is sometimes used for welding iron. A mixture of the powder and iron oxide together with a cement is packed around the ends of the metal to be joined. The mixture is ignited and the heat produced used to melt the iron and weld or join the two pieces together.

Aluminium forms one or two very important alloys ;

magnalium, formed from aluminium and magnesium, is easier to *machine* than the pure metal, while *duralumin*—a mixture of aluminium, copper and magnesium in the ratio 96 : 3 : 1—is stronger and was used extensively in the war for shell cases.

Copper

In contrast to aluminium, copper has been known and used for many hundreds of years, probably because it is often found in metallic form, although the chief sources are the ores containing copper oxide or copper sulphide.

In order to separate the metal from the ore, it is first crushed to powder and then washed to free it from earth. It is then passed into a roasting furnace where the sulphur is driven out. From this furnace it passes into a *converter* and here burning coal dust is blown into it. The carbon combines with the oxygen left in the roasted ore, and copper, still rather impure, runs off and is cast into moulds.

The final process involves the use of electricity. Plates of crude copper are suspended in a solution of copper sulphate (with a little sulphuric acid added). Between these plates are placed sheets of pure copper. The plates are connected to a source of electricity, and a current allowed to flow. Pure copper leaves the crude plates and is deposited on the thin sheets, while the impurities form a *mud* at the bottom of the tank. By this method almost pure copper is obtained.

One of the greatest uses of copper is for electrical apparatus, since it is an excellent conductor. Moreover, it is easily drawn into wires of any thickness, and its softness allows it to be beaten to any shape.

Copper is an ingredient in many useful alloys—*brass* is formed from zinc and copper, *bronze* formed from tin and copper, and *german silver* made from copper, nickel and zinc. There are others, chiefly formed from bronze by adding quantities of manganese, aluminium and phosphorus.

Coating Metals with Copper

Iron and steel, which easily rust, and brass, which easily tarnishes, are often protected by coating them with a thin layer of non-rusting metal such as copper, nickel silver or chromium, and the process by which this is done we know as *electroplating*.

Experiment 1

Mix a solution in the following proportions :

Copper sulphate	4 oz.
Sulphuric acid	$\frac{3}{4}$ fl. oz.
Water	1 pint.

Set up the apparatus as shown in Fig. 28, connecting the copper plate to the positive (anode) and the carbon rod to the negative (cathode).

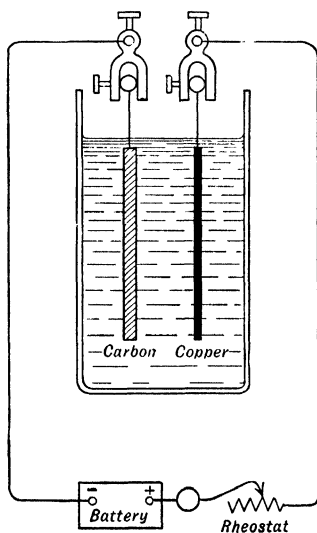


FIG. 28.

Allow an electric current to pass for five or ten minutes and then remove the carbon rod.

Replace the rod in the solution and reverse the direction of the current for ten minutes.

Observe the changes that take place.

Notice if any gas is given off from either the anode or the cathode during the experiment.

Experiment 2

Repeat Experiment 1, using a carbon rod in place of the copper plate.

Observe any changes which take place.

Notice if any gas is given off from either the anode or the cathode during the experiment.

Experiment 3

Select an object to be plated and clean it thoroughly with fine emery paper. Wash it in dilute acid and then place it in caustic soda solution to remove grease. Rinse the object in cold water and then suspend it in the plating bath as cathode.

Allow a small current to pass for about half an hour and then remove the object and rinse it.

For the time being, Experiments 1 and 2 are more important than Experiment 3, as they give some clues as to what actually happens when an electric current is allowed to pass through the solution. In the first experiment, when the carbon rod is the negative or cathode, copper is deposited on it. Now there are two possibilities : either the copper has come from the positive or anode, or it has come from the solution of copper sulphate. When the direction of the current is reversed, the copper disappears from the carbon rod, and this seems to indicate that

copper actually leaves the anode when electricity passes through the solution.

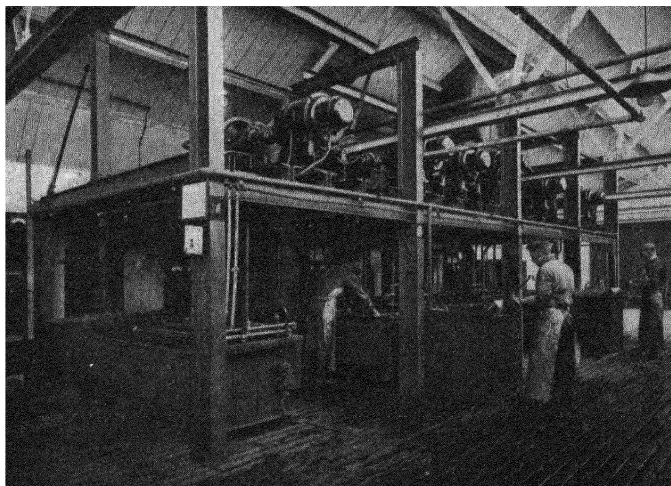
In Experiment 2 the cathode becomes coated with copper, even though a carbon rod has been used for anode, showing that the copper must have come from the solution. Thus we have two apparently contradictory results and they must be explained in some way. Scientists still differ in their explanations of what really happens, but it will be possible to give you at least one of them.

It is supposed that on dissolving, each tiny particle of copper sulphate (which contains copper, sulphur and oxygen combined) splits up into two parts called *ions*—a copper ion and a sulphate ion. These are charged with electricity and, when the plates are connected to a battery, the copper ions travel towards the cathode, give up their tiny electric charges and are deposited on it as metallic copper. The sulphate ions move to the anode, and if this is of copper, they attract tiny particles of it, forming copper sulphate, which dissolves and immediately splits up again into ions. If the anode is of carbon, the sulphate ion, which cannot exist alone, attacks the water, taking hydrogen from it to form sulphuric acid (hydrogen sulphate) and liberates oxygen. This accounts for the bubbles which appear in Experiment 2. To summarise these changes, here they are as equations :

1. Copper sulphate \rightarrow copper ions + sulphate ions.
2. $\left\{ \begin{array}{l} \text{Sulphate ions + copper ions} \rightarrow \text{copper sulphate.} \\ \text{Sulphate ions + hydrogen ions + oxygen ions} \\ \qquad \qquad \qquad \rightarrow \text{sulphuric acid + oxygen.} \end{array} \right.$

Your copper deposit will probably be soft and you will

be able to rub it off. In the electroplating works great care is taken to make the article to be plated thoroughly clean. The solution, or *electrolyte*, must be of correct strength and the current must be carefully regulated according to the size of the object. By doing this the deposit is made hard and adheres firmly to the base metal.



ELECTROPLATING MOTOR CAR RADIATORS

(By courtesy of Morris Motors Ltd.)

By using anodes of silver or nickel and appropriate compounds in solution, articles can be silver or nickel plated. These are not the only uses for electroplating; it is used a great deal for the manufacture of printing blocks. A wax mould or impression is made of the type or picture, and this is coated with graphite to make it conduct electricity. A thin plate of copper is deposited on

this, and when it is stout enough the wax is melted off and the *electrotype* backed with a block of type metal to strengthen it. Electrotype is used for most books since the plates are strong and give good, clear print.

More useful Metals

All metals are, of course, useful, but space does not allow each to receive separate, detailed treatment, and by this time you will be anxious to learn something of other substances besides metals. However, there are one or two which cannot be left out. There is lead, found chiefly as lead sulphide or *galena*; tin, found as tin oxide, and zinc, found as both oxide and sulphide.

These metals are extracted from the ores by first changing the ore to oxide by roasting in oxygen or air :

Lead sulphide + oxygen \rightarrow lead oxide + sulphur oxide.

The oxide produced is then reduced to metal by the use of carbon, thus :

Lead oxide + carbon \rightarrow lead + carbon oxides.

Experiment 4

Cut a shallow hole in a carbon block and fill this with lead oxide (litharge). Now heat the oxide with a blow-pipe as shown in Fig. 29.

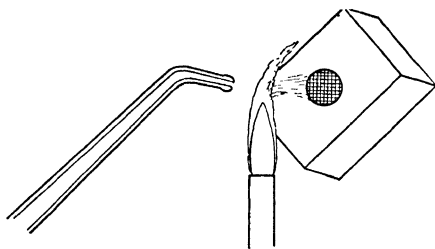


FIG. 29.

Notice the changes which take place in the oxide and observe the tiny lead beads which form on the sides of the hole.

The importance of these metals lies in their uses. Lead is used for water pipes and as a sheathing for electric cables, for its softness and the ease with which it can be soldered make it ideal for the purpose. Tin is used as a protective coating for iron and steel since it does not rust, and also as a constituent in many alloys. Solder is made from equal parts of tin and lead, and pewter contains the same metals in the ratio of 3 : 1. Zinc is used in the construction of Leclanché cells and dry batteries, and for galvanising or protecting iron since it does not rust.

SUMMARY OF THE CHAPTER

Things to Note

1. The extraction of aluminium from the ore.
2. The uses of copper, aluminium, tin, lead and zinc.
3. How electroplating is done.
4. Chemical actions during electroplating.

CHAPTER 8

ACIDS AND THEIR USES

ALMOST every household uses vinegar, and its sharp or sour taste will be well known to you. This liquid belongs to another important group of substances which we call *acids*. One of their common properties is this sharpness of taste, but, while vinegar is harmless unless taken to excess, other acids are violent poisons and must be used with great care. For this reason we cannot safely use their sharpness or sour taste as a test and must find some other means of distinguishing them. Fortunately, there is one very simple test which is quite safe and which you can all use.

Experiment 1

Test small quantities of all the liquids in the science room called acids by dipping a piece of blue litmus paper in each.

(a) What change takes place in the colour of the paper?

Experiment 2

Repeat the experiment, using vinegar, lemon juice, and any other fruit juice you can obtain.

Experiment 3

Now test the following substances : water, milk and paraffin.

Your results have shown you that those liquids labelled as acids change the colour of the litmus paper from blue to red. Vinegar and the fruit juices act in the same way and may be included in the same group.

At first, acids were believed to be special compounds of oxygen, and it was the presence of this element that was thought to give them their common characteristics. Sir Humphry Davy, a scientist whose name you have met in the previous volume, discovered that one acid at least contained no oxygen whatever. Thus the oxygen theory was proved to be false.

Later it was shown that all acids do contain hydrogen. Now hydrogen is present in thousands of compounds and you must not think that all these are acids. For example, water, methane and acetylene are all hydrogen compounds, but they are not acids.

Acids are of great use to man. They are present in the gastric juices and help in the digestion of food. In industry they play a very important part, being used in the preparation and purification of very many substances. Yet acids can be extremely dangerous. They are always eager to break up other substances and will cause very serious wounds if left in contact with human flesh. In clothing they produce holes and they soon destroy wood. To say that they *burn* is not strictly accurate, and a better and more correct description is to say that they *corrode* or wear away substances gradually. Again, acids are often used in the preparation of violent explosives—nitro-cellulose and nitro-glycerine are made by treating cotton and glycerine with nitric and sulphuric acids.

In the science room, to prevent accidents, acids are used *dilute*—that is, mixed with a considerable quantity of

water. Even then they can be dangerous and are still poisonous, so care must be taken when using them.

Acids and Metals

In Chapter 3, hydrogen gas was prepared by the action of sulphuric acid on zinc. Here are some experiments, using metals and acids :

Experiment 4

Add a small quantity of zinc to a little dilute sulphuric acid in a test tube. Notice that bubbles of gas come from the zinc. When the zinc has disappeared, add a little more and do this until the action ceases.

Test the liquid in the tube with blue litmus paper.

Filter the liquid left in the tube and allow it to evaporate in a watch glass or shallow basin.

Experiment 5

Repeat the previous experiment, using magnesium and sulphuric acid. Warm the tube to start the action.

Allow the liquid to evaporate and examine the crystals formed. Taste the crystals and then taste some Epsom salt.

Experiment 6

Into separate test tubes pour small quantities of sulphuric, hydrochloric and nitric acids. To each tube add a few pieces of zinc.

Gently warm the tubes to commence the action.

Describe what happens in each case and notice particularly the results with nitric acid.

Repeat the experiment, using copper.

From these experiments a number of important facts can be obtained. In the first, the zinc disappears and is often said to *dissolve*. The liquid which is left, however, is not acid, and by evaporation white crystals are obtained which are quite different from the original zinc. Now, when salt is dissolved in water and the solution evaporated, white crystals of salt re-appear. The salt is not changed to some other substance, but remains unaltered. Thus, if the zinc dissolved, we should expect to get zinc from the liquid after evaporation. Actually, chemical action takes place; the acid attacks the metal, forming hydrogen gas and a new substance, called *zinc sulphate*, and the hydrogen in the sulphuric acid has been replaced by zinc.

i.e. Zinc + sulphuric acid \rightarrow zinc sulphate + hydrogen.

A similar action takes place in the second experiment, but here the white crystals can easily be recognised as Epsom salts, or, to give them their chemical name—magnesium sulphate.

Magnesium + sulphuric acid
 \rightarrow magnesium sulphate + hydrogen.

Here is a good example of the great difference between the properties of a compound and those of the substances which make it. Magnesium is a metal which burns with a brilliant flame, sulphuric acid is a violent poison, but Epsom salts forms a useful medicine.

In Experiment 6, your results have shown you that the zinc is attacked by each acid. With hydrochloric and sulphuric acids, hydrogen is given off, but with nitric acid you may have obtained brown fumes or a gas which

will not burn. When cold, dilute sulphuric and hydrochloric acids have very little effect on copper, but on warming, sulphuric acid attacks the copper to form a gas with a very irritating smell of burning sulphur—the same gas that is produced when sulphur burns in air.

i.e. Copper + sulphuric acid

→copper sulphate + sulphur dioxide.

With nitric acid, the copper is attacked even when cold and brown fumes are produced. These consist of one of the oxides of nitrogen called *nitrogen peroxide*.

Sulphur and Sulphuric acid

Sulphur, as you have met it, is a yellow solid which burns in air with a blue flame, giving a gas with a very irritating smell. It is found in the neighbourhood of volcanoes in an impure state and also occurs combined with metals. Iron pyrites, already mentioned on page 69, lead sulphide or galena, and zinc sulphide or zinc blende are the chief sources of sulphur in this form and are often used for the extraction of the metals.

A great deal of the sulphur we use is prepared from the impure free sulphur earths found in Sicily. This is placed in furnaces with sloping floors, and the sulphur at the base ignited. The heat produced is sufficient to melt the remaining sulphur, and this is drawn off at intervals. To purify the sulphur still more, it is distilled from large retorts and allowed to cool in brick collecting chambers (Fig. 30).

Sulphur is a substance with many uses. In its pure form it is a valuable medicine and is used as an ingredient in gunpowder and vulcanite and for making matches.

Crude or impure sulphur is used for making sulphur dioxide, and this gas, in spite of its very poisonous nature, is very useful. Sulphur dioxide is easily liquefied, requiring only a slight increase in pressure at 0° C. to change it to a colourless liquid.

Liquid sulphur dioxide is a good solvent and is used to dissolve sulphur, iodine, phosphorus and many forms

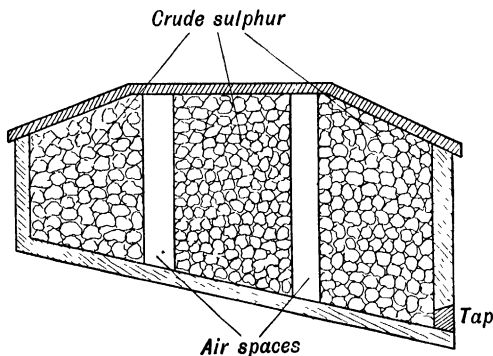


FIG. 30.

of resin. But its use does not stop here. Sulphur dioxide boils at -8° C. and, if allowed to evaporate under ordinary room temperature, extracts a considerable amount of heat from its surroundings. Very low temperatures can be obtained in this way and sulphur dioxide is often used in refrigerators (Book I, p. 55).

Sulphuric acid, too, has a great many uses, but before studying these, learn how it is made. There are two main methods of manufacture. In the first, the *chamber process*, which is over a hundred years old, sulphur dioxide is prepared either by burning crude sulphur in a current of air, or by roasting iron pyrites. A mixture of

sulphur dioxide and air is then passed through strong nitric acid and the resulting gases allowed to enter large lead tanks. Steam is forced into these tanks and many chemical changes take place. As a result, the sulphur dioxide is changed to another oxide of sulphur, called sulphur trioxide, and this dissolves in the steam and condenses as sulphuric acid. This *chamber acid* is then drawn off and purified.

A more modern process is known as the *contact process* and this is much easier to describe. Sulphur dioxide is prepared by a variety of means, from crude sulphur in rotary cylinders, or from pyrites by burning them in special furnaces. These furnaces are so designed that the sulphur or the pyrites are completely burnt to sulphur dioxide. The mixture of gases from the furnaces is carefully purified by passing it into steam chambers and then through towers down which passes a spray of water. In very modern factories, the gas is cleansed by passing an electric current through it. By this method the tiny particles of dust and impurities are deposited, and the gas issuing from these *electrostatic* cylinders is pure enough for the final processes.

The gases are next passed into what are called the *contact towers* (Fig. 31). Here they pass through perforated shelves containing platinised asbestos. This material is made by soaking asbestos in a solution of a platinum compound. After igniting, the asbestos contains platinum metal in a very finely divided state. As the gases pass through these shelves, the sulphur dioxide is changed to sulphur trioxide, and owing to the great heat which accompanies this change, special methods have to be adopted to cool the towers. This is generally done

by passing the gas through the towers twice, first through the centre in pipes. The heat produced by the combination of the dioxide with the oxygen is partly used in heating the sulphur dioxide and air in the pipes, and this has the dual effect of heating these gases ready for their next entrance,

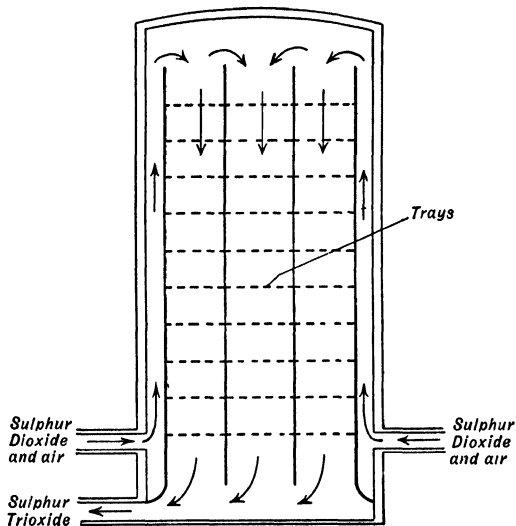


FIG. 31.

and of cooling the tower. On their second entrance into the towers they come into contact with the platinum and combine to form a whitish mist of sulphur trioxide.

After its exit from the contact towers the trioxide is cooled, and then passed into the *absorption towers* where it is changed into sulphuric acid. These towers are simply steel cylinders lined with brick and filled with broken quartz. The sulphur trioxide enters near the bottom and meets a falling spray of dilute sulphuric acid from jets at

the top of the tower. The oxide dissolves readily enough in the acid, making it stronger, and this is drawn off from the bottom at intervals. The acid produced in one tower is often used as the spray in another and gradually becomes stronger or more concentrated.

The platinum used in the contact towers does not change during the oxidation of the sulphur dioxide, but without its presence the oxidation would not take place. Another substance which is sometimes used in place of the platinum is called vanadium oxide. Both these substances are costly and this explains the great care taken to purify the gases before their entrance into the contact towers. The presence of impurities has the effect of spoiling the platinum and destroying its power.

In the preparation of oxygen on page 13, the presence of the black manganese dioxide enables the potassium chlorate to liberate oxygen quite freely, but remains unchanged itself. Substances such as these, which help, but which are themselves unchanged, are known as *catalytic agents* and the process as *catalysis*.

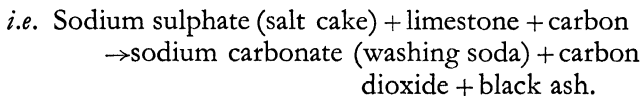
How we use Sulphuric acid

Over one million tons of sulphuric acid are produced in England every year, and this will serve to indicate that it is a very useful and essential substance in our lives.

One of the most important uses is in the manufacture of washing soda. Common salt is heated in cast iron pans with sulphuric acid and, as a result, hydrochloric acid gas is given off and a substance, sodium sulphate (Glauber's salt), left behind.

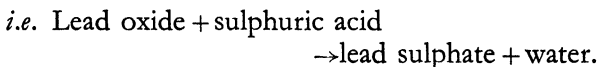
i.e. Salt + sulphuric acid
→sodium sulphate + hydrochloric acid.

The acid fumes are led away and collected, while the sodium sulphate or salt cake is mixed with limestone and coal dust and again heated. Chemical action again takes place and washing soda (sodium carbonate) with another substance called black ash are left behind, while carbon dioxide gas is produced.

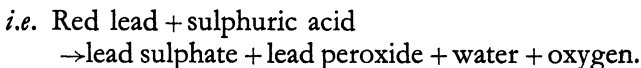


In order to separate the soda from the black ash, the mixture is washed with water. The black ash does not dissolve and can be removed by filtration while the solution of soda is allowed to crystallise.

Sulphuric acid, too, forms the electrolyte in lead accumulators. In Book I, p. 129, directions are given for the construction of a simple accumulator. Here, you will find that the plates are of lead and those marked + are filled with a paste of red lead (an oxide of lead). Those marked - are filled with litharge paste, that is, with an oxide of lead which contains less oxygen than red lead. On placing the plates in dilute sulphuric acid, the paste on the negative plates is changed to white lead sulphate :



The positive plates are changed, partly to lead sulphate and partly to a further oxide of lead called *lead peroxide*.



When the plates are connected to a supply of electricity, the hydrogen ions travel to the negative plate,

give up their electric charges and reduce the lead sulphate to grey lead and form sulphuric acid.

i.e. Lead sulphate + hydrogen \rightarrow lead + sulphuric acid.

The sulphate ions travel to the positive plate and give up their charges, combining with the water present to form sulphuric acid and liberating oxygen. This oxygen reacts with the lead sulphate, changing it to the brown lead peroxide and forming more sulphuric acid.

i.e. Lead sulphate + oxygen + water
 \rightarrow lead peroxide + sulphuric acid.

When all the lead sulphate has been changed the plates give off their gas, hydrogen at the negative and oxygen at the positive, and the accumulator is charged. As you can see from the equations, every action produces sulphuric acid and thus, during the charge, the strength or density of the acid increases.

During discharge, that is, when the cell is used to supply electricity, the hydrogen ions travel towards the peroxide plate and reduce it to lead oxide which reacts with the sulphuric acid to form lead sulphate. At the same time, the sulphate ions move to the lead plate and change it to lead sulphate. Thus, during discharge, sulphuric acid is extracted from the solution, leaving it more dilute.

Sulphuric acid is often used in the preparation of other acids and this use leads us to another section.

Chlorine and Hydrochloric acid

Chlorine was first discovered towards the end of the eighteenth century and was named from the Greek word, chloros—green, because of its colour. In the science

room it may be prepared by warming a mixture of salt and manganese dioxide with strong sulphuric acid.

Experiment 7

Mix together in a flask, one part of salt, one part of manganese dioxide, and four parts of strong sulphuric acid. Fit the flask with a delivery tube leading to a gas jar (Fig. 32) and warm the flask.

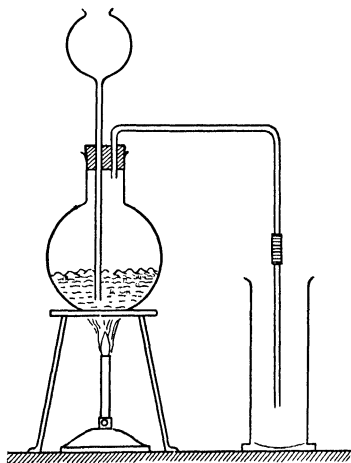


FIG. 32.

Collect three or four jars of the gas, closing them with a glass plate.

N.B.—Chlorine may also be prepared by heating manganese dioxide with hydrochloric acid.

Experiment 8

Into the first jar plunge a lighted taper.

Into the second jar plunge a tiny piece of ignited sodium in a spoon.

Into the third jar drop a piece of filter paper soaked in turpentine.

Describe what happens in each case.

During these experiments the colour and smell of chlorine will be very evident. It is, as you can imagine, poisonous, and was used in the early part of the Great War as a poisonous gas. Yet it is extremely useful and some of its properties are very interesting. You have noticed that it allows things to burn in it, the taper does not go out although its flame becomes very smoky. Sodium continues to burn brightly while turpentine, a hydrocarbon, ignites on its own accord. In the case of sodium you will find that the sides of the jar become covered with tiny white particles. Taste these and you will be surprised to find that they are salt.

i.e. Sodium + chlorine \rightarrow salt.

Here is yet another example of the great difference between the properties of a compound and those of the substances which form it. Sodium is a metal which has a great liking for oxygen, and will even extract it from water and take fire during the process, while chlorine is a poisonous gas. Yet when they combine, that very valuable substance—common salt—is formed.

Chlorine has a great affinity (liking or attraction) for hydrogen and will extract it from many hydrocarbons. Often sufficient heat is released to cause the substance to burn. A mixture of hydrogen and chlorine will explode

if exposed to sunlight, and as a result, hydrochloric acid is formed.

Hydrogen + chlorine \rightarrow hydrochloric acid^c

It is this liking for hydrogen which makes chlorine so useful. It is a powerful bleaching agent and will remove the colour from vegetable dyes. For this purpose it is sold in the convenient form of a white powder, known as *bleaching powder*. This is manufactured by passing chlorine gas into large tanks containing quicklime.

Quicklime + chlorine \rightarrow bleaching powder + water.

This powder readily gives up its chlorine when acted upon by weak acids, and the cloth to be bleached is first steeped in a solution of the powder, then in very weak acid and finally in clean water. This is repeated until the colour is removed.

Chlorine, too, in its liquid form is often mixed with drinking water to remove disease germs and bacteria.

How Hydrochloric acid is made

You have learnt that hydrochloric acid is produced when a mixture of hydrogen and chlorine is allowed to explode, and that it is obtained during the manufacture of washing soda.

Hydrochloric acid is a gas which is very soluble in water, a fact which can be shown by an interesting experiment.

Experiment 9

Prepare a flask of hydrochloric acid gas by warming some salt with sulphuric acid (Fig. 33).

Close the flask with a cork containing a narrow tube, and place this so that the end comes under water coloured blue with litmus solution.

Warm the flask for a second and then allow it to cool. Some water will enter the flask and the rapid solution

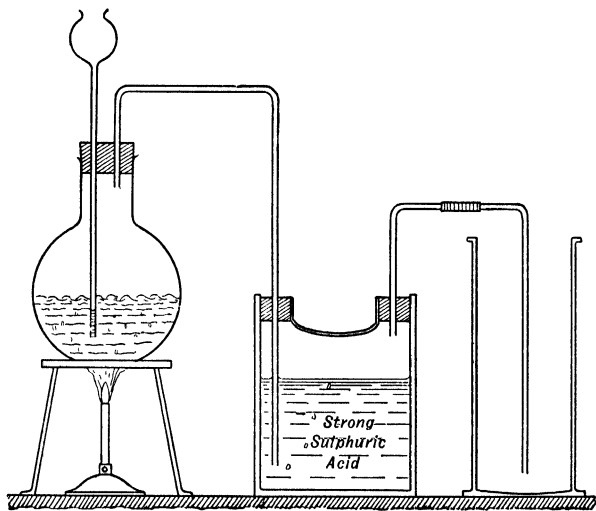


FIG. 33.

of the acid, by causing a vacuum, forces water to rush in and the acid changes the colour from blue to red.

Hydrochloric acid is known commercially as *spirits of salts*, and is used for cleaning purposes, and as a flux for soldering.

Nitric acid

Nitric acid may be prepared by forcing nitrogen and oxygen to combine in the presence of a little water. This

can be done by passing electric sparks through a mixture of the gases, but a more usual way is to heat together chili saltpetre and strong sulphuric acid.

Experiment 10

Into a retort place equal weights of saltpetre and strong sulphuric acid. Over the end of the retort place a flask and surround it with cold water (Fig. 34).

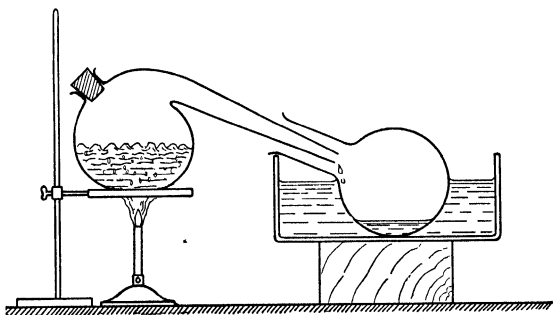


FIG. 34.

Warm the retort.

Test the liquid which collects in the retort with litmus paper.

One of the chief uses for nitric acid is in the preparation of explosives.

There are other useful acids, the most commonly known being *acetic* acid which is present in vinegar, *citric* acid present in lemons and similar fruits, and *oxalic* acid present in rhubarb and sorrel.

SUMMARY OF THE CHAPTER

Things to Note

1. The distinguishing characteristics of acids.
2. The action of acids on metals and the effect of temperature.
3. The main uses for the common acids.
4. The occurrence and separation of sulphur and its uses.

CHAPTER 9

ALKALIS AND THE SOAP INDUSTRY

IN the previous chapter you have learnt that an acid often loses its characteristics when it reacts with a metal. Thus, in Experiment 5 on page 91, the liquid remaining after sulphuric acid has been allowed to react with the metal magnesium, possesses none of the properties of an acid. The chemical action that takes place results in the formation of entirely new substances—*Epsom salt* and *hydrogen*, and neither is an acid. This change, by which an acid loses its acidic characteristics, is called *neutralisation*.

In addition to the metals there are many other substances which are able to neutralise acids. Generally, these substances are closely connected with the metals and include the majority of the metallic oxides. This is one of the distinguishing characteristics of metals, for the oxides of the non-metals are, as a general rule, acidic. Thus sulphur trioxide dissolves in water to form sulphuric acid, phosphorus oxide forms phosphoric acid, while nitrogen peroxide gives us nitric acid.

As well as the metallic oxides there are other very important substances which also possess this property, two of the most useful being caustic soda and caustic potash.

Experiment 1

Make a *weak* solution of caustic soda by shaking up one or two flakes with water.

(a) Wet your fingers with the solution and rub them together.

(b) Insert a strip of blue litmus paper. Does it change colour?

(c) Change the colour of another piece of litmus paper by dipping it in dilute acid. Now insert this in the caustic soda solution. Does the colour of the paper change?

You have noticed that the caustic soda solution felt soapy when rubbed between your fingers, and that, while it did not change the colour of blue litmus paper, it quickly restored the colour to the paper which had been made red by the acid.

Liquids or solutions which behave in this way are said to be *alkaline*, and are made from substances which we call *alkalis*. As you will discover in the following experiment, alkaline liquids possess a still more important characteristic.

Experiment 2

Make a solution of caustic soda and pour some of it into an evaporating basin. Test the liquid with red and blue litmus paper.

Add a small quantity of hydrochloric acid and mix the liquids by stirring. Test the liquid with litmus paper. If it is acid, add more caustic soda, and if alkaline, add more acid, until the liquid is just neutral.

Allow the liquid in the basin to evaporate and crystallise.

(a) What is the shape of the crystals formed?

(b) Taste the crystals and say what substance you think is present.

By your experiment you have made common salt. The alkali, caustic soda, has completely neutralised the hydrochloric acid and produced an entirely new substance.

i.e. Caustic soda + hydrochloric acid \rightarrow salt + water.

You know that hydrochloric acid is a compound of two substances—hydrogen and chlorine—and by your experiment, the hydrogen has been replaced by the sodium present in the caustic soda. It is this replacement of the hydrogen which neutralises the acid.

The new substance formed when an alkali (or a metal) reacts with an acid is given the name of a *salt*. When the acid used is hydrochloric, these salts are known as *chlorides*, and common salt possesses the chemical name of *sodium chloride*. With sulphuric acid the salts are called *sulphates*, nitric acid produces *nitrates*, citric acid—*citrates*, acetic acid—*acetates*, and so on.

In the next chapter you will learn more of the characteristics and uses of some of these salts, but in the meantime there is more to learn about the common alkalis.

Caustic soda and caustic potash

Caustic soda and caustic potash are very similar substances, the former being a compound of sodium, hydrogen and oxygen, and the latter a compound of the metal potassium, hydrogen and oxygen. The two metals, sodium and potassium, have very similar properties and this applies equally well to all their compounds.

Both soda ash and potash are present in the ash left from burning wood. By mixing this ash with lime and water and boiling the mixture, solutions of caustic soda

and caustic potash are obtained, but this method of preparation is not suitable when large quantities are required. •

One process of manufacture depends upon the fact that when slaked lime (lime with added water) is added to a boiling solution of washing soda, chemical action takes place and insoluble calcium carbonate (chalk) is produced, and the liquid remaining is caustic soda.

Washing soda + slaked lime → chalk + caustic soda.

Another more recent method of manufacture is by the electrolysis (splitting up by electricity) of a solution of common salt. The passage of an electric current through the solution results in the splitting up of the salt into sodium and chlorine. The sodium attacks the water to form caustic soda and hydrogen gas, and the latter, with the chlorine, is collected as a valuable by-product.

Caustic potash can be prepared by similar methods, using in the first, potassium carbonate in place of the washing soda, and in the second, potassium chloride in place of the common salt.

Both caustic soda and caustic potash are white substances and both have a great liking for water. They are very poisonous and corrode the skin, producing very painful wounds. Yet they are extremely useful and play important parts in the manufacture of both soap and paper.

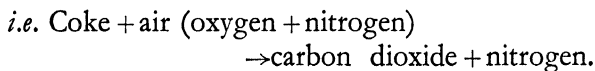
Ammonia

Ammonia, normally a gas with a very pungent smell, is a compound of nitrogen and hydrogen. Now nitrogen

is very inert, it only combines with other substances when it is forced to do so, and for many years chemists stated that it was impossible to make nitrogen and hydrogen gases combine. If the two gases were mixed together in a tube and heated for days there was still no sign of any change taking place, and yet the modern method for the manufacture of ammonia is by the direct combination of these two gases.

After repeated experiment, scientists found that a form of iron called *spongy iron* has the power to make nitrogen and hydrogen gases combine. It is, in fact, a *catalytic* agent and does not change itself during the process.

The supply of hydrogen and nitrogen is obtained in a very clever manner. Anthracite coal or coke is first heated with an air blast and, as a result, carbon dioxide and nitrogen are produced.



This air blast is then replaced by a blast of steam and *water gas* is produced. This is a mixture of hydrogen and carbon monoxide, and the two products of the coke furnace are mixed in correct proportions and washed free from tar. The carbon monoxide and carbon dioxide are removed, and the remaining gases passed over spongy iron when they combine to form ammonia.

In the science room it is more convenient to prepare ammonia by heating together sal ammoniac and slaked lime.

Experiment 3

Mix together a quantity of salammoniac with twice its weight of slaked lime. Place the mixture in a large boiling tube fitted with a delivery tube (Fig. 35) and heat.

Collect one or two jars and a flask of the gas by the downward displacement of air.

Experiment 4

Use one of the jars to test the smell of the gas (be cautious here) and to discover if it will burn.

Experiment 5

Into another jar pour a little dilute hydrochloric acid and shake. Add more acid until a piece of moist litmus paper just turns red.

Evaporate the liquid remaining.

Experiment 6

Use the flask of ammonia to demonstrate its solubility by an experiment similar to that on page 103.

These experiments have shown you several of the important properties of ammonia, you have seen that it is extremely soluble in water and that the solution is strongly alkaline. With hydrochloric acid it produces a white compound which is soluble in water, and this is the same

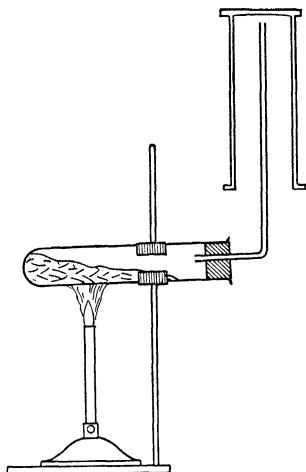


FIG. 35.

as was used in Experiment 3 for the preparation of the gas, *i.e.* salammoniac or ammonium chloride used in the manufacture of Leclanché cells and dry batteries.

i.e. Ammonia + hydrochloric acid

→ammonium chloride.

With sulphuric acid, ammonium sulphate is formed, and this is most valuable as a garden fertiliser.

Ammonia itself is also extremely useful. It is easily liquefied by compression and, like sulphur dioxide, forms a very useful material for use in refrigerators. On changing from a liquid to a vapour, a considerable amount of heat is absorbed and the temperature of surrounding objects is greatly lowered.

Another use for the gas is in the preparation of explosives.

Alkalis and Soap manufacture

To many of us a bar or cake of soap means something which cleanses the skin, and which forms a lather with water, but to a chemist all forms of soap mean the result of neutralising acids with alkalis.

The acids used are not of the same type that you have met so far, but are grouped together under the name of the *fatty acids*, so called because they are present in many of the well-known fats and oils.

In the soap industry, the fats from both animals and vegetables can be used. Those in most common use are the fats from cattle and sheep and the oils from coconuts, palms and olives.

Before passing on to a description of the method of manufacture, there are one or two simple experiments

that you can work. These will help you to understand the processes used when fats are converted into soaps.

Experiment 7

To a small quantity of pure lard add caustic soda solution. Warm the mixture in a test tube.

- (a) Does the fat dissolve?
- (b) Shake the tube to find if a lather is formed.

Experiment 8

Repeat the previous experiment, using olive oil.

Experiment 9

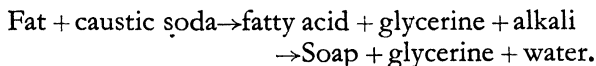
Make a strong solution of caustic soda solution in methylated spirit, and to it add a few pieces of lard.

Warm until all the lard dissolves and then boil off the spirit.

Mix the residue (if solid) with water and add strong salt solution. When cold, remove the solid on the surface and mould it into a small tablet.

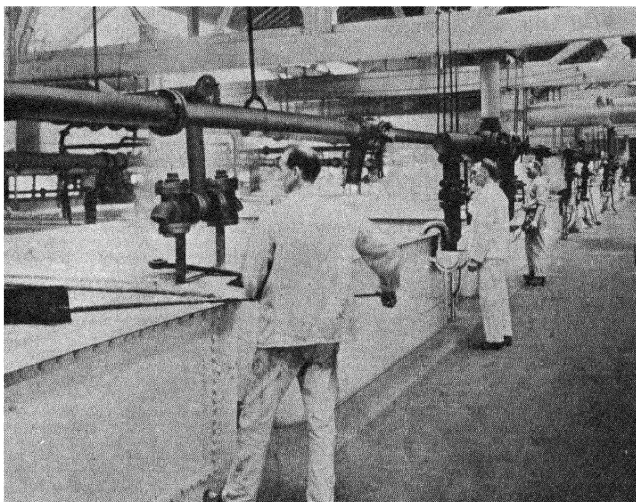
Test your tablet to find whether it will form a lather.

In the factory, the fat is placed in very large pans and heated with a sufficient quantity of caustic soda by steam (see picture on next page). During this process, the fat splits up into fatty acid and glycerine and the acid is neutralised by the alkali forming soap and water.



When salt is added, the glycerine settles at the bottom and is drawn off when kitchen soap is desired. For toilet soaps, some of the glycerine is left in the pans.

In order to still further purify the soap, the mixture in the pans is boiled and finally left to cool. Colouring matter and perfumes are added and the soap pressed into bars and cut into cakes and tablets.



THESE SOAP PANS HOLD AS MUCH AS 60 TONS
The oils, fats and alkalis pour in through overhead pipes
and are boiled by steam.

(Courtesy, Lever Bros)

An accurate balance between the quantities of fat and alkali is necessary. A soap containing too much fat would be greasy and unpleasant to use, while if too much alkali were present, it would irritate the skin and rot the clothes.

By dissolving the soap in methylated spirit, or by leaving a sufficient quantity of glycerine in the vats,

transparent soap is obtained, while by not using salt, the familiar soft soap or liquid soap is produced.

The glycerine obtained as a by-product in the soap industry is purified and used medicinally and for the manufacture of explosives.

Cellulose

To-day most of our cheap paper is made from wood pulp. By one process, soft woods are cut into chips and heated or *cooked* in a solution of caustic soda and the resulting pulp treated with china clay to fill up the pores. After drying, rolling and pressing, the pulp is changed into sheets of white paper.

The quality of the paper depends upon the materials used. For very fine papers, cotton rags are used, while for the cheap paper for newspapers, wood, straw and grasses are often employed.

The use of wood pulp for the manufacture of paper is only possible because of the structure of all forms of plant life. All plants are constantly producing a very wonderful substance which chemists call *cellulose*. Here, all that can be said about cellulose is that it is one of the many carbon compounds, containing carbon, hydrogen and oxygen. In plants it goes to the building up of the structure of the millions of tiny cells of which the plant consists.

Cellulose is one of the most useful substances we possess. In addition to its use in the manufacture of paper, it is present in large quantities in linen, and is used for the manufacture of paints.

By treating it with nitric and sulphuric acids, the violent explosive, *gun cotton*, is produced, but by the use of a

smaller proportion of nitric acid *collodion* can be made. This substance is of great value and is sometimes used in the manufacture of artificial silk, when it is „squirted through tiny holes and emerges as threads of silky material.

By mixing collodion with camphor and alcohol, *celluloid* is obtained, and this is used to a great extent for the manufacture of photographic films. Its inflammable nature makes it a dangerous material to handle, and cinematograph films are now made from cellulose acetate, obtained by treating cellulose with acetic acid. This material is very similar in appearance to celluloid, but it does not burn.

SUMMARY OF THE CHAPTER

Things to Note

1. The test for an alkali.
2. The reactions between acids and alkalis.
3. The meaning of neutralisation.
4. The methods of manufacture of caustic soda and ammonia with their uses.
5. The action between fats and the alkalis and the manufacture of soap.
6. The uses of cellulose.

CHAPTER 10

MORE ABOUT SALTS AND THEIR USES

The Formation of Salts

In the two preceding chapters you have learnt something about the properties of the two groups of substances—*acids* and *alkalis*. These have been distinguished chiefly by their effects on litmus paper. This test, however, is not really conclusive, for there are some substances which turn blue litmus red, but which are not classed as acids, and there are some acids which do not affect litmus at all. The same applies to the alkalis, and if you wish to be more accurate in your descriptions of these groups, you must learn a little more about each.

Experiment 1

Examine a specimen of zinc oxide. Test a small quantity to see if it will dissolve in water.

To a little of the oxide, add dilute sulphuric acid. Add more oxide until a little is left in the bottom of the tube. Pour off the liquid and allow it to evaporate. Examine the crystals formed.

Experiment 2

To a little copper hydroxide add some dilute sulphuric acid. Warm the mixture until the solid disappears.

Then add a little more and repeat until a small quantity of solid remains at the bottom of the tube.

Pour off the blue liquid and allow it to evaporate.

Experiment 3

To some dilute caustic soda solution add dilute sulphuric acid until the liquid just turns blue litmus red.

Allow the liquid to evaporate.

In these experiments you have prepared specimens of salts called *sulphates*. In the first, zinc oxide neutralises the acid to form white zinc sulphate and water.

i.e. Zinc oxide + sulphuric acid \rightarrow zinc sulphate + water.

In the next experiment, copper hydroxide changed to blue copper sulphate and water.

i.e. Copper hydroxide + sulphuric acid
 \rightarrow copper sulphate + water,

while in the third, caustic soda (or sodium hydroxide) neutralises the acid forming sodium sulphate and water.

i.e. Sodium hydroxide + sulphuric acid
 \rightarrow sodium sulphate + water.

By similar experiments, using other acids, more simple salts may be made. Thus, sodium nitrate may be prepared by neutralising caustic soda with nitric acid, and potassium nitrate by using caustic potash instead of caustic soda. These nitrates are of great importance; sodium nitrate found in large quantities in Chili as saltpetre is used extensively as a garden fertiliser. Potassium nitrate (saltpetre or nitre) which is also found naturally in hot countries, is used chiefly in the manu-

facture of gunpowder, as it is a very powerful oxidising agent. It differs from Chili saltpetre since it does not take water from the air when left exposed.

By using hydrochloric acid, *chlorides* are formed. With caustic soda, sodium chloride or salt is produced; with lime (calcium oxide), calcium chloride, a very useful drying agent, is formed. Copper hydroxide gives copper chloride; zinc oxide—zinc chloride, etc.

Hydroxides

Now you are familiar with the term oxide, but you will be wondering about hydroxides. The word itself gives you a clue and suggests that there is hydrogen present. The hydroxides are all formed from water by replacing part of the hydrogen by a metal. Thus, sodium hydroxide is formed when sodium meets water,

i.e. Sodium + water → sodium hydroxide + hydrogen,
and calcium hydroxide when calcium oxide (lime) reacts with water,

i.e. calcium oxide + water → calcium hydroxide,
but this method does not always apply and many hydroxides have to be made in other ways and usually sodium hydroxide is employed.

e.g. Magnesium salts + sodium hydroxide
→ magnesium hydroxide + sodium salts,
and Copper salts + sodium hydroxide
→ copper hydroxide + sodium salts,

and in these equations, as you can see, the sodium in the hydroxide has been replaced by another metal.

Experiment 4

To a solution of zinc sulphate add a few drops of caustic soda solution. "

Describe what happens.

Experiment 5

Repeat Experiment 4, using copper sulphate solution and caustic soda.

Repeat, using magnesium sulphate solution and caustic soda.

Here you have prepared hydroxides by the use of caustic soda, and in each experiment the hydroxide has been insoluble in water and has separated from the liquid when the caustic soda is added.

In your experiments, 1 and 2, zinc oxide and copper hydroxide have acted in exactly the same way as alkalis, but this term is reserved chiefly for the hydroxides of sodium and potassium, generally called the alkaline metals. The more general name of *base* is given to all those substances which have the power of neutralising acids producing salts and water, and our alkalis of the previous chapter must be included in this more general group.

Bases are generally the oxides or hydroxides of metals, although ammonia, which is a strong base, is an exception.

The Salts of Silver

Silver occurs in nature mainly in combination with sulphur as *silver glance*, but a chloride called *horn silver* is also found. The metal itself is not affected by hydrochloric acid, but neutralises sulphuric acid to form silver

sulphate. This salt is only slightly soluble in water. With nitric acid the soluble silver salt—silver nitrate—is produced. This substance crystallises in colourless plates, but on exposure to light it darkens and becomes almost black in colour. It blackens the skin and is used chiefly as an ingredient in marking ink.

Experiment 6

To a freshly made solution of silver nitrate add a solution of sodium chloride. Allow the tube to remain in the light and observe the colour changes which take place.

Experiment 7

Repeat the experiment, using a solution of potassium bromide in place of the salt solution.

Again observe the colour changes.

N.B.—Bromides are salts of hydrobromic acid, just as chlorides are salts of hydrochloric acid.

In your previous experiments you found that acids and bases react together to form salts, and here you have seen that some salts also react to form other salts. Thus, in Experiment 6—when sodium chloride solution is added to silver nitrate—a white, curdy solid is formed. This is silver chloride and the liquid remaining in the tube is a solution of sodium nitrate.

i.e. Silver nitrate (soluble) + sodium chloride (soluble)
 → silver chloride (insoluble) + sodium nitrate (soluble).

A similar change takes place in Experiment 7, and by the action, silver bromide and potassium nitrate are formed.

i.e. silver nitrate (soluble) + potassium bromide (soluble)
 → silver bromide (insoluble) + potassium nitrate (soluble).

These changes are made evident only by the insolubility of the silver chloride and bromide, and if these were soluble, no change would be visible.

The changes in colour in both experiments suggests that the action does not cease with this simple interchange. The silver chloride, at first white, soon becomes darker on exposure to light, and after a while is almost black. Silver bromide, too, acts in a similar manner and changes in colour from yellow to black. This is one of the most important and useful properties of the silver salts, and on it is built the foundations of photography.

The Chemistry of Photography

The construction of the camera is fully explained in Book I (page 112) and you have studied the purpose of the various parts. It will be advisable to make quite sure of this work before proceeding to the construction and working of the photographic plate or film.

The film in your camera is made from celluloid and is coated on one side with an emulsion containing gelatine and silver bromide. This has to be done in the dark, or rather, in the presence of red light, as it has been found that light of this colour does not change the silver salt in any way (see picture on next page). Ordinary white light, as you have seen, does affect the silver salt, and films or plates have to be well wrapped with black paper to make them *light tight*.

The inside of your camera is dark when the shutter is closed and the film does not change. When the shutter is opened for an instant, light enters and, on meeting the film, causes colour changes in the silver bromide. Now you have learnt that we see objects by the light reflected

from them, and that some appear very bright because they reflect more light than those which seem dull or dark. Thus, when taking a photograph, the brightest parts of the object cause the greatest changes in the colour of the silver salt on the sensitive film. The darker parts of the object and the shadows reflect very little light and cause very little change.



COATING PHOTOGRAPHIC PLATES IN A DARK ROOM
Ilford Infra-Red By courtesy of Ilford Ltd.

A film or plate after exposure to light in a camera appears to the eye to be unchanged, it is still of its original cream colour, but by the aid of still more chemicals it is possible to make the photograph visible or, to use the more common term—to *develop* the film.

When a film is developed

When the exposed film is developed, it is treated in a dark room with a *reducing agent* (see page 17), one of the

cheapest being a mixture of soda and pyrogallic acid in water. This liquid continues the changes that have already taken place and very finely divided silver is deposited. Where the light was most intense, there the deposit will be greatest, while those parts of the film not exposed to light will be unaffected by the developer. Thus, on the film, the brightest parts of the object are dark in colour while the darkest parts are light. For this reason the image on the film is called the *negative*, being directly opposite to the object itself.

The process of developing does not end here, for it is necessary to remove all the remaining silver bromide, or this would change when the negative was exposed to the light. This *fixing* of the image, as it is called, is accomplished by the use of a chemical called sodium thio-sulphate, more generally known as *hypo*. A solution of this salt easily dissolves silver bromide, and the photographic plate or film is dipped into it to remove all traces of silver bromide still left on it. The image on the negative is thus *fixed*, and no further changes will occur when it is exposed to the light.

Making a photograph from a negative

Some printing papers are similar in construction to the film, that is, they are made from paper coated with a mixture of gelatine and silver bromide. When a piece of sensitive printing paper is exposed behind a negative, a positive image is formed on the paper, since the darker parts of the negative transmit only a little light, while the lighter parts allow light to pass more easily. Thus, the image on the paper corresponds exactly to the object and

to make it permanent it must be developed and fixed as if it were a film.

Another type of paper, which is more popular as it does not have to be developed, is made by using silver chloride instead of the bromide. On exposure to light behind a negative, the silver chloride is changed to what we call the *photo-chlorides*, and these vary in colour according to the intensity of the light. No development is necessary, but the photograph has to be fixed by immersion in hypo solution.

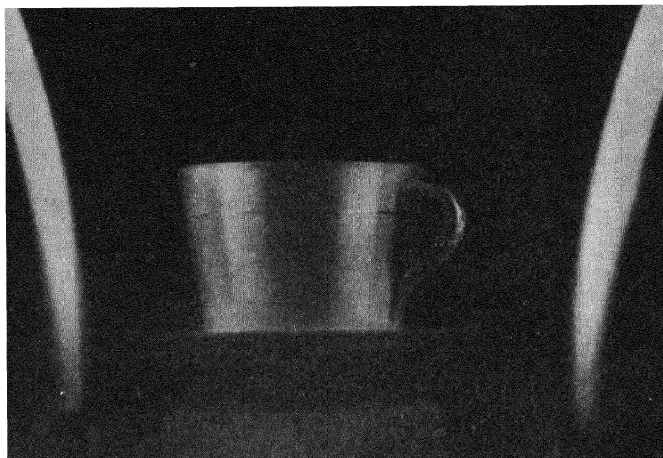
The colour of the photograph using this type of paper is not pleasing, and to tone the print, it is dipped into a solution of gold chloride. Some of the silver is replaced by gold and a rich, dark brown colour is obtained. By the use of a platinum salt in place of the gold chloride, a dark grey image is obtained which is often more popular than the brown.

Photographic paper often contains these substances in the sensitised coating, and no dipping in gold or platinum solution is necessary. Such paper is said to be *self toning* and is very popular, since it makes the process of photography much more simple.

Infra red Photography

When using the ordinary photographic plate it is necessary to have a good light, and photographs taken on a dull day are often poor, unless the shutter has been left open for the correct interval of time. To-day it is possible to take photographs in darkness by the use of specially prepared plates, but this is a modern development that was only discovered after a great deal of experiment. The earliest photographic plates were sensitive

only to blue light, and the colours in an object were not reproduced in correct shades on the final picture. The reds came out almost black while blue was nearly white, but by mixing certain dyes and chemicals with the silver salts, it was found possible to make the film sensitive to light of all colours. Such films are called *panchromatic*,



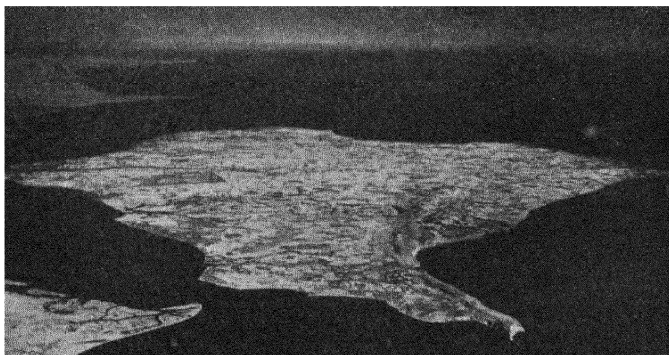
A CUP AND SAUCER MADE VISIBLE BY THE HEAT RADIATION FROM TWO ELECTRIC IRONS

Ilford Infra-Red Photograph By courtesy of Ilford Ltd

and when used produce photographs with the colours represented by varying degrees of light and shade in monochrome (single colour).

Experiment did not stop here and scientists, by the use of certain chemicals, found that they could make a photographic plate which, in addition to being sensitive to white light, was also sensitive to the invisible red rays or heat rays (Book I, p. 121).

Now a piece of iron when about to melt appears visible in a dark room and gives out white light. We say that it is white-hot, and if it is allowed to cool, the white light gradually changes to red and the iron emits, or gives out, rays of red light. On cooling still more, the iron becomes invisible to the eye, but still gives out rays, and these we know as the *infra red* rays. The modern infra red plate is sensitive to these invisible rays and by its use



THE ISLE OF WIGHT
From a height of 15,000 feet.
Ilford Infra-Red By courtesy of Ilford Ltd.

photographs may now be taken in darkness. The picture on page 126 shows quite clearly that these heat rays make nearby objects visible to this new type of photograph plate.

As the infra red plates are sensitive to light of all colours, special measures must be taken to exclude all but the infra red rays when a photograph is to be taken. To do this a special *filter* is fitted over the lens. To the human

eye this appears to be an opaque glass disc, but such a filter, while stopping the passage of the visible light rays, allows the infra red rays to pass through.

Now for the uses of these special plates. In the past it has been possible to take photographs only on a clear day, but as the infra red rays are not absorbed or dispersed by mist, it is now quite easy to take good, clear photographs, even in a thick mist. The pictures in Book I on page 122 show the prints obtained by the use of an ordinary panchromatic plate and an infra red plate on a misty morning. Both were taken at the same time, and a mere glance will show the remarkable difference in the results obtained.

Another great use is in the taking of long distance photographs. With ordinary plates, the mist and haze present in the air made a long range out of the question, but the infra red plate has added many miles to the range of the camera. On page 127 is shown an aerial picture of the Isle of Wight, taken from a height of nearly three miles, and when enlarged by the use of a magnifying glass you will be amazed at the detail it contains. It is now possible to take photographs in Dover and see on the prints the houses on the coast of France, and who can say what marvels of science will follow to make this art of photography even more wonderful.

SUMMARY OF THE CHAPTER

Things to Note

1. The characteristics of bases and their uses in making salts.
2. Chemical action between salts.

3. The light properties of silver salts.
4. How the image is produced on the photographic plate.
5. The need and purpose of developing and fixing a negative.
6. How the negative is made to give a positive photograph.
7. The use of the infra red rays in photography.

CHAPTER II

CHEMISTRY IN THE GARDEN

In spite of the many huge factories which are to be found in almost every country of the world, agriculture is still the world's greatest industry. Upon the farmer the many millions of the world's people depend for their food, for almost everything we eat comes from either plant or animal.

Science has helped the farmer just as it has helped the manufacturer. It has shown him how to produce new and better varieties of crops and animals, how to use machines to lighten his labour, and how to obtain from his land, in abundance, crops of the best variety. It will be the last of these that we are concerned with here, for the work of the chemist has done much to increase and improve the production of crops.

Before entering upon a study of the methods by which the growth of plants can be improved, it will be necessary to say a word or two about the soil in which they grow.

What is Soil ?

The soil that we find all over the land surface of the earth, was once hard, solid rock. By the action of great heat, rain, snow, wind and frost, the surfaces of rocks were worn down and broken up into tiny particles of dust. Even the air helped as the oxygen and carbon dioxide in

it joined up with certain of the substances present to form new compounds in the form of powders.

The dust and powder from the rocks were blown about by the wind or washed away by the rain, and settled in our valleys and on the hillsides as a layer of soft soil, steadily increasing in depth as the years passed by.

In addition to this broken up rock, there is in all soil the products of decayed animal and vegetable matter. This is generally called *humus*, and the amount present varies considerably with the district. Thus, in the Sahara Desert it is almost entirely absent and the soil is sand, while in the lowlands of England humus is abundant and the soil is darker in colour, heavier and more compact.

Generally, soils may be classed in one of three groups :

(a) *Sandy soils*—containing a large proportion of sand (or *silica*) together with humus.

(b) *Clay soils*—containing a compound of silicon and aluminium together with humus.

(c) *Chalk soils*—containing over 10% of chalk.

Each group has its own special properties and advantages. Thus sandy soil is light and easily worked or dug, but often lacks sufficient humus. It dries quickly and is easily moved by the wind, leaving the roots of plants uncovered and exposed to the weather.

Clay soil, on the other hand, usually contains a large proportion of humus. It retains or holds water, but, when wet, becomes very sticky and is difficult to work. When it dries there is a tendency for it to bake hard and crack.

The presence of chalk or limestone is found to have a great effect on the growth of plants. There are some, such as the foxglove, which will not grow in a soil which

contains a large proportion of chalk. On the other hand, many plants thrive in it.

For agricultural purposes it has been found that the best soil contains enough clay and humus to make it retain moisture, sufficient sand to make it easily worked, and a proportion of limestone to prevent it from becoming sour.

The work of the Soil

The soil plays a part of the greatest importance in the growth of plants. It has to supply the necessary water, to allow air to reach the roots, and it has to supply the plant with the greater proportion of its food. Now you can understand the disadvantage of very light sandy soils and of very heavy clay soils. The former, while they admit air very readily, will not retain moisture, and during periods of drought the plants often die through water shortage. The latter, although they retain water, do not easily admit air, and the growth of the plant is greatly retarded.

Without food no plant can live, and for the most complete growth, the soil must contain all the necessary plant foods in correct proportion. Now what are these plant foods? By examining specimens of many different plants, chemists have found they contain, or are built up of, a very large number of different substances. All contain nitrogen, carbon, oxygen, hydrogen, sulphur, potassium, phosphorus, calcium, magnesium, sodium and iron. Of these, only the first three are obtainable from the air and the remainder have to be supplied from the soil.

When the vegetation is natural, these substances are replaced in the soil when the plants die and decay, but

with cultivated plants, which are removed for food, some other means of replacement has to be found. This is accomplished by the use of *manures* and *fertilisers*.

Manures and Fertilisers

Of all the substances given in the previous section, nitrogen seems to rank as one of the most important foods. When it is available in plentiful supplies, growth is vigorous, while if the supply is limited, growth is slow and stunted.

Experiment 1

Prepare solutions as follows :

Solution A

Magnesium sulphate	-	-	·25 gr.
Ferrous phosphate	-	-	·6 gr.
Potassium nitrate	-	-	1·0 gr.
Calcium sulphate	-	-	·25 gr.
Water	-	-	1·5 litres

Solution B

Magnesium sulphate	-	-	·25 gr.
Ferrous phosphate	-	-	·6 gr.
Potassium chloride	-	-	1·0 gr.
Calcium sulphate	-	-	·25 gr.
Water	-	-	1·5 litres

Experiment 2

Select a number of barley seeds and allow these to germinate on blotting paper. From the seedlings choose two of approximately equal size and growth.

Experiment 3

Into one bottle or test tube pour solution *A*, and into another solution *B*. Fit each with a cork containing a piece of glass tubing and a hole for the seedling (Fig. 36).

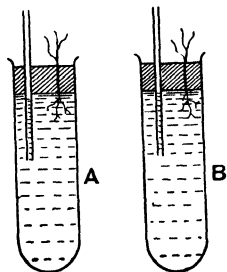


FIG. 36.

Place one seedling with its roots in solution *A* and the other with its roots in *B*, and place the bottles in earth. Set them aside and pump air into the liquids at intervals.

Observe and record the growth of each seedling.

N.B.—These solutions or *water cultures* will be met again in Book 3, where plants are studied in greater detail.

The experiments here are designed to show you the great importance of nitrogen. Solutions *A* and *B* are identical, except that *A* contains nitrogen in the form of potassium nitrate, while in solution *B* it is entirely absent. The great difference in the growth of the seedlings will be very clear. With nitrogen present growth is vigorous and the seedling in solution *A* will be strong and healthy, while that grown in *B* will be weak and stunted.

In the pictures on page 135 are shown the results of an experiment on apple trees. Those in *A* were not supplied with any additional manure while those shown in *B* were given supplies of nitrogen and potash. Study these and notice the great difference in the size of the trees.

Similar experiments may be worked to discover the effect of the absence of other substances—phosphorus, magnesium, calcium, iron, potassium and sulphur, but



A. - NO MANURI

Photo G. A. Cowie



B. - NITROGEN AND POTASH

Photo G. A. Cowie

the results for some of these will not be as clearly marked as those from the experiment you have already worked.

The majority of plants depend upon the soil for their supplies of nitrogen, and although four-fifths of the air is free nitrogen, only a few plants have the power of extracting this for use. The roots of peas, beans, clover and alfalfa are invaded by small rod-like bacteria, and these have the power of extracting nitrogen from the air. At first, these bacteria live on the food supplies of the plant, but they are absorbed by it during the later stages of growth and their supplies of nitrogen used up by the plant. Such crops, in themselves, form valuable manure and they are often grown, not for actual use as food, but to be dug into the ground and so enrich it with their supplies of nitrogen.

During the decay of both animal and vegetable matter, the gas ammonia, already met in Chapter 9, is produced. This gas is converted into an acid, nitric acid, by types of bacteria present in the decaying matter, and the acid combines with the bases present in the soil to form nitrates. This process, by which ammonia is converted into nitrates, is given the name *nitrification*, and the bacteria causing it are known as *nitrifying bacteria*.

When soil is repeatedly cultivated, the supply of nitrogen and the other essential substances is gradually exhausted, and if crops are still to be grown it is necessary to replace them by the use of suitable fertilisers and manures. For replacing nitrogen, one of the best known fertilisers is sodium nitrate (Chili saltpetre) found in the nitrate deposits of Chili. Another is sulphate of ammonia, produced as a by-product in the gas industry, while yet another is calcium nitrate.

The composition of some of the common manures is given in the following table, and the importance of the correct use of fertilisers will be appreciated on reference to the pictures on page 135, and the summary of the weights of the crops obtained (below).

COMPOSITION OF COMMON FERTILISERS

Name	Nitrogen	Phosphorus	Potash
	per cent.	per cent.	per cent.
Guano - - -	12	22	2.5
Basic Slag - - -	—	30	—
Bone Meal - - -	4	45	—
Fish Manure - - -	8	14	1.5
Sulphate of Potash -	—	—	49
Sulphate of Ammonia -	21	—	—
Superphosphate - -	—	35	—
Hop Manure - - -	1.5	5	1

Effect of Manure on Growth of Apple Trees

	No Manure	Potash only	Nitrogen and Potash
Weight of crop	105 lb.	207 lb.	256 lb.
No. of apples -	445	803	980

How the plant obtains its carbon

So far we have dealt only with the soil and the part it plays in the growth of the plant, but before leaving this chapter, some mention must be made of the great part played by both the air and sunlight.

Now chemical analysis of plants shows that they contain a large proportion of carbon in the form of starch and sugar. This carbon is obtained from the air, where it is present in the form of carbon dioxide. As you have learnt, this gas is breathed out by both human beings and animals, and it is a product of combustion when any form of carbon is burnt.

The carbon dioxide from the air enters the tissues of the plant by way of the leaves and there meets with water which has found its way to the leaves from the roots. Now under ordinary conditions carbon dioxide does not combine with water, although it dissolves in it. In the plant, however, there is a green colouring matter known by the name of *chlorophyll*, and this has the power of absorbing some of the sunlight falling on it. Some of the energy of the sunlight, for light is a form of energy just as heat and electricity are, is *transformed* (Book I, p. 57) and used to produce chemical action between the carbon dioxide and the water. As a result of this action, sugar is formed and oxygen is given off.

Experiment 4

Take a fresh leaf and dip in into boiling water.

Now soak the leaf in alcohol to remove the chlorophyll which dissolves.

Place the leaf in a weak solution of iodine and notice the coloration produced.

The blue coloration formed in this experiment is a very clear indication of the presence of starch and the action that has taken place may be represented by the following equation :

Carbon dioxide + water \rightarrow sugar (or starch) + oxygen.

Some of the sugar so formed is used by the plant for growth, while the remainder is stored. Many vegetables contain valuable stores of sugar in its various forms; carrots contain *glucose*, while sugar beet and sugar cane contain what we know as *cane sugar*. From the last two we obtain almost all our supplies of sugar.

There is one remarkable fact about this chemical action and that is that the volume of oxygen produced is exactly equal to the volume of carbon dioxide absorbed. This is of great importance, and because of it, the proportion of oxygen in the air does not change. In spite of the fact that we are continually using it and changing it into carbon dioxide, a perfect balance is obtained by nature, by means of the plants and trees which use carbon dioxide as one of their main foods.

Daylight is necessary for this absorption of carbon dioxide, since it supplies the necessary energy for the chemical action, and thus, during the night there is no production of sugar and carbon dioxide is not absorbed.

SUMMARY OF THE CHAPTER

Things to Note

1. The weathering of rocks to form soil.
2. Types of soil and their advantages.
3. The substances necessary for plant growth.
4. The necessity for both air and water.
5. The methods used for the provision of nitrogen.
6. Why is it necessary to manure cultivated ground.
7. The formation of starch in the tissues of plants.

CHAPTER 12

THE FOODS WE EAT

MANY of the applications of science to our lives you have already met. In this volume alone, you have studied coal gas and its use for heating and lighting our homes. You have learnt something of the manufacture of such common commodities as a bar of soap and a box of matches, but what of that large variety of substances we call foods ?

Why the human body needs food

In the previous volume you have made a brief study of some of the common engines ; you met the steam engine, the turbine, the petrol engine and the diesel or heavy oil engine. All of these have to be supplied with fuel of some kind, and the energy contained in this fuel is used to do mechanical work. In the motor-car, for example, the energy contained in the mixture of petrol and air is transformed into work when the car moves along the roadway. In the locomotive, heat energy from coal is used to change water into steam, and the energy stored in the steam is transformed into work when the engine pulls its load along the track. The human body is often called an engine, but if we are to use this term, we must remember that it is far more delicate in its construction, and that its performance is far more wonderful than that of any man-made engine. It requires fuel from which to obtain its

necessary energy, and here is one of the reasons why we must eat food, for the things we eat represent stored up energy, and this is released and used to keep us warm and to enable us to move our limbs.

The human body, while it resembles a mechanical engine in this use of fuel, differs from it in many other respects. No motorist expects his motor-car to grow in size or increase in weight, nor does he expect it to do its own repairs. He finds that as the age of the car increases, he has to replace worn parts if he wishes it to give good service. However, the human body grows in both size and weight, and repairs are being carried out continuously. Here again, food is the most important factor and so far we can list the uses of food as :

1. To supply the necessary warmth and energy.
2. To add to the bones and tissues to produce growth.
3. To repair or replace worn out tissues.

Types of foods

It is evident that if our bodies are to be well nourished, then the food we eat must supply all the necessary needs listed above. No one article of food will do this and it is necessary for us to eat foods of different kinds. In addition to these, water is essential, for life would not be possible without it. While the bulk of the supply is obtained through the liquids we drink, some is supplied by fruits and vegetables as well as by eggs, fish and meat, all of which contain a proportion of water. Even now the list is not complete, for within recent years it has been discovered that other very complex substances, called *vitamins*, are necessary if our bodies are to be healthy.

These are of different kinds and are present in tiny quantities in many types of food, and it is believed that they may also be produced, to some extent, by sunlight.

Foods which build the body

Two types of food go to the building of the body—those we call *proteins* and certain of the *mineral salts*. The proteins we obtain from meat, eggs, fish, cheese and certain cereals. Actually, we depend upon the plants for our supplies as the animals from which we obtain our meat, eggs and cheese all live on plants, while fish, even though some live by feeding on their smaller neighbours, depend upon the tiny plants and animals found in the rivers, lakes and seas.

These proteins are complex chemical compounds built up mainly from carbon, hydrogen, oxygen and nitrogen. There are various tests for their presence and two are given for you to try.

Experiment 1

To half a test tube of distilled water add a small quantity of pea meal. Shake well and then allow it to stand for some hours. When the sediment has settled, pour off the liquid into another tube.

Experiment 2

To a part of the liquid obtained in Experiment 1, add a little strong caustic soda solution and just a drop or two of copper sulphate.

Note the colour changes.

Experiment 3

Powder two or three peas and mix them with a small quantity of lime. Heat the mixture in a test tube and test the gas which is given off by its smell to find if it is ammonia. Also test by a piece of moist red litmus paper.

Now, while the human body is able to store up supplies of heat and energy giving foods, it cannot store proteins. Moreover, they are difficult to break up, and a diet containing protein in excess may be just as harmful as one where they are insufficient.

Certain mineral salts are also necessary for growth, for it has been found that the bones and teeth contain calcium and that the blood contains iron. Sodium, potassium and phosphorus are also essential, and the absence of the salts of these substances from our diet will result in what are sometimes called *deficiency* diseases. Thus a diet lacking calcium will result in weak bones and poor teeth. It is far better to take these substances in our food rather than rely upon medicines for our supply, for by including them in our diet, we are sure of having them regularly in the small quantities that can be dealt with by the mechanism of the body.

Foods containing mineral salts include milk, fruit, vegetables and some kinds of fish. Thus milk and parsnips contain calcium and potassium salts, cabbage and raisins contain iron salts, peas contain phosphorus and potatoes contain potassium salts. In addition, of course, there is common salt, used to flavour our food, and this, as you know, contains sodium.

The value of mineral salts is not confined to body building, for some are necessary to keep the internal

machinery in good working order. They help in the digestion of food and in the removal of the waste material.

Foods which supply heat and energy

These foods may be split into two classes—those which contain fat and those containing starch or sugar. Everyone has seen fat and most of you will have had the opportunity of examining starch and sugar. They appear very different in form and yet, chemically, they are very similar. In fact, fat, whether it is contained in plant or animal, is made from starch. You have seen in the previous chapter that the leaves of plants contain starch, made from the carbon dioxide of the air and water under the energy of light. While in some plants the starch remains unchanged, in others it is changed into fat or oil and stored in the seeds or fruit. Animals, too, manufacture their supplies of fat from starch contained in the foods they eat. Part of the fat is consumed to provide the necessary supply of energy while the rest is stored in the tissues as natural fat.

When fat is *burnt* in our bodies, energy is provided in the form of heat, part of which is used in keeping our bodies warm. The remainder is transformed into mechanical energy and does work during the movement of our limbs and muscles. The products of this burning of fat are carbon dioxide and water, the former being passed out of the body by way of the lungs, or through the pores while the water is used in the body to aid digestion.

Additional supplies of energy come from starch and sugar, and while you accept the latter as an article of food, you may have some doubts about starch. Yet starch and

sugar are very closely related, in—fact, they contain the same substances, carbon, hydrogen and oxygen, but in slightly different proportions. Now the body cannot use starch in its natural form as it is insoluble in cold water, but the saliva and the gastric juices react with it and change it into a form of sugar which is then used to provide energy. Surplus supplies of sugar are changed into fat and stored as reserves to be called upon in case of need.

A simple test for starch has been given on page 138, and you will need this for the following experiments.

Experiment 4

Powder a small lump of starch and gently heat it on a clean tin lid. When it is hot, try to ignite it.

(a) What is left behind when the starch burns ?

Now heat the residue strongly and test for the presence of carbon dioxide, using a drop of lime water held in the end of a piece of glass tubing.

(b) What does this experiment show you about the composition of starch ?

Experiment 5

Repeat Experiment 4, using sugar.

Do the results of these two experiments agree ?

Experiment 6

Cut up a raw potato and make it into a pulp. Add water and stir thoroughly for some minutes. Pour off the liquid and add one or two drops of iodine solution.

(a) Does this test show the presence of starch ?

Experiment 7

Test a number of common foods to discover if they contain starch. Make a list of those in which starch is present.

We obtain our supplies of starch and sugar from bread, which is still the food in greatest use, from such vegetables as potatoes, from cereals such as rice, peas, beans and oatmeal, and from fruits, most of which are rich in sugar. Then, of course, we use sugar itself and such foods as treacle and honey.

The percentage composition of some of the common articles of diet is shown in the table on the next page. From this you will see the relative values of certain foods, and you will understand why they are so necessary if we are to give the body the correct materials to work with.

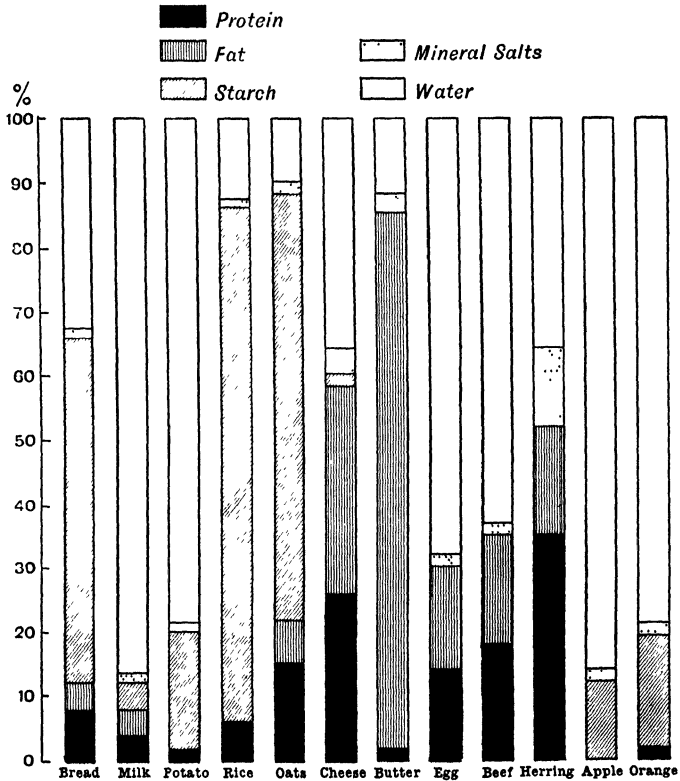
Vitamins

Before the coming of the steamship, in those days when sailors were away from land for considerable periods, a disease known as *scurvy* was very prevalent. Their food consisted chiefly of salted meat and dry biscuits, for fresh fruit and vegetables could not be carried, and it was the absence of these which caused the disease. In these fresh fruits and vegetables scientists have found a substance they call *vitamin C*, and it was the absence of this substance that gave rise to scurvy. The value of this vitamin is lost by boiling, so if we are to preserve good health, our diet must contain fresh raw fruit, of which oranges are probably the best, or uncooked vegetables such as lettuce.

In addition, scientists have discovered other vitamins. Vitamin A, present in green vegetables, in fish, in

animal fat and in milk, cheese and eggs, is necessary if we are to escape infection from disease. It is necessary for

HOW YOUR FOODS ARE BUILT UP



correct and healthy growth and hence is of extreme importance for children. Vitamin B is found in cereals and in wholemeal bread, and is probably present in the husk.

White bread and specially prepared cereals such as polished rice are lacking in vitamin B, as the husk has been removed during the refining processes.

This vitamin helps in the expulsion of waste matter from the body and strengthens the nervous system. Lastly, there is vitamin D, produced in the skin after exposure to natural sunlight, and also found in eggs and in oily fish. Vitamin D is essential for the correct development of bones and teeth, and its absence in children causes rickets. Cod liver oil and halibut liver oil are rich in this vitamin and also contain vitamin A, and these form valuable tonics for children.

Just one last word about food. You have met the various types and the work they do, but there is one thing more. If we are to be healthy, these classes of foods must be present in our diet in correct proportions, for excess of one form can never make up for a deficiency or lack of another.

SUMMARY OF THE CHAPTER

Things to Note

1. The foods necessary for health.
2. The place of carbon and nitrogen in diet.
3. The necessity for variety in our meals.
4. The types and importance of vitamins.

ONE HUNDRED QUESTIONS

THESE questions will test your knowledge of the work you have done and of the books you have read. When you find one that you cannot answer, make a point of revising the work and of hunting up the facts in your reference books.

1. What do you understand by the terms : *burning* and *rusting* ?

2. Give a full description of the methods in use for the prevention of rusting.

3. Describe an experiment to prove that when a substance burns in air, some other substance becomes joined to it.

4. It is often said that oxygen is essential for substances to burn. Give examples to show that this is not always true.

5. What substances are present in air ? How does the air in a town differ from that in the country ?

6. What scientists do you think of in connection with air ? Give a brief account of the work of those you mention.

7. Explain the difference between the terms *compound* and *mixture*. Give examples to illustrate your answer.

8. Why is it dangerous to hold a piece of phosphorus between the fingers ? What method is used when storing this substance ?

9. Name two substances with low ignition temperatures.

10. What do you understand by an unstable compound ? Give examples of such compounds.

11. Why are nitrogen compounds valuable for use in explosives ?

12. How would you test a gas to discover if it were carbon dioxide ?

13 Compare the characteristics of carbon monoxide and carbon dioxide.

14. How does Nature preserve the correct balance of oxygen in the atmosphere ?

15. Describe the method used for the manufacture of glass.

16. Why is it dangerous to run the engine of a motor-car in a garage with the doors shut ?

17. Give examples to show that the characteristics of a chemical compound differ considerably from those of the substances contained in it.

18. Describe an experiment by which water can be split up into its separate substances.

19. How would you prove that water contains only hydrogen and oxygen ?

20. What happens when hydrogen burns in air ?

21. What chemical action takes place when sodium meets water ?

22. Why is helium more suitable than hydrogen for use in airships ?

23. Why are substances often powdered before being dissolved ? Name two substances which are more soluble in cold water than in hot.

24. Explain what happens when a substance dissolves.

25. What is an *amalgam* ? Describe how zinc amalgam can be made.

26. Why are petrol and benzine useful liquids for removing dirt from clothes ?

27. When is a solution said to be *saturated* ? Describe how you would prepare a saturated solution of common salt.

28. When common soda is left in contact with air, it changes from white crystals to a powder. Describe the changes that cause this.

29. If you were given a few crystals of copper sulphate and an unknown liquid, how would you decide whether this liquid were water ?

30. Why is the water in some districts said to be hard ?

31. What is the difference between temporary and permanent hardness of water ?

32. What are the methods in use for the purification of our water supply ?

33. Describe some of the uses of *fractional distillation*.

34. Give a description of the formation of the coal seams in the earth. What are the main types of coal ?

35. What scientist is associated with safety in our coal mines. Describe some of his work.

36. Describe the characteristics and uses of the various forms of carbon.

37. Why is it essential to have good ventilation when using gas appliances ?

38. Why is charcoal often used in the construction of gas masks ?

39. Make a drawing to show the working parts of a bunsen burner.

40. Give in tabulated form a list of all the by-products of coal which you have met in your work.

41. Name some of the most useful hydrocarbons.

42. The chimney of a coal fire needs to be regularly swept while that of a gas fire does not. Explain this difference.

43. How is iron found in nature ? Describe how you imagine these compounds to have been formed.

44. Why is an air blast necessary in the blast furnace ? What is the advantage of making this hot ?

45. What do you understand by a *reducing agent* ? Describe the part played by (a) coke, (b) limestone in the blast furnace.

46. Name four alloys in common use. Describe how they can be made.

47. Describe the process by which iron can be converted into steel.

48. What is *mild* steel. How does this differ from *high speed* steel ?

49. What do you understand by the (a) malleability and (b) the ductility of a metal. Name one metal which is very malleable and another which is very ductile.

50. Why are blast furnaces kept working continuously ?

51. Explain how aluminium can be used to separate metals from the ores.

52. What is meant by *welding* ? What methods of welding are in common use ?

53. Describe, with diagrams, the methods used for the separation of copper from the ore. How is the crude copper refined ?

54. What methods of cleaning are used in the electroplating industry ?

55. Describe an experiment which shows how lead can be separated from lead oxide. What is the reducing agent used ?

56. Why is copper of such great value to the electrical industry ?

57. Name two alloys of tin and describe their uses.

58. Describe some of the uses for lead.

59. How would you test a liquid to discover if it were an acid ? Is your test conclusive ?

60. Why must great care be taken when using strong acids ?

61. What substances are produced when sulphuric acid reacts with magnesium ? How does the action differ when copper is used ?

62. What are the uses of sulphur dioxide ? Is this the only oxide of sulphur ?

63. Name and describe one of the methods of manufacture of sulphuric acid. What are some of the uses for this acid ?

64. What are *catalytic agents*? Describe two processes employing them.

65. What substance is present in all acids? Are all compounds containing this substance acids?

66. Describe some of the uses of chlorine.

67. Draw up a table showing the difference in characteristics of metals and non-metals. Name two substances which may be considered as border-line cases.

68. Describe the modern process for the manufacture of ammonia. What is the catalytic agent in use?

69. What are the commercial uses of ammonia?

70. Describe the chemical change which takes place when an acid is neutralised by an alkali.

71. What are the names of the salts obtained from (a) hydrochloric, (b) sulphuric, (c) nitric acids. Describe how you would prepare a specimen of potassium nitrate.

72. What would be the characteristics of a soap containing too much alkali?

73. Describe the stages in the manufacture of a bar of soap.

74. Write a short essay on the value of cellulose.

75. What is the difference between an oxide and an hydroxide?

76. What is the importance of the silver salts in photography?

77. Why is it necessary to reduce the size of the aperture of a camera in brilliant sunlight?

78. Why are photographic films wrapped in black paper?

79. What do you understand by the photographic terms: (a) Negative, (b) Self toning, (c) Panchromatic, (d) Over exposed?

80. When using an infra red plate it is necessary to cover the lens with a *filter*. Why is this?

81. Describe the uses to which the infra red plate may be put.
82. Explain the changes that take place when a solution of silver nitrate is added to a solution of sodium chloride.
83. Describe the process of *fixing* a photographic negative.
84. What is meant by the *weathering* of rocks ?
85. Compare the characteristics and advantages of a sandy soil and a clay soil.
86. What is the value of lime to the gardener ?
87. Describe the methods in use for replacing the supplies of plant food in the soil.
88. Describe a simple test for starch. How would you show that the leaves of a plant contain starch ?
89. Describe how a plant obtains its supplies of carbon.
90. Why is it necessary to dig the soil at regular intervals ?
91. In what ways does the human body resemble an engine ? How does it differ ?
92. What are the various uses to which the foods we eat are put ?
93. How would you test for the presence of (a) water, (b) starch, in a specimen of food ?
94. Write a short essay on the value of nitrogen.
95. What are deficiency diseases and how are they caused ? What is meant by a *balanced diet* ?
96. What are the classes of vitamins ? Name foods which are rich in (a) vitamin A, (b) vitamin D.
97. Why is milk considered as a complete food for young children ?
98. State what classes of foods are necessary for good health.
99. Draw up a suitable diet for one day. State those foods which supply the protein.
100. How does the science of Chemistry differ from that of Physics ?

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