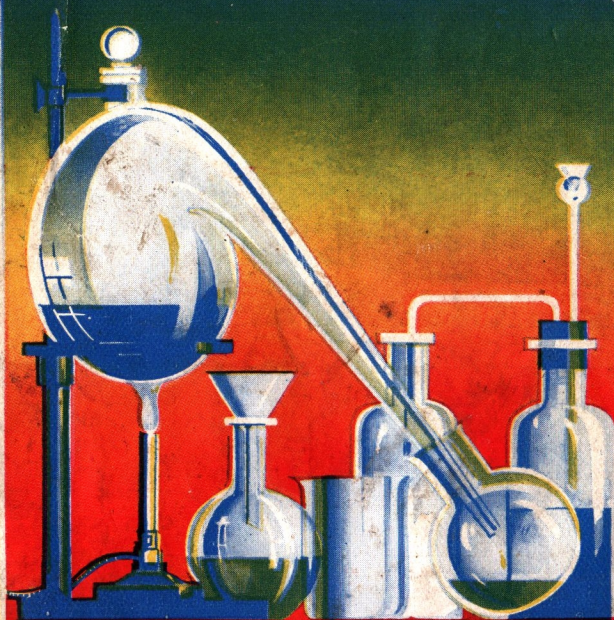


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REGISTERED

CHEMICAL EXPERIMENTS



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CHEMICAL EXPERIMENTS

Everything in the world around us is built of chemical elements and their compounds. These compounds undergo ceaseless changes, and the study of these changes forms the basis of the wonderful science of chemistry. The Kemex Outfits have been introduced to provide apparatus, materials and instructions for carrying out a series of fascinating experiments, in which the secrets of chemical science are revealed.

The No. 2 Outfit contains the necessary apparatus and chemicals to make a chemical garden that actually grows, to dye wool and silk, to make soap, and to smelt metals from their compounds.

With this Outfit 250 interesting experiments can be carried out.

The No. 3 Outfit is planned on a much wider scale, and with it from 350 to 400 experiments can be made. It covers the whole range of the No. 2 Outfit, and has additional apparatus and materials for a further series of experiments showing how chemistry is applied in the home and in the factory.

To get the greatest fun from your Kemex Outfit you should read the "Meccano Magazine," in which special articles link up with the Outfits and describe new and interesting experiments in all branches of chemistry.

Kemex Outfits for Chemical Fun!



Kemex Chemical Experiments

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1. THE BEGINNINGS OF CHEMISTRY

The earliest chemical experiments date back to prehistoric times, and were concerned with the crude operations by which primitive Man produced metals from their ores. The first metals to be noticed by Man were probably gold, silver and copper, which would attract his attention by reason of their bright colour. He would soon find that gold and silver were of little practical use on account of their softness, but that copper could be made into satisfactory weapons, tools and domestic utensils of many kinds. We do not know how or when the art of smelting copper was discovered, but it has been suggested that it came about in the later stages of what is known as the Stone Age, through the accidental use of copper ore stones in the fire-pits built for cooking purposes. Later came the discovery that copper could be hardened by the addition of tin, thus producing the alloy bronze. Possibly this discovery was made accidentally during the smelting of copper ores along with which were ores containing tin.

It is probable that copper and bronze were in use as far back as 10,000 years ago. Certain Egyptian relics of these metals are believed to date back 7,000 years,

and there is evidence that copper tools were made in Ireland some 2,500 years B.C. while in Great Britain bronze is thought to have been in use until 1,000 B.C.

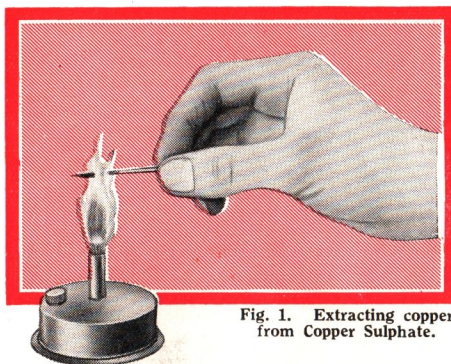


Fig. 1. Extracting copper from Copper Sulphate.

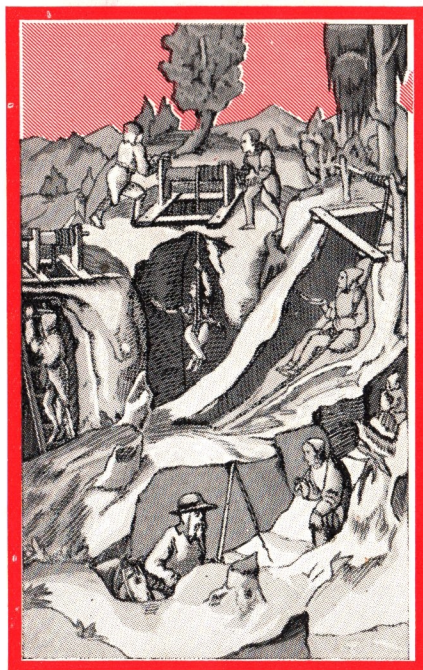


Fig. 2. Mediaeval miners descending shafts by means of ladders, ropes and steps cut in the rock.

Experiment in Copper Smelting

In view of the great antiquity of copper smelting it will be interesting to commence our chemical experiments with one of this nature. For this we require a piece of wood about 4 in. in length and about twice the thickness of a match. Hold the end in the flame of the Spirit Lamp (Part No. K22) until the wood begins to char, and then rub it with a large crystal of washing soda. Some of the soda is absorbed, and by repeating the action the end of the stick becomes transformed into a piece of charcoal thoroughly soaked in soda.

Crush a few small crystals of Copper Sulphate (No. K108); place as much as possible of the substance on the charred stick, and hold this in the flame (Fig. 1). The blue powder becomes white, and further colour changes take place as the match continues to burn, the mass becoming black with a reddish tint in places. The soda prevents the wood from burning away quickly, and also acts as a "flux." A flux is a substance used to help in separating metals from the other

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constituents of their ores; in this case the soda assists the materials on the wood to become liquid.

Stop the heating before the charred end of the wood burns away or falls off, and break the end off into the Evaporating Dish (Part No. K8), half filled with water. With the point of a penknife, or the end of a Glass Rod (Part No. K16), crush the charred end and stir it into the water, and then give the dish a circular movement in order to wash the remains. Add more water and continue the movement, tilting the dish slightly so that the water swirls out over the edge, this process being similar to that used by gold prospectors when washing gold-bearing gravel in a pan. The light charcoal is washed away, the remaining soda is dissolved, and finally reddish-brown particles are left at the bottom of the dish. These particles are flakes of copper smelted out of the Copper Sulphate. Most of the water remaining on the copper may be removed by blotting paper, and gentle warming of the dish completes the drying.

Analytical chemists make use of tests of this kind when they suspect an unknown substance to contain compounds of copper. They complete the test by adding dilute nitric acid drop by drop until the reddish-brown flakes are dissolved. This may be done before removing the last trace of water, and gentle warming may be necessary. A blue solution is formed, and the colour becomes far more intense if a few drops of ammonia are added to the liquid. This striking colour change is characteristic of copper and makes itself evident when only a minute proportion of copper is present.

Miniature Smelting Furnace

The copper-smelting experiment may be carried out more conveniently on the Charcoal Block (Part No. K21). Near one end of this scrape a round depression about $\frac{1}{2}$ in. in diameter and $\frac{1}{4}$ in. deep, and place in it a mixture of Copper Sulphate (No. K108) with an equal amount of washing soda, first crushing the materials together. The charcoal block is then held in the hand while the flame of the spirit lamp is directed on to it by means of the Blowpipe (Part No. K20) as shown in Fig. 4, thus forming a miniature smelting furnace.

The blowpipe is used partly to direct the flame in the direction required and partly to make it hotter by introducing more air. The nozzle should be slightly outside the flame of the lamp, and on the opposite side to that on which the charcoal block is held. It is not

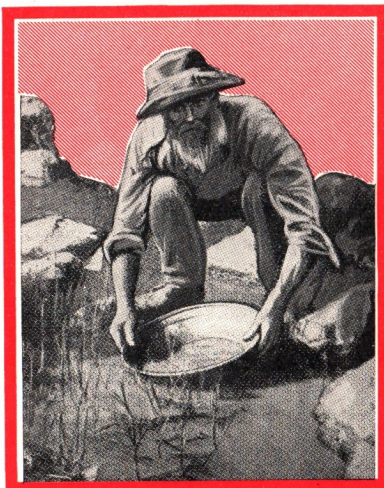


Fig. 3. A prospector at work, panning gold-bearing gravel.



Fig. 4. A miniature smelting furnace formed by scraping a depression on a Charcoal Block.

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necessary to blow hard, for if the flame is shielded from draughts it is easily directed on to the mixture of Copper Sulphate and washing soda, which at first should be held in the tip of the side flame produced, and afterwards brought nearer to the lamp.

The mixture in the hollow of the charcoal block melts and finally becomes a dry hard mass, which is scraped out into water in the porcelain dish and washed as already described.

Three Metals that are Magnetic

Other metals that may be extracted similarly from their compounds are iron, nickel,

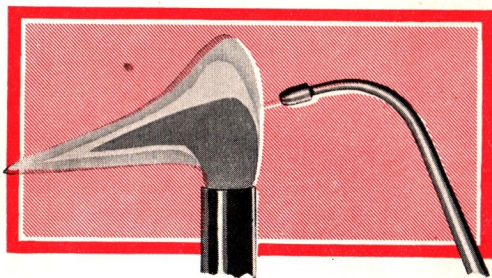


Fig. 5. When metals are being smelted on charcoal the nozzle of the blowpipe should be held just outside the flame.

and cobalt, the experiments being carried out with the Ferrous Ammonium Sulphate (No. K110), Nickel Ammonium Sulphate (No. K120) and Cobalt Chloride (No. K105) contained in the Outfit. These metals are obtained in the form of dark grey powders. There is no difficulty in recognising them, however, for they are magnetic, and in each experiment the dark grains left in the porcelain dish adhere to a magnet dipped into them.

If the experiment is repeated with Lead Nitrate (No. K113) the metal is obtained in the form of a small grey bead, for the temperature is high enough to melt the lead formed. This bead marks paper. In addition, a yellow stain forms on the charcoal block on the side furthest from the blowpipe. This is due to the burning of some of the lead, and it consists of a compound known as

litharge, or lead oxide. How it is formed will be explained later.

An interesting experiment may be made

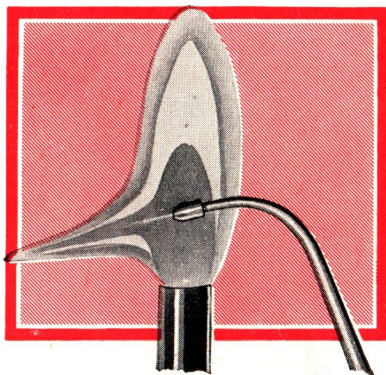


Fig. 6. Higher temperatures suitable for burning Zinc are obtained when the nozzle of the blowpipe is placed inside the flame.

with the Granulated Zinc (No. K134) included in the Outfit. Place a few small pieces in a hole scraped in the charcoal block, and heat these strongly with the tip of the blowpipe flame. The Zinc burns and is changed into a powder, part of which spreads over the surface of the block. This powder has the peculiar property of being yellow when hot and white when cold, and the colour change may be brought about repeatedly by alternately heating and cooling.

Since Zinc burns so readily when heated in the blowpipe flame, it cannot be obtained by heating its compounds with soda on a charcoal block. These compounds give zinc oxide as a powdery stain that is at first yellow and becomes white on cooling, and this is as useful a guide to the analyst as the production of the metal itself would be.

The Iron and Steel Industry

Wood was burned in the early furnaces in which metals were smelted, and this was the source of the charcoal required to bring about the reduction of the ore to the metal. When larger furnaces came into use, air was blown in through tubes in order to produce higher temperatures, exactly as in our experiments with the blowpipe. To this day the native ironworkers of Central Africa feed their furnaces with wood and iron ore, and fan the flames by means of bellows that drive air through bamboo tubing. The temperature

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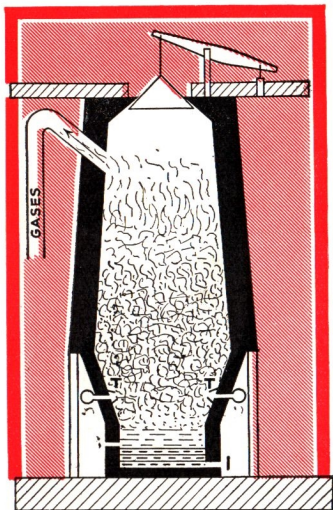


Fig. 7. Diagram of a blast furnace. The blast of air enters through the tuyeres T, and molten iron is tapped at the bottom I of the furnace.

thus produced is scarcely sufficient to melt the iron, but a doughy mass is obtained that can be worked by hammering.

The huge furnaces in which iron is now produced are known as blast furnaces, and Fig. 7 shows how they are built. Iron

ore, limestone and coke are fed in at the top of the furnace. In the fierce heat of the interior the limestone forms compounds that serve as fluxes, like the washing soda in our experiments, and the coke replaces the charcoal. As the mass descends through the furnace its temperature increases until at the bottom it is white hot, and the iron ore is reduced to the metal. The blast required to raise the temperature sufficiently high to melt the metal is delivered through heavy steel tubes called tuyeres, and from time to time molten iron is run out of the furnace into moulds, where it solidifies in the form of bars or "pigs."

Iron produced in the blast furnace is known as cast iron, and it contains about four per cent. of carbon. It is converted into steel, which contains less carbon, by blowing air through the molten metal in a huge pear-shaped vessel called a convertor, in which the unwanted carbon is burned out. Wrought iron contains less carbon than steel and is made by melting pig iron with scrap iron in a furnace lined with iron ore. The impurities of the cast iron are removed, leaving a pasty mass of wrought iron.

2. ALCHEMY AND CHEMICAL CHANGES

In addition to smelting operations for producing metals, the ancients carried out experiments with a variety of chemicals, including salt, potash, and soda, and means of making these were devised. Many important investigations were carried out by the chemists of ancient Egypt, a land in which metals were used at a very early period, and by those of Greece, Rome and other countries that later became prominent. The most famous chemists of the Middle Ages were the Arabs, and the name of the science is a reminder of this. It is derived from "alchemy," the name employed for the science during the Middle Ages, and this in turn was formed from *al kemia*, the name given to chemistry by the Arabs.

Changing Base Metals into Gold

Unfortunately many of the later alchemists spent their time in searching for the "Philosopher's Stone," an imaginary stone or mineral compound by means of which common metals could be changed into gold. In order to keep their supposed discoveries secret, they described their experiments in the most fantastic language. Their efforts to bring about this transformation failed, although many deceived themselves into the belief that they had



Fig. 8. An alchemist in his laboratory searching for means of transforming base metals into gold.

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succeeded, and innocent people, eager to share in the riches such a discovery would bring, were often swindled by imposters. Similar failure followed the attempts of the alchemists to find the "Elixir of Life," a wonderful liquid that was to confer eternal life on those who drank it! Although the alchemists failed in these researches, their experiments led to great advances in chemical industry. They also improved medical science, for they tested the value of almost every known chemical as a remedy for some form of disease.

Alchemy may be said to have come to an end in the 16th century when modern chemistry began with Robert Boyle (1627-1691), who is known as the "Father of Chemistry."

Boyle was a great experimental chemist, but is chiefly famous for his suggestion that all things were built up of a certain number of elements, or simple substances that cannot themselves be split up. The idea of elements was not new, for the ancient Greeks believed that all things were built up from fire, air, water, and earth, in varying proportions. At a later period it was thought that all metals contained a large proportion of mercury, while readily combustible substances were believed to consist chiefly of sulphur. To-day 91 chemical elements, or simple substances, are recognised.

Compounds Built Up from Elements

Substances that are not elements are known as compounds. They are built up from elements by combining these in various proportions, but they must be formed by what is known as a chemical change, in which is produced a substance that is entirely different from its components.

The two following experiments will make

clear the distinction between mere mixing and real chemical action. A Scoop (Part No. K36) is included in the Outfit for roughly measuring the quantities of chemicals required in experiments, and in this Manual one scoopful will be described as one measure. Add two measures of common salt to a Test Tube (Part No. K1) half full of water. Close the end of the test tube with the thumb and invert the tube rapidly several

times so as to shake the salt and water together. The salt disappears, but its presence in the liquid may be shown by tasting it. The salt has not combined with the water to form a new substance; it has only dissolved to form a solution, and the change is not a chemical one.



Robert Boyle (1627-1691).

The salt in this solution may be recovered by evaporating the solution, that is by heating it in order to drive off the water in the form of steam. For this

purpose the solution is poured in the evaporating dish and this is placed on the Gauze Square (Part No. K9) supported on the Evaporating Stand (Part No. K35), made by fixing the Ring (Part No. K29) at a convenient height on the Pillar (Part No. K28) fitted over the stud on the Base (Part No. K27). When all is ready the spirit lamp is placed under the gauze and the heating is continued until a white solid is left in the evaporating dish. On cooling this may be shown to be salt by tasting it. Evaporation is not a chemical change, for no new substance is produced.

Now mix two measures each of Sulphur (No. K131) and Iron Filings (No. K112) on the lid of a small tin. The particles of iron may readily be seen in this mixture, and may be extracted from it by means of a magnet. Thus no chemical change has taken place, for no new substance has been formed.

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Heat the mixture by holding the lid in a pair of tongs over a fire, or in the flame of the spirit lamp. Some of the Sulphur burns with a blue flame, but a black substance is formed on the lid. When this substance is cool, break it up and place part of it in a test tube. Add one measure of Sodium Bisulphate (No. K125) and cover with water. Bubbles are formed in the liquid, and a gas with the unpleasant smell of bad eggs is produced. This gas cannot be obtained by the action of Sodium Bisulphate on either Iron Filings or Sulphur. A new substance therefore has been formed by heating these two substances together, and the change has been chemical.

The burning of magnesium is another chemical change that may be brought about by heating. Cut a piece of Magnesium Ribbon (No. K116) an inch in length, and hold one end of it in a flame by means of the Test Tube Holder (Part No. K4), or a pair of pliers or small tongs. The metal burns with an intensely brilliant flame and a white ash that is a new substance is formed.

Water May Cause Chemical Change

The addition of water may bring about a chemical change. For instance, mix one measure of Tartaric Acid (No. K133) with one measure of baking soda in a dry test tube. There is no change until water is added, when a violent effervescence or bubbling immediately takes place owing to the production of a gas, which clearly is a new substance.

Other types of chemical changes take place in solution. Dissolve one measure of Cobalt Chloride (No. K105) in a test tube one-third full of water, and add to it a solution of washing soda made by dissolving a few crystals in a test tube containing water to a

depth of 1 in. A light blue solid is formed in the liquid.

A solid formed in this manner is called a precipitate, for it is precipitated, or thrown out of the liquid. In this case the solid is cobalt carbonate, an interchange taking place that gives also sodium nitrate, which is soluble in water and remains in solution.



Fig. 9. Displacing copper from Copper Sulphate by means of Zinc.

In other interesting examples of chemical change, one metal may be made to displace another. Dissolve two measures of Copper Sulphate (No. K108) in a test tube half full of water. Pour half of the solution into a second test tube, and dip in it the blade of your penknife, after cleaning the steel with emery paper. A reddish-brown coating is

formed on the steel. This is copper, displaced from the solution of Copper Sulphate.

To the remainder of the Copper Sulphate solution add three small pieces of Granulated Zinc (No. K134) and heat, holding the test tube in the test tube holder with its lower end in the flame, and its mouth pointing away from the face (Fig. 9). *This precaution should always be taken when heating a liquid in a test tube.* The Zinc acts on the hot Copper Sulphate solution and a reddish-brown powder is formed. This is copper, displaced by the zinc in the same manner as by the steel in the previous experiment. The solution becomes colourless and now contains zinc sulphate instead of Copper Sulphate.

Some of the early alchemists thought that they had brought about the transmutation (or change) of one metal into another when they made experiments of this kind, and their apparent success caused them to redouble their efforts to produce gold from less valuable metals.

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The breaking up of sugar into charcoal, or carbon, and steam when heated in a dry test tube is another example of chemical change, for in this case also new substances are produced.

3. BURNING AND BREATHING

In one of the experiments already made we showed that the burning of magnesium is a chemical change. The burning of any substance must be a change of this kind, for a new substance is always produced, and many striking experiments reveal the mysteries of burning, and also of breathing.

For the first experiment fix a short piece of candle on a cork, and float this on water about 2 in. in depth in a shallow bowl or basin, screwing Meccano bolts into the underside of the cork if necessary in order to make it float upright. Light the candle and cover it with an inverted glass jar placed with its mouth on the base of the bowl, and of course under the surface of the water (Fig. 10), in order that the candle may have a definite volume of air in which to burn.

The flame soon becomes paler and at last dies out altogether, and when the jar has become cold the level of the water in it is higher than at the beginning of the experiment. Thus some of the air has disappeared, although its escape was cut off. With the jar still upside down in the water extract the cork and candle. Then close the mouth of the jar under water with a sheet of paper or thin cardboard, and lift it out, placing it upright on the table. Test the remaining air by

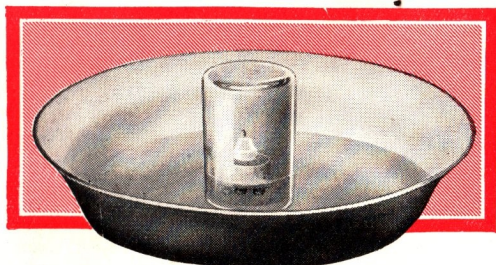


Fig. 10. Candle burning in a limited volume of air.



Fig. 11. Testing the gas left after candle has burned.

putting the lighted end of a taper in it, pulling the paper cover aside for this purpose (Fig. 11). The flame is immediately extinguished.

This experiment shows that air contains two gases, one of which is used up when the candle burns. The proportion removed in this manner is about one-fifth, and the remaining four-fifths is of no use for burning. The first of these gases is oxygen, formerly known as "fire air" because it is necessary in burning, and the second is nitrogen. Both are elements.

The experiment may be repeated with Sulphur (No. K131). This substance is best placed on a small tin lid resting on the cork, and its burning may be commenced by means of a lighted match.

Rusting is Slow Burning

Now let us try a similar experiment with a metal that does not burn, but apparently undergoes a change of another kind when exposed to the air. This is iron, which rusts readily, especially in moist air. Collect a number of small fragments of unruled iron, such as tacks and small nuts and bolts. Wrap

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these up in a piece of muslin, and suspend them from the top of the glass rod inside a jar standing upside-down in water (Fig. 12). Leave the jar undisturbed for a day or two, and then examine the contents of the muslin bag; they will be found to have acquired a coating of rust. As the iron slowly rusts, the level of the water in the jar rises, showing that something is taken out of the air. No matter how much iron is enclosed in the muslin, however, or how long it stays in the jar, the proportion of air removed is never greater than one-fifth, and the gas remaining in the jar will not allow a taper to burn in it. It is in fact nitrogen, and the iron in rusting has removed the oxygen.

This experiment shows that the rusting is a similar chemical change to burning, the only difference being that rusting takes place much more slowly than burning, and is not accompanied by a flame. Oxygen is necessary for both, and a candle will not burn, and iron will not rust, if this gas is absent.

These experiments give the solution to the problem of burning and rusting. What happens is that the substance concerned unites with the oxygen of the air, and for this reason the new compound formed is said to be an oxide. Thus iron rust is oxide of iron; the white ashes formed when magnesium and zinc burn are magnesium oxide and zinc oxide respectively.

Detecting Invisible Gases

It may be asked why a candle and sulphur do not form oxides of this kind when they burn in the air, for in burning they seem to disappear. The answer is that oxides are formed, but they are invisible because they are colourless gases. The gas formed when sulphur burns is sulphur dioxide, and may be

detected by its choking smell. Carbon dioxide is formed when a candle burns, for carbon is the chief element present in the material of the candle. This gas has no smell, but is easily recognised, although colourless and invisible, as explained in the following experiment.

For this purpose lime water is required, and this is made by putting one measure of Calcium Oxide (Lime) (No. K103) in a test tube containing water to a depth of about 2 in. Shake the test tube in order to dissolve as much lime as possible, and allow the remainder to settle to the bottom of the tube. The clear liquid is a solution of lime, and is usually known as lime water. Pour it into a second test tube and close this with a Small Cork (K18).

Place the cork with the piece of candle on it, used in a previous experiment, on the bottom of the jar and cover the jar loosely with a piece of card. When the candle is lit it burns for some time but eventually is extinguished. Then lift out the cork and candle and pour part of the lime water into the jar. On shaking, the lime water becomes milky in appearance. This is a test for carbon dioxide, and the effect of the gas in the jar on the lime water is a proof that the burning of the candle has resulted in the formation of this gas.

It is interesting to know that chemical changes similar to the burning of a candle take place within our bodies. We are built up of complex chemical substances containing carbon, and the oxygen of the air taken into our lungs when we breathe comes in contact with carbon compounds in the blood, which circulates through the lungs for this purpose. A gentle burning process occurs, with the formation of carbon dioxide; and the air we breathe out contains this gas in

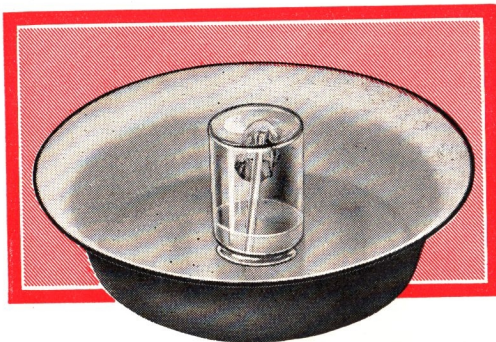


Fig. 12. Showing that the rusting of iron is slow burning.

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addition to nitrogen, and a proportion of oxygen that has not been used. In order to prove this, gently bubble the breath through a glass tube dipping below the surface of the lime water remaining in the test tube (Fig. 13). The liquid turns milky almost immediately, showing that carbon dioxide is present. The heat of the slow burning that produces this gas in the lungs helps to keep up the temperature of the body.

4. OXYGEN— THE ELEMENT THAT SUPPORTS LIFE

Oxygen clearly is a very important gas, and it is interesting to prepare it in the pure state. It is not easy to extract it direct from the air, but fortunately it can be obtained by heating certain chemicals that contain it. One of these chemicals is mercury oxide, and this was used by Dr. Joseph Priestley, the famous English chemist, when he discovered the gas in 1774.

Another chemical we may use as a source of oxygen is nitre or saltpetre, known to the chemist as Potassium Nitrate, and an interesting experiment will show the presence in it of this gas. Heat three measures of Potassium Nitrate (No. K124) in a small test tube until it melts and begins to bubble. Then remove the tube from the flame, and holding it almost vertically, and with the mouth pointing away from you, as in Fig. 14, drop into it half a measure of Charcoal (No. K104). This takes fire and burns as it floats on the liquid in the tube, and the flame is more brilliant than in the air because it is fed with pure concentrated oxygen

from the melted Potassium Nitrate.



Fig. 13. Breathing through lime water shows that breathing is a similar chemical change to burning.

When replacing the hot tube in the Test Tube Stand (Part No. K3), put a little Asbestos Fibre (Part No. K38) under it in order to prevent breakage by too rapid cooling. Although it is not absolutely necessary, this plan may be followed with all test tubes that have been heated.

A similar experiment may be carried out with half a measure of Sulphur (No. K131), which burns with a very brilliant flame.

Writing with Fire

Potassium Nitrate also may be used for fire writing. Add a few drops of water to one measure of the chemical in a test tube and warm until it dissolves, holding the test tube by means of the test tube holder and pointing its mouth away from you. Use the solution as an invisible ink to write or draw on thin paper, marking the beginning with a pencilled cross, and making all the strokes heavy and the writing or drawing continuous. Allow the paper to dry and touch the beginning of the writing with a red hot wire. A spark then travels slowly over the paper, tracing out the design or words, for the oxygen

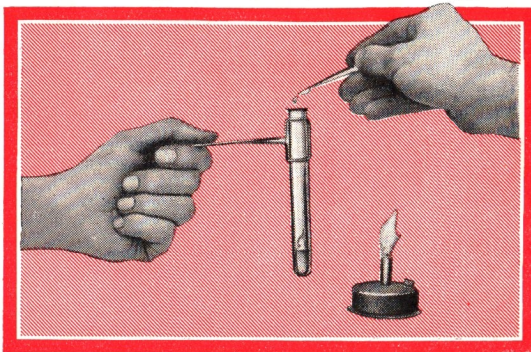


Fig. 14. Dropping Charcoal into melted Potassium Nitrate, in which it takes fire.

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of the Potassium Nitrate causes the paper on this track to be easily burned.

A more convenient chemical from which to

KI22) and Manganese Dioxide (No. KI18), the latter being a black substance that has the remarkable power of assisting the oxygen to escape from Potassium Chlorate at a lower

temperature than would be required in its absence. Crush the Potassium Chlorate that still remains and mix with it about one-quarter of the amount of Manganese Dioxide. Put the mixture in a test tube that has been thoroughly dried by keeping

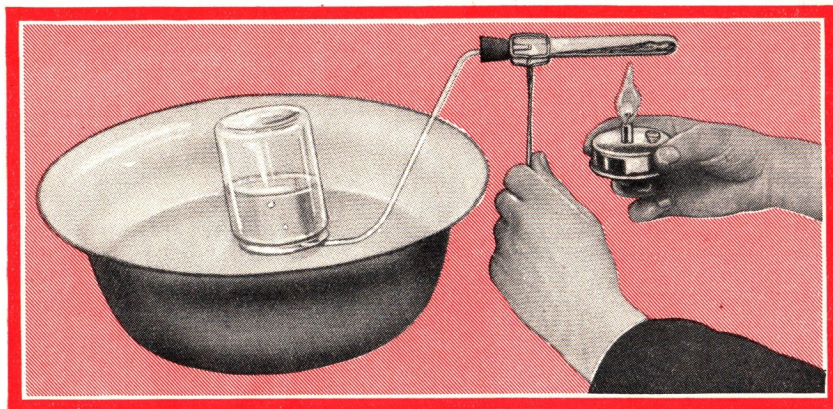


Fig. 15. Heating Potassium Chlorate and Manganese Dioxide and collecting the oxygen produced.

obtain oxygen is Potassium Chlorate (No. KI22). Crush two measures of this to powder by means of a piece of hard wood, and heat it in a small test tube. It melts, and on further heating, bubbles of oxygen rise to the surface.

In order to show that the gas produced is oxygen, put the glowing end of a half-burned chip of wood inside the tube. The wood will immediately burst into flame, for the glowing chip burns so furiously in the oxygen that the temperature rises to ignition point again. A glowing wood chip is an excellent means of testing for oxygen, and the wooden spills sold by tobacconists as pipe-lighters are specially suitable for the purpose.

Making Fire Air

Making oxygen on a large scale and collecting it is one of the most interesting and attractive chemical experiments. The gas is best prepared by heating a mixture of powdered Potassium Chlorate (No.

it for some time in a warm place.

Next, carefully work one end of the Double Angle Delivery Tube (Part No. KI4) through one of the small Bored Corks (Part No. KI9) with a screwing motion, a direct push being avoided as this may cause breakage. Fit the

cork thus equipped into the test tube containing the oxygen mixture.

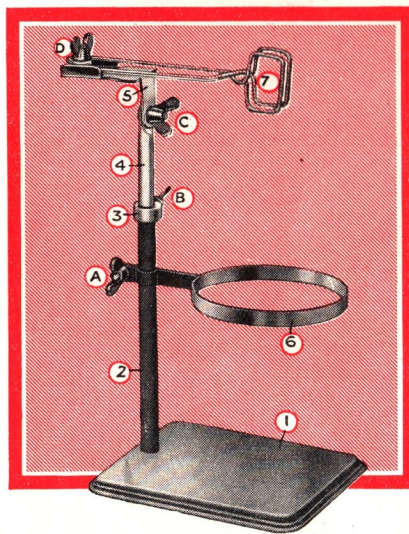


Fig. 16. The Universal Stand.

Means of collecting the gas must be provided. Test tubes may be used as gasholders, but small glass jars are more suitable, and the quantity of Potassium Chlorate gives

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sufficient oxygen to fill three vessels of the size of 1-lb. jam jars. A further requirement is a small bowl containing water to a depth of about 2 in.

The first jar to be used is filled with water and its mouth closed by pressing on it a piece of paper. It is then placed upside down on the bottom of the bowl, and when the paper is removed the jar remains full of water, which is kept in position by the pressure of the atmosphere. In order to collect the gas to be given off in our experiment, all that is necessary is to hold the test tube by means of the test tube holder so that the open end of the delivery tube is under the mouth of the inverted jar, when the gas produced on heating will bubble up and displace the water already occupying the jar (Fig. 15).

The Universal Stand (Part No. K26), included in the No. 3 Kemex Outfit also may be used to support the test tube holder. This is illustrated in Fig. 16, and is assembled by fitting the Pillar 2 (Part No. K28) over the stud on the Base (Part No. K27). Slip the Extension 4 (Part No. K30) into it, tightening the Wing Screw B (Part No. K42) in the Clamp 3 (Part No. K31) at the top of the pillar in order to fix the extension firmly in position. The end of the test tube holder 7 fits over a threaded rod on the Top Bracket 5 (Part No. K32), where it is held by means of a Washer

(Part No. K34) and Wing Nut D (Part No. K33). The Ring 6 (Part No. K29) and the Wing Nut A complete the Universal Stand, but are not required in this experiment.

Slacken the wing nut C immediately below the bracket, which is then turned round so that when the test tube containing the oxygen mixture is fitted in the holder it is horizontal. Then, by raising or lowering the pillar extension, bring the test tube to the required height, with the free end of the double angle delivery tube below the surface of the water in the basin (Fig. 17).

Tap the tube gently until the mixture forms a layer along the lower side,

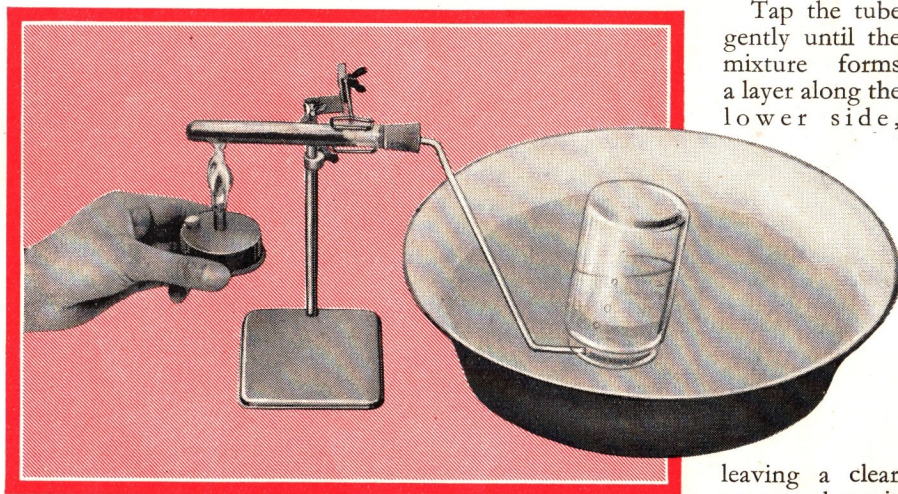


Fig. 17. The apparatus for the preparation of oxygen when the Universal Stand is employed to support the test tube containing the mixture to be heated.

leaving a clear passage above it for the gas; and commence to

heat the mixture by moving the lighted spirit lamp backward and forward under the test tube. Bubbles of air driven out by the expansion due to heating first escape, but soon the stream becomes more rapid owing to the production of oxygen, and then the inverted jar is placed above the end of the delivery tube. The spirit lamp may be put down for the brief time required for this, but a better plan is to call in the assistance of a friend, who may make himself useful by holding the jar in order to prevent it from toppling over, for it becomes unsteady when full of gas. Then continue to heat the tube steadily and evenly.

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Fig. 18. A candle burns furiously and with a very brilliant flame in oxygen.

Stop the heating when all the water in the jar has been displaced by gas, and immediately lift the delivery tube away in order to avoid the inrush of cold water that would follow the cooling of the gas inside it, for this would spoil the experiment and perhaps break the test tube. Then slip a sheet of paper over the mouth of the jar of oxygen and lift this out, keeping the paper tightly pressed down. To prevent the subsequent escape of the gas invert a saucer over the wet paper. The second jar may then be placed ready for use as a gasholder, and the experiment continued until all three jars are full of oxygen. Retain the mixture remaining in the test tube for use in a later experiment.

Furious Burning in Oxygen

Pour water into the first jar to a depth of about half an inch, and then lower into it a lighted piece of candle impaled on a wire (Fig. 18). The candle burns so furiously that the tallow melts rapidly in the great heat developed.

The flame becomes duller and smokier when the oxygen in the jar has been used up. Remove the candle, cover the mouth of the jar, and shake it in order to bring the water into contact with the gas remaining in it. Then place the vessel on one side for a later experiment.

Similar experiments with Sulphur (No.

K131) and Magnesium Ribbon (No. K116) are equally striking. In each case the bottom of the jar should be covered with water. One measure of Sulphur is sufficient, and may be burned in a tiny spoon made by suspending a piece of tin from a stout wire. Ignite the Sulphur by means of the flame of the spirit lamp and lower it into the second jar of oxygen, when its pale blue flame becomes far more brilliant. As in the case of the candle, shake the jar in order to cause the gas produced to dissolve in the water, and keep the solution for a further experiment.

Hold a piece of Magnesium Ribbon about $1\frac{1}{2}$ in. in length in the test tube holder and ignite it before lowering into the third jar of oxygen. The Magnesium burns with a very brilliant flame in air, but the flame becomes dazzling in intensity the instant it is surrounded by pure oxygen. A white ash is formed, and this may be allowed to drop into the water at the bottom of the jar.

5. CHEMICAL DETECTIVES

Heat a measure of powdered Litmus (No. K114) in a test tube half filled with water. The liquid rapidly becomes deep blue in colour as the Litmus dissolves. Shake the powder with the water and then allow the solution to cool. A little undissolved Litmus may remain in it, and it is therefore filtered in order to give a clear solution.



Joseph Priestley (1733-1804).

The Glass Funnel (Part No. K6) and a Filter Paper (Part No. K7) are required for filtering. Fold the circle of paper across its centre, and then again in order to form quadrants. The paper is

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then opened out into the shape of a cone (Fig. 19), with three thicknesses on one side and one thickness on the other. Fit the cone into the funnel, moisten the paper and press it gently into contact with the glass. Then place the funnel in a clean test tube, supported in the Test Tube Stand (Part No. K3), and pour the liquid into it (Fig. 20). Undissolved Litmus remains in the filter paper, and a clear blue solution is collected in the test tube under the funnel.

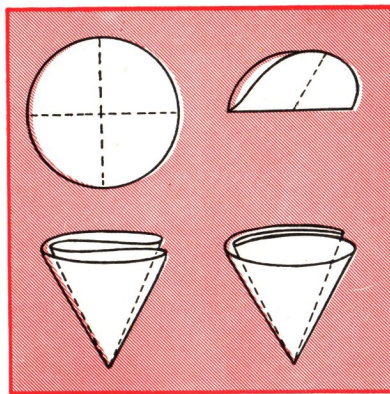


Fig. 19. How Filter Paper is folded in order to make it fit the glass funnel.

Pour a few drops of Litmus solution into the jars in which the candle and the sulphur were burned in oxygen. In the first jar the Litmus solution changes to a dull red, and in the second to pink. The oxides of sulphur and carbon are said to be acid oxides, and the change in colour of the Litmus solution from blue to pink or red is an indication of this.

It is not always convenient to use Litmus solution as an indicator and Litmus papers are sometimes employed instead. These are easily made from the solution already prepared by dipping into it strips of fine-grained blotting paper and allowing them to dry. A convenient size for the strips is 2 in. by $\frac{1}{4}$ in.

There are substances that have the power of bringing about the reverse change, and the red Litmus solution obtained by the addition of a few drops of acid to the blue solution we have prepared is once more turned blue when one of them is added to it. These substances are called alkalis. Red Litmus papers to be used as tests for alkalis may be prepared by dipping the blue ones in a solution formed by adding a few drops of vinegar, or dilute hydrochloric acid obtained from a chemist, to a test tube almost full of water.

Use Litmus papers to test solutions of common substances such as washing soda, baking soda and ammonia, and of Sodium Bisulphate (No. K125), Sodium Bisulphite (No. K126), Calcium Oxide (No. K103), and other chemicals included in the Outfit.

Other interesting indicators of a similar nature include the juices formed by crushing elderberries or dark coloured cherries, or by boiling shreds of fresh red cabbage with water. Test these also with substances already

known to be acids or alkalis.

In addition to Litmus, three other substances included in the Kemex Outfits may be used as indicators. One of these is Congo Red (No. K106) which undergoes opposite

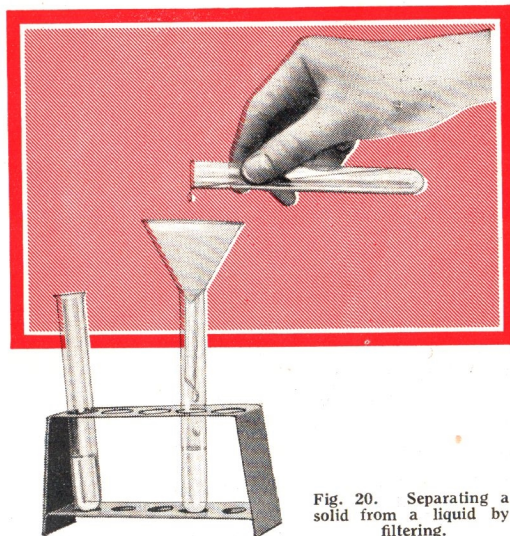


Fig. 20. Separating a solid from a liquid by filtering.

changes to those of Litmus. Make a solution by dissolving one measure of Congo

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Red in a test tube full of water and from this prepare test papers as in the case of Litmus. These turn blue when dipped into an acid



Fig. 21. Burning hydrogen at the mouth of the tube in which it is produced.

liquid, and back again to red when tested with alkalis.

The second indicator in the Outfit is Phenolphthalein (No. K121). This is a very complex chemical that is provided in the form of a solution in a mixture of alcohol and water. The solution is colourless except in the presence of alkaline substances, when it immediately becomes a brilliant pink.

The third indicator is Logwood (No. K115). Boil two measures of Logwood in a test tube half filled with water, and filter in order to obtain a clear solution. This is turned yellow by acids and blue by alkalis.

Mystifying Tricks with Indicators

A knowledge of the behaviour of indicators with acids and alkalis enables simple but mystifying tricks to be carried out. For instance, a strip of red paper may be dyed blue by dipping it into a red solution! The red solution is prepared by dissolving a few crystals of washing soda in a wineglassful of water and adding

a few drops of Phenolphthalein Solution (No. K121). The indicator immediately turns red, because of the presence of the alkali, and the colour of a red Litmus paper dipped into the solution is changed to blue for the same reason.

Changing water into "wine" is a very effective trick in which Phenolphthalein solution is used. In the neck of a wine glass place a drop of this indicator. The "water" for the trick is prepared in a separate glass vessel by dissolving washing soda or some other alkali in it, and on pouring it into the wine glass a pink colour is immediately produced.

This trick may be inverted and "wine" changed into water in an equally easy manner. The "wine" consists of water containing a few drops of Phenolphthalein solution, to which only just sufficient washing soda solution has been added to develop the pink colour. Stirring the "wine" with a glass rod changes it to water—provided that the rod has first been dipped into vinegar or some other acid solution! It is good fun to allow other people to try this trick, giving them a *clean* glass rod for use as a stirrer!

6. THE LIGHTEST GAS

Dissolve two measures of Sodium Bisulphate (No. K125) in a test tube containing water to a depth of one inch and drop into

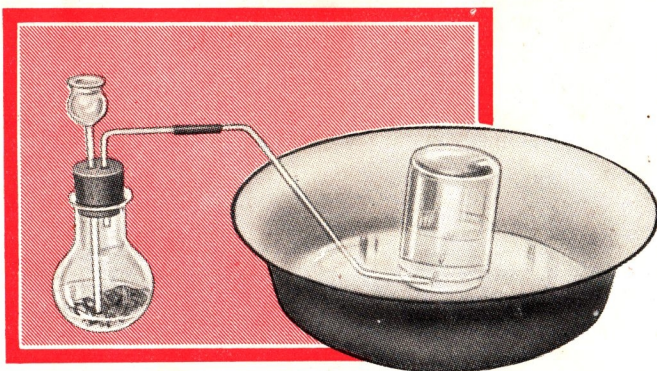


Fig. 22. Filling a jar with hydrogen given off when dilute sulphuric acid acts on Zinc. The gas takes the place of water in the inverted gasholder

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the liquid a strip of Magnesium Ribbon (No. K116) half an inch in length. There is a violent action, and a gas is given off that burns with a blue flame when the mouth of the test tube is brought to the flame of the Spirit Lamp (Fig. 21). Similar results follow the use of Granulated Zinc (No. K134) or Iron Filings (No. K112) instead of Magnesium, but in these cases the action is less violent.

The name of the gas that burns so readily is hydrogen. It is colourless and has no smell. It is lighter than air, and may be collected in a test tube held above the mouth of the tube in which it is being produced. When collected in this manner the gas usually is mixed with a proportion of air, and there is a loud but harmless explosion when a light is applied to it.

Making Hydrogen in the Gas Generator

Hydrogen is more readily obtained if dilute sulphuric acid of accumulator strength is employed. In order to collect the gas for further experiments place half the Zinc included in the Outfit in the Wide-necked Flask (Part No. K10) and insert the Large Cork (Part No. K17), into the holes of which the Thistle Funnel (Part No. K11) and the small Right Angle Delivery Tube (Part No. K12) have been carefully worked. The lower end of the tube of the thistle funnel must be near the bottom of the wide-necked flask. This constitutes the gas generator, and in order to collect the hydrogen we may follow the plan adopted in the case of oxygen, using test tubes as well as jars as gasholders (Fig. 22).

Pour sufficient water down the thistle

funnel to cover its lower end, and slowly add dilute sulphuric acid of accumulator strength until effervescence begins. Hydrogen then passes down the delivery tube and bubbles upward through the water in the large basin into the test tubes placed to receive it. As each test tube is filled, close it with the thumb and test its contents by bringing its mouth to the flame of the spirit lamp, which in this experiment must be kept as far away from the apparatus as possible. The gas first collected probably has a little air in it, and the mixture explodes with a loud noise, but harmlessly. When pure hydrogen is being collected the

gas will burn quietly, a blue flame travelling up the test tube.

Pouring an Invisible Gas Upward

Pour more acid down the thistle funnel if necessary, and fill two test tubes with hydrogen. Hold these side by side with their mouths open, but with

one inverted, and count 10 slowly. Then test each by means of a flame. No hydrogen is present in the test tube held with its mouth upward, but an explosion indicates that gas remains in the one that was inverted. This shows us that hydrogen is lighter than air.

A jar may be filled with hydrogen and the gas in it poured upward into a second jar as easily as the heavy gas carbon dioxide can be poured downward. Air is nearly 15 times as heavy as its own volume of hydrogen, and 1,000 cu. ft. of hydrogen give a lift in air of about 67 lb. Hydrogen therefore would seem to be ideal for use in airships, but is so readily inflammable that there is always danger of fire in envelopes filled with it.

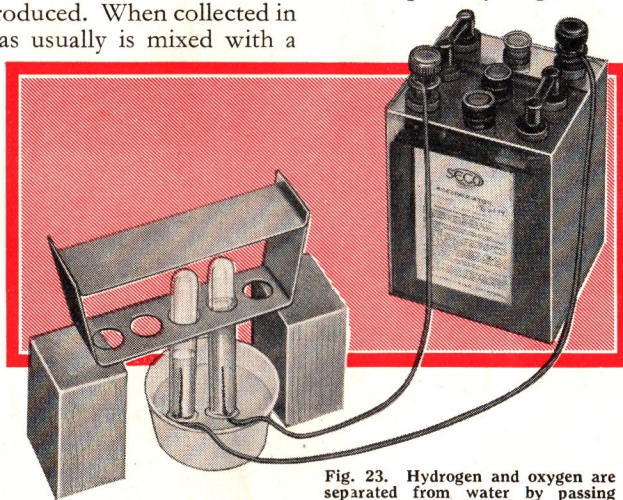


Fig. 23. Hydrogen and oxygen are separated from water by passing an electric current through it.

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Inflammable Gas from Water

To many who are not familiar with chemistry it is a surprising fact that hydrogen can be obtained by passing an electric current through water. This can readily be demonstrated by fitting up the apparatus shown in Fig. 23. The small basin under the test tube stand contains water, to which enough Sodium Bisulphate (No. K125) to cover a shilling has been added. Fill each of two test tubes in turn with water, close it with a finger, and place it upside down with its mouth under water in the basin. When both test tubes are ready, invert the test tube stand over them in order to keep them in position, blocks of

wood or other suitable objects being used for supporting the stand at the required height.

As a source of current the Meccano Accumulator or a flashlamp battery may be used, insulated copper wires being attached to the positive and negative terminals. A stout piece of copper wire is attached to the negative lead, and the length of Nickel Chrome Wire (Part No. K39) contained in the Outfit is similarly attached to the positive lead, these being carefully turned upward into the test tubes, as illustrated in Fig. 23. There they form "electrodes," the name given to the wires or plates that lead current into a solution.

As soon as the current is switched on, bubbles of gas are seen to rise in the two test tubes, and it is found that the quantity of gas formed at the negative pole is about twice that formed at the positive pole. The experiment may be allowed to continue until the test tube over the copper wire is nearly full, and the current is then switched off and the wires are removed.

The next step is to test the two gases formed. Close with the thumb the mouth of the tube containing the larger quantity and remove it from the basin. Then bring the mouth to the flame of the spirit lamp, remove the thumb, and the gas will take fire, burning with a blue flame. It is in fact hydrogen, and if the experiment is repeated it may be subjected to tests similar to those already described.

The second of the two gases is oxygen, as the introduction of a glowing chip of wood shows.

The hydrogen and the oxygen produced in this experiment have been formed from the water. The Sodium Bisulphate was added to enable a current to be passed more easily through the

liquid, for water is a poor conductor. Thus the liquid is shown to contain hydrogen and oxygen in the proportion of 2 to 1 by volume. Water therefore is hydrogen oxide.

7. WATER, NATURE'S SOLVENT

Water is one of the most important chemicals, as well as being one of the most abundant, for it is as necessary as oxygen to animal life, and plants cannot grow without it. In nature it occurs in the form of rain and in springs, rivers and oceans pure water is practically never met with, for the liquid is a wonderful solvent, being capable of dissolving a large number of substances. The purest natural water is rain water, but even that may contain dissolved carbon dioxide and other impurities often present in the air. As the water percolates through the soil and rocks it dissolves out other soluble components, and these are carried down to the sea, where they have been accumulating ever since the oceans were formed. Thus the sea is really a solution

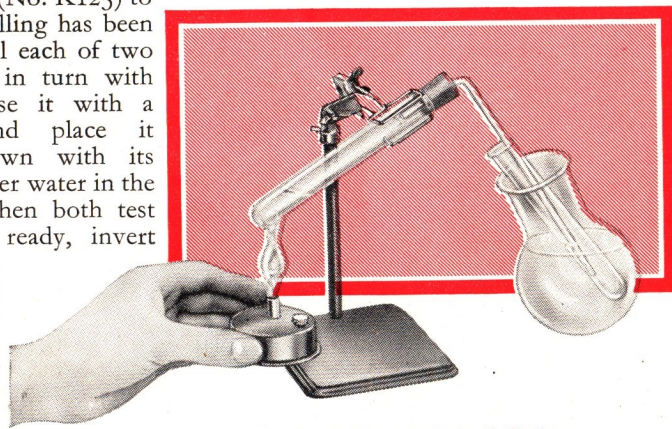


Fig. 24. Purifying water by distillation.

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in water of a large number of substances, the one present in the largest proportion being salt, which gives the sea its flavour.

How Pure Water is Made

Pure water may be obtained by distillation, the name given to the double process of evaporation or change of liquid to vapour by heating, and condensation or the opposite change from vapour to liquid, brought about by cooling. Place the small right angle delivery tube in a small bored cork and fit this into a test tube. Fix the test tube in the universal stand and attach a long straight piece of glass tubing to the free end of the delivery tube by means of the short length of Rubber Connection Tube (Part No. K37). This leads into a clean dry test tube, which stands in water in a basin or in the wide-necked flask (Fig. 24).

Water to a depth of about 1 in. is placed in the test tube held in the stand, and in order to show the effect of distillation a few crystals of common salt are added. Heat the salty water to boiling. Steam passes into the dry test tube, where it is cooled and condensed.

The distilled water collected in this manner does not taste of salt, and also lacks the pleasant flavour of tap water. Pure water may be prepared on a larger scale by means of the distillation apparatus shown in Fig. 25. One hole in the cork carries the small right angle delivery tube and in the other is a short piece of glass tube, one end of which has been heated in the flame of the spirit lamp until the glass has melted and sealed the tube.

Curious Properties of Water

Water is a poor conductor of heat, and a piece of ice with sufficient wire wrapped round to make it sink to the bottom of a test tube nearly full of water remains unmelted while the liquid above it is boiled (Fig. 26). Water may be heated quickly if advantage is taken of the fact that hot water is lighter than an equal bulk of cold water. If a flame is placed under a vessel containing water, the liquid

heated rises, forming what is called a convection current (Fig. 27). This explains why water in a kettle or pan must be placed above a fire instead of by its side, when the water in it is required to boil quickly.

Convection currents carry heat away rapidly from the bottom of a vessel containing liquid. This is shown in an interesting experiment in which water is boiled in a vessel made of paper. Fold a sheet of glazed paper as shown in Fig. 28. Place water in the vessel thus formed and suspend it by means of string above the flame. The liquid becomes hot and may even boil before the paper is burned, for the heat transmitted to the paper is communicated to the water and the temperature of this cannot rise above boiling point.

How the Chemist Uses Water

To the chemist water is of great importance because so many substances dissolve in it, and chemical changes readily take place in solution. In order to find whether a chemical is soluble in water, place one measure of it in a test tube containing water to a depth of an inch, close the end of the test tube with the thumb, slowly invert it,



Fig. 25. Another method of distilling water.

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and restore it to its normal position in order to mix the substance with the water. Test common salt, Strontium Nitrate (No. K130), Ferrous Ammonium Sulphate (No. K110) and Magnesium Sulphate (No. K117) in this manner. In each case the solid disappears from sight, and therefore is soluble in water.

Add a second measure of salt to the solution of salt already prepared and shake the tube. If the further quantity dissolves, add more. Eventually a stage is reached when the water in the tube will dissolve no more salt, and the solution is then said to be saturated. Different amounts of various soluble chemicals are required to saturate a given quantity of water. Repeat the experiment with Manganese Dioxide (No. K118). Shaking a measure of this substance with water makes no apparent difference to the quantity to be seen in the tube. In order to find if a small proportion has dissolved, filter and evaporate the filtrate to dryness in the evaporating dish. No residue is left, and therefore Manganese Dioxide is insoluble in water.

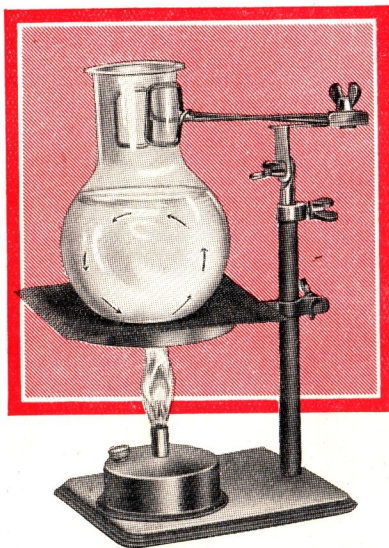


Fig. 27. The arrows show the direction of convection currents formed when the water in the flask is heated.

Test Calcium Oxide (No. K103) in a similar manner. It is not easy to see if any of the Calcium Oxide has dissolved, but a white residue is left after evaporation, and some therefore must have passed into solution. This chemical is said to be slightly soluble in water.

Separating Sand from Sugar

Mix four measures of sugar with an equal quantity of sand. Half fill a test tube with water, add the mixture and shake. Then filter the liquid. Undissolved sand is left on the filter paper. Wash it free from sugar by pouring successive small quantities of water through it, and put it in a warm place to dry. Evaporate the filtrate in order to recover the sugar.

This experiment shows how an insoluble substance may be separated from a soluble one. Repeat it with the residue in the test tube used in the preparation of oxygen. Unchanged Manganese Oxide is left on the filter paper. The solution gives a white powder on evaporation. The chemical name of this is potassium chloride.

Most chemicals are more soluble in hot water than in cold. Add one measure of common salt to the saturated solution already prepared and heat the liquid to boiling point. A little more salt dissolves.

Now try the same experiment with Potassium Nitrate (No. K124). The extra measure dissolves rapidly as the temperature is raised, and further measures of Potassium Nitrate added to the hot liquid also readily dissolve. A given volume of water dissolves nearly 20 times as much Potassium Nitrate at boiling point as at freezing point.

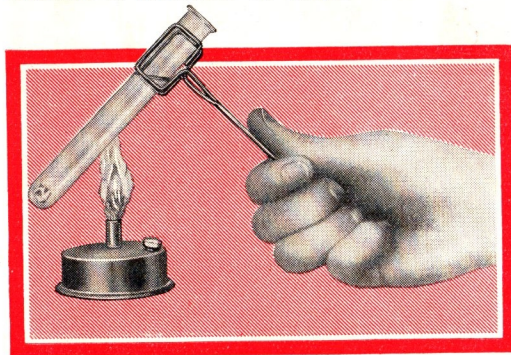


Fig. 26. Ice still unmelted in a vessel containing water that is heated to boiling point.

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Mysteries of Crystallisation

Allow the solution of Potassium Nitrate to cool and the extra quantity dissolved then separates out. Pour off the liquid and shake the solid on to a piece of blotting paper, and spread it out in a thin layer in order that the paper can absorb the remaining liquid. Examine the residue when it is dry. It consists of a number of needle-shaped pieces that have long flat sides and are of definite geometrical shape. These are crystals of Potassium Nitrate.

When a chemical separates from its solution in water in this manner it usually does so in the form of crystals. These vary in shape—each substance having its own peculiar crystalline formation—and also in size. Larger crystals are best obtained by allowing the solution from which they separate to cool slowly.

Prepare crystals of Copper Sulphate (No. K108), Tartaric Acid (No. K133), Iron Alum (No. K111), Magnesium Sulphate (No. K117), Lead Nitrate (No. K113), and other soluble chemicals contained in the Outfit. After drying them on clean blotting paper, compare their shapes, using a magnifying glass if one is available.

Small Crystals Grow into Large Ones

Make sufficient saturated solution of Copper Sulphate (No. K108) to fill an egg cup or a similar vessel, and by means of fine thread passing over a piece of wood laid across the top of the vessel hang in the solution a well-shaped crystal of Copper Sulphate obtained in one of the previous experiments. Place the vessel in a cool

place where it will not be disturbed, and examine the crystal from time to time. As the solution slowly evaporates the crystal grows, but always preserves its shape, the additional material being laid down in regular layers on the existing faces.

It is interesting to grow a string of crystals of sugar in this manner. Dissolve as much moist sugar as possible in a test tube nearly full of hot water, and place the tube in an upright position where it will not be disturbed. Tie one end of a piece of string, or strong thread, to a strip of wood, and place this strip across the top of the test tube with the thread, which must be nearly as long as the tube, hanging in the solution. A Meccano nut tied to the lower end of the thread will keep it in position. Allow the solution in the tube to cool slowly. Sugar begins to crystallise on the string and in time the small crystals grow into large ones, forming sugar candy.

Many crystals contain water. There is so much water in crystals of sodium carbonate, or washing soda, and sodium thiosulphate—which is the familiar “hypo” used in photography—that these substances

dissolve in the water set free when the crystals are heated. Place a few crystals of washing soda in a dry test tube and heat them gently. As the temperature rises the solid mass becomes a liquid and steam is given off.

Common salt, Potassium Nitrate and Potassium Chlorate contain no water of crystallisation, and crystals of them dried by pressing between sheets of clean blotting

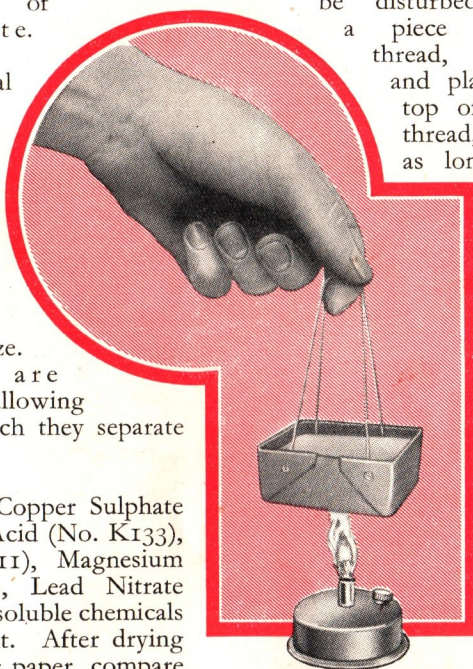


Fig. 28. Heating water in a vessel made of paper.

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Sir Humphry Davy (1778-1829).

paper give no steam when heated in a dry test tube.

Interesting colour changes sometimes take place when crystals lose water

of crystallisation. Heat a small crystal of Copper Sulphate in a dry test tube. Water is given off, and the colour of the Copper Sulphate changes from blue to white. Allow the tube to become cold and then add a drop of water. The white solid soaks up the liquid and then turns blue once more. This is sometimes used as a test for water.

Crystals of Cobalt Chloride (No. K105) show an even more interesting change when heated, the red colour changing to blue. Addition of water after cooling restores the red colour.

Ink for Secret Writing

Dissolve two measures of Cobalt Chloride in sufficient water to cover them at the bottom of a test tube. Using this solution as an ink, write on a sheet of glazed paper with a clean pen or a sharpened match stick. When the "ink" dries, the writing is practically invisible. Now hold the paper in front of a fire or over a flame, and the writing is revealed in strong blue lines. When the paper cools the colour fades away, and may be made to disappear more quickly by breathing on it. This solution may be employed as a secret ink.

The changes that have taken place on the paper are similar to those already described, the red Cobalt Chloride crystals losing water of crystallisation when heated

and becoming blue in colour. Water is then absorbed from the atmosphere, with the production of the red crystals.

The colour change of Cobalt Chloride may be put to excellent use in constructing a simple barometer. Steep a square of muslin of any convenient size in a solution of Cobalt Chloride made by adding one measure of the crystals to one-third of a test tube of water. Allow the muslin to dry, fix it to a piece of cardboard, and hang this in any suitable position. When the atmosphere contains much moisture and rain threatens, the muslin is pink in colour, but it becomes blue in dry sunny weather.

The piece of muslin containing the Cobalt Chloride may be shaped to fit into a pictorial design. For instance, it may be cut to represent the dress of a dancer, or that of a figure in a small picture, the muslin being fastened over the dress.

Sympathetic Inks and Magic Solutions

Liquids such as Cobalt Chloride solution are sometimes called "sympathetic" inks. There are many varieties, but not all are as



Fig. 29. Evaporating a solution to dryness on the Evaporating Stand (Part No. K35).

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good as Cobalt Chloride, for with this substance the writing may be made to appear and disappear as often as required. The juice of a lemon or of an onion, milk, and vinegar may be used as sympathetic inks in exactly the same manner as the Cobalt Chloride solution, but in these cases the heat of the fire reveals the writing by scorching the paper readily where it has been in contact with the sympathetic ink. The writing therefore cannot be made to disappear after it has once been brought out.

Writing with Water

Mix together one measure of Tannic Acid (No. K132) and one of Iron Alum (No. K111) that has been crushed to a powder. Place the mixture on a sheet of writing paper and rub it into the paper thoroughly with the aid of a dry wad of cotton wool or paper. Shake off the powder that has not been absorbed, and write or draw on the paper with a pen dipped in water. The result is a great surprise to those who have no knowledge of chemistry, for the water acts like a black ink, producing writing that is easily readable!

The explanation of this experiment is that in solution the two chemicals form a black precipitate of iron tannate. A very good ink may be made by dissolving one measure of each in separate Test Tubes half filled with water and mixing the solutions. A little gum should be stirred into the black liquid formed.

This chemical change enables Tannic Acid to be used as a secret ink. Make a solution of this in the manner already described, write with it on white paper, and allow the writing

to dry. Moisten a strip of clean blotting paper with a solution of Iron Alum of the strength already given and press this on the paper. The invisible writing immediately shows up as strongly as if freshly written with black ink.

This effective trick will mystify most observers if the writing with Tannic Acid is secretly carried out beforehand. It may be

reversed, the Iron Alum solution being used as the ink and the blotting paper dipped in a solution of Tannic Acid.

Similar experiments may be carried out with Iron Alum (No. K111) and Sodium

Ferrocyanide (No. K128), which give blue letters.

Analysing Water

Impurities that occur in natural water do not necessarily prevent it from being fit for drinking purposes, but it is important to ensure that water supplies of all kinds are free from harmful materials.

The first tests to apply to a sample of water are to examine its colour, taste and smell. Water that has a distinct smell may contain organic matter of animal or vegetable origin. It may therefore be unfit for drinking, for the organic matter may be accompanied by bacteria of a harmful kind. Safe water should have no smell, and also should be colourless and tasteless. Water that contains a trace of iron and is harmless may be slightly brown in colour, however, and good moorland water occasionally has a slightly peaty taste.

In order to test a sample of water for odour

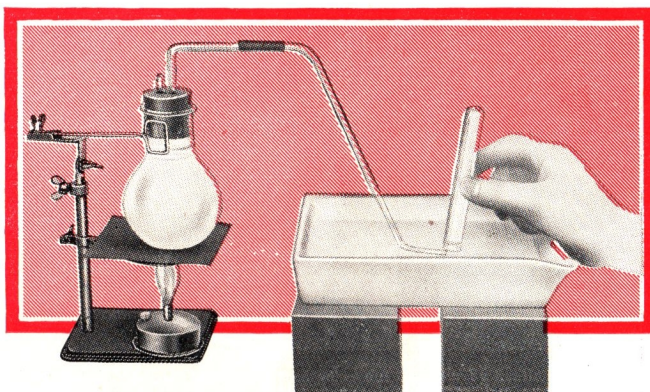


Fig. 30. Driving dissolved gases out of water by heating.

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fill a test tube with it, shake thoroughly, and warm slightly before smelling. The colour is best seen by placing a test tube filled with the water in front of a white surface. Occasionally tap water has a milky appearance, but this may be due simply to air bubbles, and on standing for a few minutes the water becomes clear.

The chief substances that are likely to be dissolved in water supplies are common salt and compounds of the metals calcium, magnesium and iron. In order to find if any quantity of solid matter is dissolved in the water under test, fill an evaporating dish with it and evaporate slowly to dryness (Fig. 29). Samples of water from different sources, including sea water if available, may be compared in this manner. River water

and that from wells and springs do not contain a large amount of dissolved solids, except in special instances, such as the medicinal waters of spas and health resorts.

How Muddy Water may be Purified

Occasionally water may contain insoluble solids in the form of fine powders that give it a muddy appearance. In order to see how it may be cleared, prepare a sample of muddy water by adding a little crushed clay to a test tube filled with water. Pour half this into another test tube and to this add half a measure of Aluminium Sulphate (No. K100) and half a measure of washing soda. Thus one test tube contains muddy water alone, and the other

similar water that has been specially treated. Shake the two for an equal length of time and then place them side by side in the test tube stand. Examine them from time to time without disturbing them, and it will be seen that the solid contents of the treated water are settling to the bottom of the test tube, while the untreated water continues muddy.

The clearing effect of the addition of the two chemicals is due to the precipitation of aluminium hydroxide. This is gelatinous, and

as it settles to the bottom of the test tube it carries the suspended impurities with it.

Testing Water for Metals

In order to test for salts of the metal calcium in a sample of water, dissolve two measures of crushed

sodium carbonate in a few drops of water and add the solution to a test tube half filled with the sample. The water slowly becomes cloudy or turbid if calcium is present. Calcium salts are the cause of hardness in water, and an artificially hard water that will answer to this test may be made by shaking a pinch of Calcium Oxide (No. K103) in a test tube filled with water. The explanation of hardness and its removal will be given later.

There are two simple tests for iron in water. For the first fill a test tube with the sample of water to be tested. Then dissolve half a measure of Iron Alum in a second test tube full of water, and place the two tubes side by side. To each add a few drops of a solution made by



Fig. 31. A charcoal burner at work.

dissolving one measure of Sodium Thiocyanate (No. K129) in water to a depth of one inch in a test tube. The solution containing iron becomes red in colour, and if iron is present in the water being tested, this will show a similar colouration, the depth of the colour depending on the proportion of iron present.

The second test for iron is made in exactly the same manner, substituting Sodium Ferrocyanide (No. K128) for Sodium Thiocyanate (No. K129). With this chemical solutions of iron compounds, and water containing them, give a blue colouration.

Gases More Soluble in Cold Water than in Hot

Certain gases are soluble in water. Carbon dioxide is an example, and mineral waters are charged with this gas under pressure, which has the effect of forcing more gas into solution.

Unlike Potassium Nitrate, salt and other solid chemicals already tested, gases that dissolve are less soluble in hot water than in cold. Pour a quantity of mineral water into a test tube and allow it to become "flat" as the result of the escape of the gas held in solution. Then warm it, and more gas is evolved, the hot liquid being unable to retain as much as could be dissolved in cold water. The gas produced may be proved to be carbon dioxide by pouring it downward into a tube containing lime water, or by fitting a bored cork and delivery tube to a bottle of mineral water in order to bubble the gas through lime water.

Dissolved Oxygen from Water

Water exposed to the air dissolves small proportions of oxygen and nitrogen, and the dissolved gases may be recovered by heating. For this purpose close one of the two holes in

the large bored cork by means of a piece of glass tube sealed at one end by heating in the flame. Pass the small right angle delivery tube through the second hole, but do not push its end beyond the cork. Connect the open end with the double angle delivery tube by

means of the length of rubber connection tube. Fill the wide-necked flask with water and press the cork into it, surplus water passing through the delivery tubes and completely filling them. Dry the wide-necked flask and place it on the ring of the Universal Stand with the free end of the double angle delivery tube under water in a basin (Fig. 30). Fill a test tube with water, invert it in the basin with its open end under water, and place it over the end of the delivery tube.

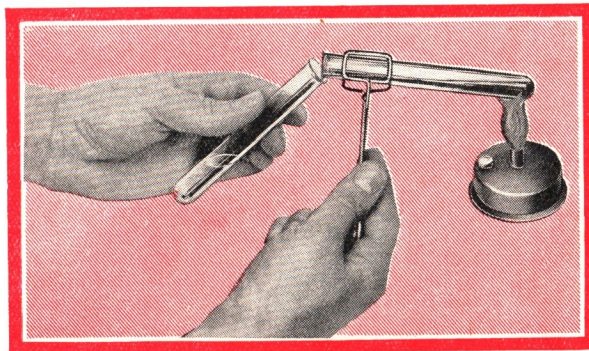


Fig. 32. An experiment to prove that sugar contains carbon.

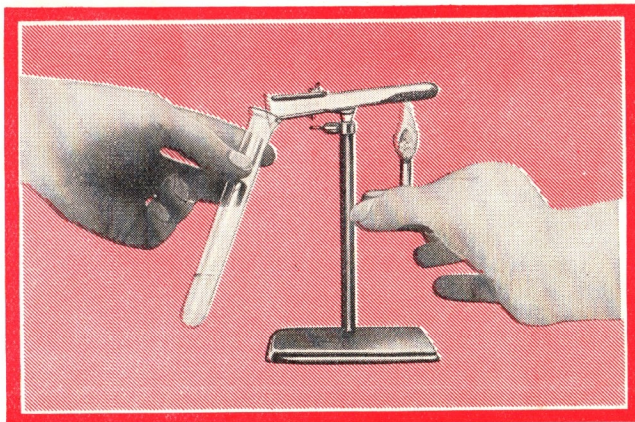


Fig. 33. Making carbon dioxide by heating an organic substance with copper oxide.

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When all is ready heat the water in the flask. Minute bubbles of gases dissolved in the water are formed and these rise to the top of the Wide-necked Flask and pass through the Delivery Tube to be collected in the inverted Test Tube.

Only a very small volume of gas can be obtained in this manner. When no more seems to be entering the collecting tube, close the end with the thumb and remove it from the basin, holding it right way up. Plunge a

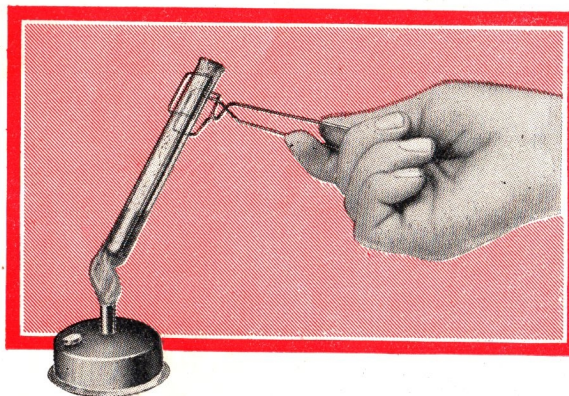


Fig. 34. Organic substances such as starch, sawdust, and sugar char when heated and give off an inflammable gas.

glowing wood splint in the small quantity of gas collected. This burns more brightly than in ordinary air, because the gas extracted from water in this manner is richer in oxygen.

8. AN ELEMENT WITH MANY DISGUISES

So far we have made experiments with three elements, oxygen, nitrogen and hydrogen. These are gases, but in combination with other elements they form solids and liquids. Now we turn to deal with a remarkable element that appears in several different forms. Its chemical name is carbon, but in its various disguises it is known to us as diamond and graphite, and also as charcoal, lampblack, gas carbon and coke, all of which are impure forms of the element.

It is surprising to learn that charcoal, obtained by slowly burning wood, is

chemically the same as diamond, which has long been prized for its brilliance and hardness. The difference between the two is that the diamond is crystallised carbon. Charcoal and certain other forms of the element that are not crystalline are described as "amorphous," a word derived from the Greek and meaning shapeless. For centuries charcoal has been used as a fuel, and charcoal burning, as its production is called, is one of the oldest of human occupations. The charcoal burner collects his logs into a pile and covers them with turf, leaving small holes at the top and near the ground in order to allow a steady current of air to pass through the heap. The wood at the bottom of the pile is lighted and slowly burned by regulating the quantity of air admitted, and the heat transforms the greater part of the mass into charcoal. The slow burning occupies several days (Fig. 31).

Finely powdered charcoal absorbs colouring matters from solutions. Dissolve a pinch of Congo Red in a test tube half filled with hot water and add four measures of Charcoal (No. K104). Shake the liquid for a few minutes and filter. The filtrate is colourless. Similar experiments may be carried out with Litmus and Logwood, and also with vinegar. The filtrate in the case of vinegar is acid to Litmus and may be used as a dilute solution of acetic acid, an organic substance that is the chief constituent of vinegar.

Brilliantly Coloured Dyes Contain Carbon

Many thousands of carbon compounds are known, and they are grouped together as organic substances, this name having been given to them because it was formerly supposed that they could only be produced by living organisms. They include the chemicals of which living things are built up, in addition to wonderfully coloured dyestuffs, perfumes, and such compounds as starch and sugar. It seems difficult to believe that some of these contain the element most familiar in the form of charcoal. A simple and conclusive proof of the presence of carbon in these and other organic materials is to obtain carbon dioxide

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from them by heating them with Copper Oxide, which yields the necessary oxygen.

Mix four measures of crushed sugar with one measure of Copper Oxide (No. K107), and heat this mixture in a small test tube held horizontally, placing a test tube containing a few drops of lime water with its mouth below that of the tube containing the mixture (Figs. 32 and 33), so that the heavy gas will flow downward into it. After heating strongly for a few minutes shake the lime water with the gas in the second test tube. The liquid turns milky, proving that the sugar contains carbon that has united with the oxygen of the Copper Oxide to form carbon dioxide. Similar tests prove the presence of the element in cheese, meat, bread, sawdust, paper and the other organic substances.

When the organic substances already mentioned are heated, they char and an inflammable gas and a tarry liquid are formed. Heat small quantities of each in a small test tube, and ignite the gas given off by bringing the mouth of the tube to the flame of the lamp. In all cases the gas burns with a yellow smoky flame (Fig. 34).

Coal is the most convenient source of inflammable gas, and coal gas is manufactured on an enormous scale for lighting and heating purposes. The coal is strongly heated in fireclay retorts, and the gas produced is passed through pipes leading first to pits in which tar and a dark liquid are deposited, and then through purifiers to giant gasholders.

A Miniature Gasworks

A miniature gasworks may be made from a test tube and a wide-necked flask, as shown in Fig. 35. Finely powdered soft coal is placed in the tube and strongly heated. The gas passes through the large right angle delivery tube into the flask, where a tarry liquid is deposited, and then through the upright glass tube to be collected in an inverted test tube, for it is lighter than air.

During the experiment the test tube acting as gasholder is raised slowly upward until its mouth is just clear of the upright glass tube. It is then closed with the thumb, and its contents are tested by means of a flame. The first tests made in this manner may give no result, for the gas produced has to displace the air in the glass tubes and in the flask; but later the presence of the gas is shown by a sharp explosion. The gas burns with a smoky luminous flame.

Coal gas is not now used directly for lighting purposes. Instead a proportion of air is admitted to the gas through special holes in the burners employed, and the blue flame thus obtained is used to heat a mantle composed of oxides of certain rare metals. The blue flame is not smoky, and is hotter than the flame of coal gas alone, and therefore is used in gas fires and cookers, where a clean flame is required.

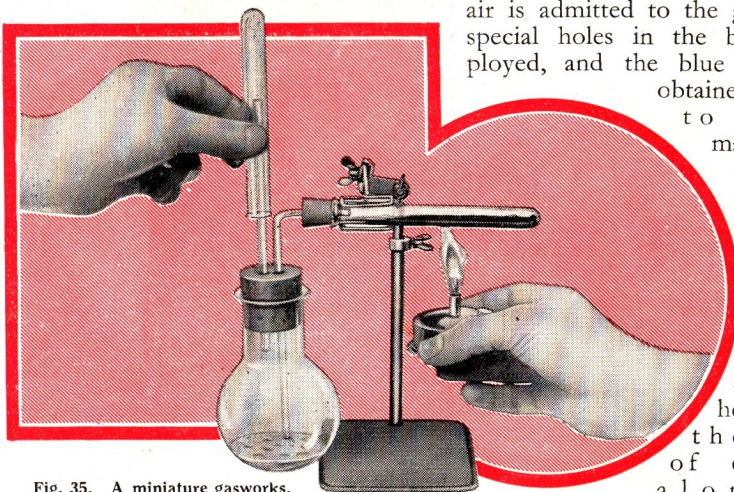


Fig. 35. A miniature gasworks.

A glass flame is hollow and contains unburnt gas. Similarly the interior of a candle flame contains unburnt vapour, which

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may be extracted by means of a glass tube, and lighted (Fig. 36). Another curious property of a flame is its inability to pass through the mesh of a Gauze Square until this becomes red hot (Fig. 37).

The liquid that collects in the wide-necked flask contains a tarry substance. At one time this was regarded as a nuisance, but it is now greatly valued as the source of benzene and other important organic substances from which dyes are made. Many thousands of dye-stuffs derived from coal tar are known, and explosives such as trinitrotoluene or T.N.T., drugs and anæsthetics, and the scents and essences used in many well-known perfumes are obtained from it.

Even this does not exhaust the uses of this amazing product. Mixed with the tar in the wide-necked flask is a brown liquid. Pour a little of this into a test tube, add a measure of Calcium Oxide (No. K103) and warm gently, when the smell of ammonia will be detected. The ammonia that is obtained by distilling coal in gasworks, and in special ovens erected at collieries, is used in the manufacture of artificial fertilisers. These supply the nitrogen that is required by many important food plants.

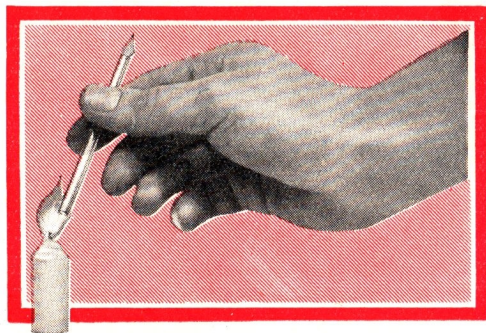


Fig. 36. Extracting unburned vapour from the interior of a candle flame by means of a length of glass tubing.

Experiments with Carbon Dioxide

One of the most important compounds of carbon is carbon dioxide, the heavy invisible gas we have already made by burning organic substances. It is more easily prepared by

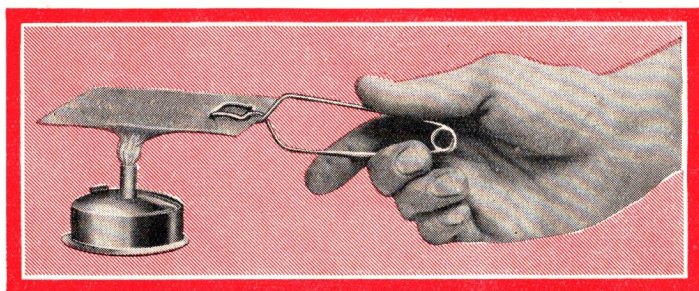


Fig. 37. The flame of the spirit lamp cannot pass through a gauze square.

pouring vinegar on crystals of washing soda placed in a test tube. A vigorous effervescence occurs, and the gas produced turns lime water milky when it is carefully poured downward into a second test tube containing this liquid. Washing soda is known to chemists as sodium carbonate, and is one of the salts of carbonic acid, the weak acid that is formed by dissolving carbon dioxide in water.

Many other common substances are carbonates. Test oyster shells, whiting, chalk, limestone, tooth powder, old mortar and marble by adding a little vinegar to a small quantity of each in turn, warming slightly if necessary (Fig. 38). In all cases carbon dioxide is given off. The chalk tested must be natural chalk, and not the so-called chalk employed for writing on blackboards.

Acid Drives Gas from Marble

Carbon dioxide for further experiments may be prepared by placing washing soda crystals in the wide-necked flask of the gas generating apparatus previously employed in preparing hydrogen, and pouring vinegar down the thistle funnel. A bottle fitted with the small bored cork and the small right angle

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delivery tube may be used instead of the wide-necked flask (Fig. 39).

A steadier stream of the gas is obtained from Calcium Carbonate (No. K102), included in the form of marble chippings, and dilute hydrochloric acid. Allow half the quantity of marble chippings in the Outfit to slide gently down the side of the flask and add sufficient water to cover the marble and the lower end of the funnel. Then pour down the thistle funnel sufficient dilute hydrochloric acid to give a brisk effervescence. The gas given off is led through the large right angle delivery tube into a small dry jar, or a dry test tube (Fig. 40).

The gas is heavier than air, and fills the lower part of the vessel first. At any moment the level to which it has risen may be found by carefully lowering a lighted splint of wood, which is extinguished when it reaches the gas. When the light goes out immediately inside the mouth of the jar, lift the delivery tube out and cover the mouth of the jar with a close-fitting card or sheet of paper. A second jar is then placed in position to collect the gas, and if necessary more acid is poured down the thistle funnel in order to keep up a brisk effervescence. Throughout the experiment it is advisable to add acid in small quantities as required.

Four jars of carbon dioxide may readily be collected in this manner. Pour half a test tube full of water into one jar and shake the gas with it, and add a few drops of Litmus solution after shaking. The Litmus is turned red, showing that the gas dissolves in water to form an acid solution. This solution contains carbonic acid.

Pour the gas in the second jar into another jar containing air. Carry out this experiment exactly as if pouring water, and then test each jar in turn with a lighted taper in order to follow the progress of the carbon dioxide. The invisible gas may be poured repeatedly backward and forward between two jars in this manner.

Pour the carbon dioxide in the third jar over the flame of the spirit lamp. The flame is extinguished immediately, and without the fuss that would follow the use of water for similar purposes (Fig. 41).

Carbon dioxide acts by blanketing out the air and thus choking the flame, for ordinary combustibles do not burn in the gas.

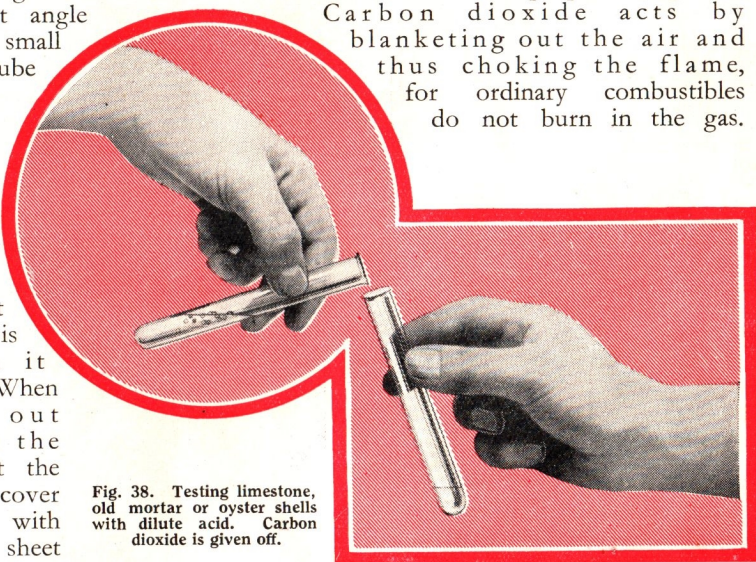


Fig. 38. Testing limestone, old mortar or oyster shells with dilute acid. Carbon dioxide is given off.

Burning Magnesium Robs Carbon of Oxygen

Ignite one end of a piece of Magnesium Ribbon (No. K116), about 2 in. in length, held in a pair of tongs or pliers, and lower the burning end quickly into the third jar of carbon dioxide. The metal continues to burn with a brilliant flame and its white ash is mixed with black specks of carbon.

Although a candle or a wood splint does not burn in carbon dioxide, burning magnesium is not extinguished because this metal has a

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greater attraction than carbon for oxygen. It robs the carbon of its oxygen therefore to form the white ash we know to be magnesium oxide, turning the black element out of the invisible gas.

We have already seen that carbon dioxide is soluble in water, and it is even more readily soluble in a solution of caustic soda, the chemical name of which is sodium hydroxide. In order to prepare this compound, place six measures of Calcium Oxide (No. K103) with 12 measures of

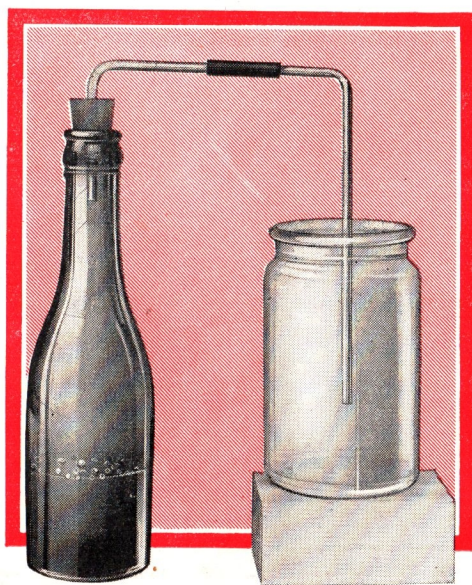


Fig. 39. Preparing carbon dioxide by the action of vinegar on washing soda.

washing soda in the evaporating dish, add half a test tube full of water and boil for a few minutes, adding a little water from time to time to replace that lost by evaporation. Allow the liquid to cool and filter it into a clean test tube. A test with Litmus paper shows the filtrate to be alkaline and it has a greasy feel when rubbed between the finger and thumb.

Pour the caustic soda solution obtained in this manner into the evaporating dish. Then fill a dry test tube with carbon dioxide,

and close it with the thumb. Invert the tube and open it with its mouth under the caustic soda. The liquid rises slowly and when a little has entered the tube, again close this with the thumb and shake. The gas is then absorbed and the lowering of the pressure in the tube is shown by the suction effect on the thumb, and still more strikingly by the rush of water into the tube when its mouth is opened under water.

Carbon dioxide is absorbed by caustic soda with the formation of sodium carbonate. Pass the gas for five minutes through the remainder of the caustic soda solution in a test tube. Evaporate the solution nearly to dryness, and crystals of sodium carbonate are formed. Dry these crystals and test with vinegar or dilute hydrochloric acid in order to show that carbon dioxide is liberated.

How "Washing Soda" differs from "Baking Soda"

Sodium carbonate is capable of absorbing even more carbon dioxide to form a new substance, known as sodium bicarbonate, because it contains a higher proportion of carbon dioxide than does washing soda. Make a concentrated solution of washing soda by placing sufficient of the crystals in a test tube to occupy about a quarter of its length and covering them with water, warming to cause the crystals to dissolve. Pass the gas through this solution and a white solid is precipitated. This is sodium bicarbonate, usually known as "baking soda" to distinguish it from "washing soda."

Filter and dry the new substance by pressing with clean blotting paper. Heat part of it in a dry test tube, and it loses carbon dioxide, again forming ordinary sodium carbonate. Dissolve the rest in a small quantity of water in a test tube and heat to boiling point. Again the sodium bicarbonate is decomposed and carbon dioxide is driven out of the solution.

Changes of an equally interesting kind occur when carbon dioxide from the

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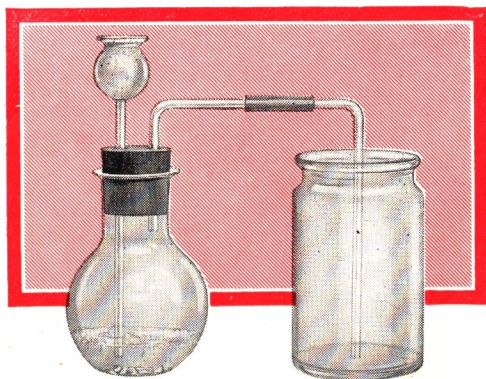


Fig. 40. A steady flow of carbon dioxide is obtained by the action of dilute acid on marble in the gas generating apparatus.

generating apparatus is passed through lime water contained in a test tube. The solution immediately turns milky, but eventually again becomes clear. Boil the clear liquid and once more it becomes milky, bubbles of carbon dioxide being given off.

The first action of the carbon dioxide on the lime is to form calcium carbonate, which is a white substance insoluble in water and is therefore precipitated. On continuing to pass carbon dioxide, calcium bicarbonate is formed, and as this is soluble in water the solution becomes clear once more. Finally, boiling the solution of calcium bicarbonate causes this substance to decompose, calcium carbonate being reformed with the loss of carbon dioxide, and thus the solution again becomes milky.

A similar series of changes are brought about by breathing through a tube dipping into lime water, but this may take a little longer owing to the small proportion of carbon dioxide in the breath. Add two drops of Phenolphthalein Solution (No. K121) to the lime water. The indicator turns pink, and the colour is removed when calcium carbonate is formed.

From Shell Fish to Marble

The changes that carbon dioxide brings about in lime water take place on an

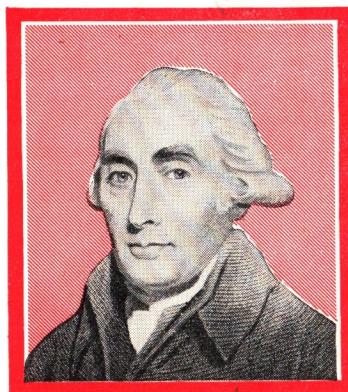
enormous scale in nature, for calcium carbonate is found in immense quantities in the Earth, where it is known as chalk, limestone or marble, according to the form in which it occurs. Chalk consists of the shells of dead marine animals. In past ages these shells have accumulated at the bottom of the oceans and geological changes have caused the deposits to be lifted above sea level. Limestone and marble are more compact forms of calcium carbonate, produced from chalk by the combined action of heat and pressure. Marble is actually crystalline, and often beautifully coloured owing to the presence of compounds of iron and other metals.

The carbon dioxide that takes part in these chemical changes comes from the atmosphere, and a proof of its presence there may be obtained by leaving lime water in the evaporating dish for a few hours, when its surface becomes covered with a crust of calcium carbonate. It is interesting to note that carbon dioxide in the atmosphere causes the lime in mortar to change slowly into calcium carbonate and thus helps in the "setting" or hardening process.

Gigantic Caverns Carved out of Limestone

Carbon dioxide from the air is carried to limestone or chalk in the Earth by means of rain water.

As this percolates through the chalk or limestone, it slowly dissolves these materials to form calcium bicarbonate, exactly as in our experiments,



Joseph Black (1728-1799), who first recognised carbon dioxide.

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and gigantic underground caverns have been eaten out of the Earth in districts where these forms of calcium carbonate are found. Stalactites and stalagmites, columns of limestone descending from the roof and rising from the floor, are well-known features of the limestone caves produced in this manner. These columns sometimes unite to form gigantic pillars that continue to grow in thickness. These often take fantastic forms and are coloured if compounds of iron and other metals are present.

The explanation of these curious formations is the reformation of calcium carbonate when water containing calcium bicarbonate in solution comes into contact with the air. A proportion of carbon dioxide is then lost and insoluble calcium carbonate is deposited, the chemical change being the same as that taking place when calcium bicarbonate solution is boiled. As the water drips from the roofs of caves, the small amount of calcium carbonate set free forms tiny projections, and these are slowly lengthened and made thicker by the continual dripping. More carbon dioxide is lost as the drops fall to the floors of the caverns and minute quantities of calcium carbonate are deposited there, thus beginning the erection of stalagmites. The action is slow, but many of the world's great limestone caves have been in existence for thousands of years, and wonderful stalactites and stalagmites have been built up in them by means of this remarkable chemical change.

Carbon Dioxide in Bread Making

Carbon dioxide is used for giving mineral waters their sparkle, being forced under pressure into these liquids. It is also used in making bread. When required for this

purpose it is obtained by the addition of living yeast to the flour and water employed in making dough. Yeast is a vegetable organism consisting of microscopic cells that grow and multiply by the simple process of the division of one cell into two when placed in a warm solution of sugar, on which they feed, with the production of carbon dioxide. The bubbles of gas formed in this manner grow larger when the dough is left in a warm place to rise, and they blow the dough up into a light and spongy mass.

In order to show the production of carbon dioxide by this change, dissolve a teaspoonful of sugar in water to a depth of an inch in the wide-necked flask and add a few fragments of dried yeast, which is pale yellow in colour and consists of cells

that are in a dormant condition. Then introduce the double bored large cork with a short piece of glass tube sealed at the end in one hole, and the small right angle delivery tube in the other. Connect this tube to the large right angle delivery tube dipping into a test tube containing lime water (Fig. 42).

Leave the apparatus in a warm place for a few hours. The yeast will then have grown to such an extent that the mass of active cells nearly fills the liquid, which has become frothy because of the evolution of gas. This gas is proved to be carbon dioxide by its action on the lime water, which in the meantime has become milky.

The yeast splits the sugar into carbon dioxide, and alcohol which remains in the liquid. Bread does not contain alcohol, however, for this liquid is evaporated during baking. The whole process is described as fermentation, and the yeast is known as a ferment.

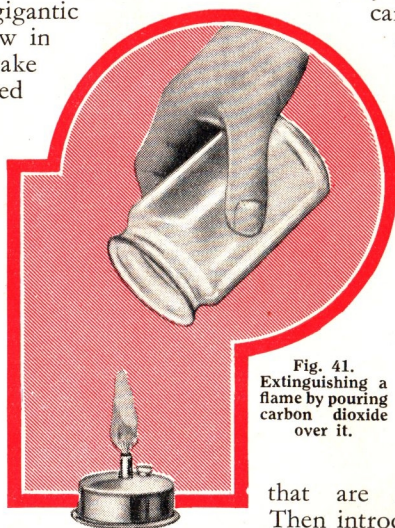


Fig. 41.
Extinguishing a
flame by pouring
carbon dioxide
over it.

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Plants absorb carbon dioxide and give out oxygen, thus behaving in exactly the opposite manner to animals. Pass carbon dioxide through water in a jar, place water plants in this and invert the jar in a basin containing water (Fig. 43). Bubbles of oxygen form on the leaves of the plants, and these collect at the top of the jar. This interesting process can be watched at any time in a home aquarium that has been well stocked with fresh water plants.

Why Hard Water Wastes Soap

Washing with water from wells and reservoirs in chalk and limestone districts is sometimes difficult, for the effect of rubbing a little soap into the water is to form a hard scum, and an excessive quantity is required to form a satisfactory lather. Rain water behaves quite differently, forming a copious lather with a very small proportion of soap. Rain water is therefore said to be "soft," while water that does not readily lather is described as "hard."

Hardness due to calcium bicarbonate is known as temporary hardness, for it may be removed by simply boiling. As our experiments have shown, this has the effect of precipitating calcium carbonate, and thus getting rid of the cause of the hardness. The "fur" formed in kettles that have been in use for a considerable time is calcium carbonate produced in this manner from hard water. A piece of this material chipped out from an old kettle readily gives an effervescence when dilute acid is poured on it, carbon dioxide being displaced from it.

Hardness is sometimes due to the presence of small proportions of calcium sulphate and of compounds of magnesium. Hardness of this kind is said to be permanent, because it cannot be removed by boiling; and chemicals, such as sodium carbonate, that precipitate insoluble compounds of calcium and magnesium, must be added in order to overcome the difficulty.

Most of the carbonates of the metals are insoluble in water, and are prepared by precipitation. Nickel carbonate is an example. Make a solution of washing soda by adding half a teaspoonful of small crystals to a test tube full of water. Then dissolve one measure of Nickel Ammonium Sulphate (No. K120) in a test tube half filled with water and add washing soda solution drop by drop (Fig. 44). A light green precipitate of nickel carbonate is obtained. Filter and wash the precipitate by pouring successive small quantities of warm water through it. Finally allow it to dry, and test a portion in a dry test tube by adding dilute acid to it. An effervescence takes place, carbon dioxide being produced.

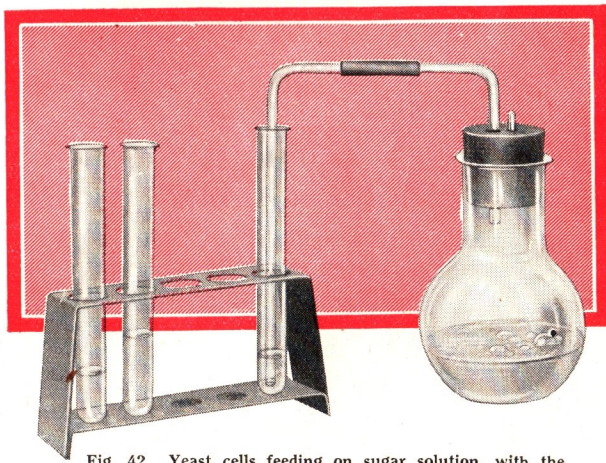


Fig. 42. Yeast cells feeding on sugar solution, with the formation of carbon dioxide.

Other insoluble carbonates that may be similarly prepared and tested are copper carbonate, green; magnesium carbonate, white; manganese carbonate, pale pink; cobalt carbonate, blue; lead carbonate, white. Wash and dry these

precipitates in the manner already described and in each case test a portion by adding a few drops of dilute acid. Compare the results with those obtained in the experiment with nickel carbonate.

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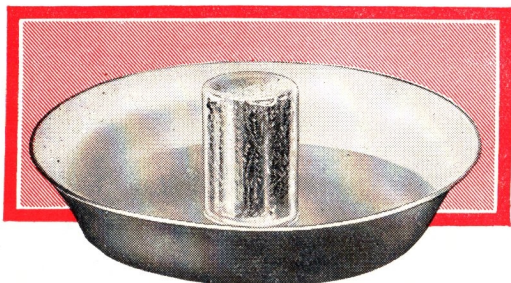


Fig. 43. Water plants in water saturated with carbon dioxide. Bubbles of oxygen form on the leaves and collect in the jar.

Heat the remaining portion of each carbonate prepared in this manner in a small dry test tube, holding a glass rod on the end of which is a drop of lime water in the mouth of the test tube while heating is taking place. In each case the lime water is turned milky, showing that carbon dioxide is given off. The residue is the oxide of the metal. Interesting colour changes often take place. Thus green copper carbonate becomes black copper oxide, and white lead carbonate becomes yellow lead oxide.

Most carbonates are readily decomposed by heating, sodium carbonate being an exception in this respect. Iron carbonate is decomposed so readily that iron oxide is formed immediately sodium carbonate solution is added to one containing iron alum. The precipitate formed in this case is brown and is really iron hydroxide, a compound of iron oxide and water. Precipitated copper carbonate also changes easily, becoming black copper oxide when the liquid in which it is formed is boiled.

Quicklime from Marble

Lime is made by heating marble, limestone or chalk, and is calcium oxide. Place a tin lid containing a few small pieces of marble on a bright fire. The marble soon loses its crystalline appearance and becomes dull white. Lift the tin lid from the fire by means of a pair of tongs and allow it to stand in a dry place until cool. Then very cautiously allow one or two drops of water to fall on the residue, holding the lid as far away from the face as possible. The white mass crumbles and steam is given off.

The residue that is formed when marble is heated is known as quicklime. It readily combines with water to form calcium hydroxide, or slaked lime, and as we have seen, this action takes place with the production of heat. Lime left exposed to the air takes up moisture slowly to form slaked lime.

Shake the lime prepared in this experiment in a test tube half filled with water. Carefully pour off the clear liquid left after settling has taken place, filtering if necessary to remove all undissolved lime. Then pass carbon dioxide into the filtrate, or bubble your breath through it, in order to show that it behaves in exactly the same manner as the lime

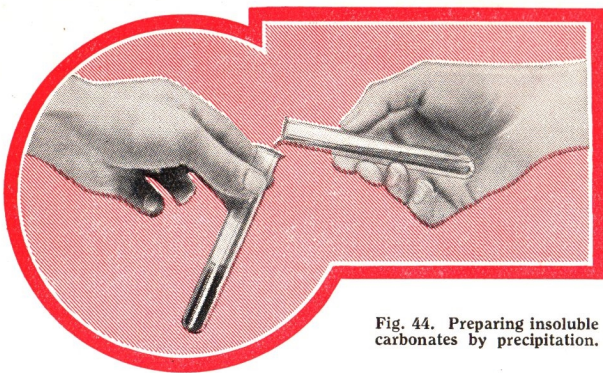


Fig. 44. Preparing insoluble carbonates by precipitation.

used in previous experiments. The liquid becomes milky because of the formation of a white precipitate of calcium carbonate. It is interesting to note that by this simple means the chemical

from which the lime was made has been reformed. It differs in appearance from the marble, because it is not crystalline.

9. NITROGEN—A SLUGGISH ELEMENT

Our experiments have shown that about four-fifths of the atmosphere consists of nitrogen, a colourless gas that prevents burning, breathing and other changes in which the air takes part, from proceeding too quickly. Nitrogen does not seem eager to unite with other elements, but although sluggish and lazy, it forms part of many complex chemicals that are present in plants and in animals. Life indeed would be practically impossible without it, and it must be present in a certain proportion in the food of both plants and animals.

In order to show the presence of nitrogen in a substance of animal origin, add three measures of Calcium Oxide (No. K103) and a few drops of water to a small quantity of wool in a test tube. Place a piece of moist red Litmus paper over the mouth of the tube and heat the mixture gently. The Litmus paper becomes blue, and on cautiously smelling the vapours given off the odour of ammonia is recognised. Ammonia is a compound of nitrogen, and its production in this experiment is a proof of the presence of the element in wool. Make similar tests with hair and cheese.

Ammonia is a gas and is more easily prepared by heating one measure of Ammonium Chloride (No. K101) with one measure of Calcium Oxide (No. K103). The smell of ammonia is detected immediately on mixing the two chemicals, and gentle warming produces the gas in greater quantity.

Dip the end of a clean glass rod into dilute hydrochloric acid, and hold it near the open end of the test tube in which ammonia is being produced. Thick white clouds are immediately formed by the combination of the ammonia with the hydrochloric acid to form ammonium chloride (Fig. 45). This compound is a white solid and is produced in the form of tiny particles.



Fig. 45. Ammonia formed by heating Ammonium Chloride with Calcium Oxide. The gas forms white fumes with hydrochloric acid.

Ammonium Chloride Smoke Rings

An interesting experiment with clouds of ammonium chloride may be carried out with the aid of the box illustrated in Fig. 46. This is made by cutting a small hole about $\frac{1}{2}$ in. in diameter in the centre of the base of a small cardboard box, and a rectangular opening in the lid, this opening being nearly as large as the box itself. Over the rectangular opening gum a sheet of parchment paper to replace the cardboard cut out.

Stand the box on end, with shallow saucers or dishes containing warm household ammonia and hydrochloric acid respectively inside it.

Thick clouds of ammonium chloride are formed, and on smartly tapping the parchment cover rings of smoke are shot out of the circular hole in the base. These are known as vortex rings, and it is interesting to see the smoke composing them whirling round as the rings drift away from the box. Smaller rings may be obtained by pasting paper or thin card over the hole and cutting out a new opening of smaller diameter. This is a particularly attractive experiment and one that never fails to interest and amuse all who see it performed.

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An Ammonia Fountain

Ammonia gas is lighter than air and may be collected in dry test tubes by means of a delivery tube turned upward. It is very soluble in water, as the following attractive experiment shows.

Fit the mouth of a perfectly dry test tube with a small cork carrying the small right angle delivery tube, and to the end of this attach a length of glass tubing by means of the short piece of rubber tubing. The experiment is most effective when the outer end of this length of glass tubing is drawn out into a jet.

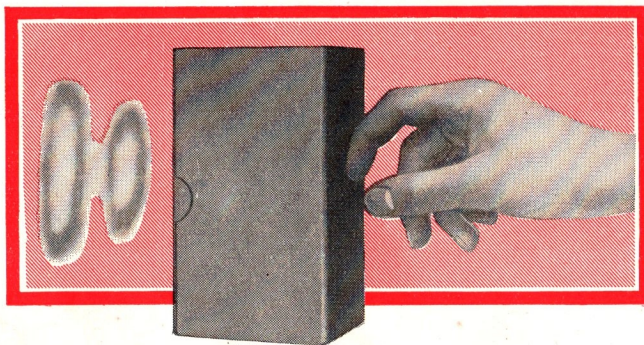


Fig. 46. Vortex rings of the fumes formed by ammonia and hydrochloric acid.

For this purpose hold the tubing horizontally, with the flame of the spirit lamp playing on a point near its end, and rotate it until the heated section becomes soft. Then pull gently, continuing to rotate the tube. This has the effect of narrowing the bore at the point at which the glass tube has been heated. After cooling make a scratch with the corner of a file in the narrow portion, and the unwanted piece at the end may then be broken off easily.

Close one hole in the double bored large cork by means of a piece of glass tube one inch in length sealed at one end by heating in the flame. Through the second hole work the long glass tube, pushing the jet so far that it projects halfway into the wide-necked flask when the cork is placed in position.

In the test tube place a mixture of half the Ammonium Chloride (No. K101) contained in the Outfit and an equal quantity of Calcium Oxide (No. K103), and fit the apparatus together as shown in Fig. 47. Do not

allow the cork to close the wide-necked flask, which must be perfectly dry; but leave a small opening for the displacement of the air in the flask by the ammonia that is formed when the mixture is heated. In order to complete the preparations have at hand a jar containing at least twice the quantity of water required to fill the flask, and colour this by pouring in a few drops of Litmus solution that has been made red by the addition of one drop of very dilute hydrochloric or sulphuric acid.

When all is ready heat the mixture in the test tube. Ammonia passes through the delivery tube into the wide-necked flask, and the heating is continued until the smell of ammonia is readily detected near the apparatus, showing that sufficient ammonia to fill the flask has entered it.

It is advisable to carry the heating a little further in order to ensure that the air originally in the apparatus is displaced. Then close the flask by means of the cork and immediately detach the jet from the rubber connection tube, placing the finger on the end of the glass tube in order to prevent the escape of ammonia. Put the end of the jet under the coloured water in the jar, holding the glass tube upright with the inverted flask at the top, and remove the finger. Water rises slowly up the tube, dissolving the ammonia as it does so, and as soon as a few drops have oozed out of the tube, the gas is dissolved so rapidly that the coloured liquid rushes up like a fountain and fills the flask (Fig. 48). As it does so, its colour is changed from red to blue. When Congo Red is used as the indicator instead of Litmus solution, the opposite colour change takes place.

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If the fountain does not commence to play immediately, it may be started by closing the end of the glass tube with the finger again, lifting the apparatus out of the water and turning it the other way up. The column of water in the tube runs slowly down towards the flask when the pressure of the finger is cautiously released. Allow part of this to enter the flask containing the ammonia, and then tighten the pressure and replace the end of the glass tube in the coloured Litmus solution in the jar.

On removing the finger the solution rushes up the glass tube and changes colour as it enters the wide-necked flask.

Making Ammonia Solution

The mixture of Ammonium Chloride and Calcium Oxide used in preparing the ammonia

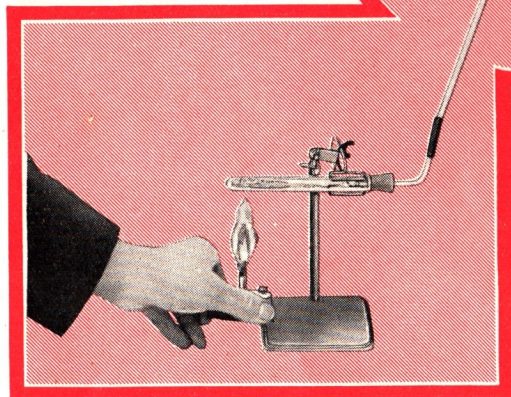


Fig. 47. Filling the wide-necked flask with ammonia when preparing an ammonia fountain.

fountain will not have been exhausted, and may be used for making ammonia solution for further experiments. The large delivery tube may be used for passing the gas through water (Fig. 49). The delivery tube must dip only just below the surface of the water, however, and must be lifted out as soon as the stream of bubbles of gas ceases, in order to prevent water from

entering the hot tube, and breaking it. A better plan is to use the thistle funnel with the small right angle delivery tube. Turn the right angle delivery tube to point downward and attach the tube of the thistle funnel to it, allowing the broad end of the funnel to dip just below the surface of water to a depth of an inch in the wide-necked flask (Fig. 50). Water cannot reach the hot test tube, for the end of the funnel is uncovered before any of the liquid that may enter during the experiment is able to rise beyond the thistle-shaped head.

The ammonia solution thus prepared is similar to that sold in shops for household purposes. It has the same smell, and a Litmus paper dipped into it turns blue. Chemically it is ammonium hydroxide, formed by the union of ammonia with water; but it splits up on warming, ammonia gas being given off. When a few drops of the solution are added to dilute solutions of metallic compounds such as Copper Sulphate (No. K108), Nickel Ammonium Sulphate (No. K120) and others contained in the outfit, precipitates of various colours are obtained.

Acid Prepared from Saltpetre

Saltpetre or nitre contains nitrogen, and unlike the element itself, is chemically very active. It is used in making gunpowder, and we have already employed it to make sulphur and carbon burn, for it contains oxygen with which it readily parts.

Saltpetre is a salt of a strong acid known as nitric acid and its chemical name is potassium nitrate. Place two measures of

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Potassium Nitrate (No. KI24) in a dry test tube. Add two measures of Sodium Bisulphate (No. KI25) and a few drops of water, and heat. Brown fumes are given off, and a moist blue Litmus paper placed in them turns red. The fumes are those of nitric acid, turned out of the saltpetre by the action on it of the stronger acid that is derived from the Sodium Bisulphate.

The new acid may be prepared by condensing the fumes. For this purpose fit the test tube with a small bored cork carrying the large right angle delivery tube, the lower end of which dips into a second test tube cooled by standing it in a vessel containing water. Heat the mixture and a small quantity of a pale yellow liquid that turns Litmus paper red and Congo Red paper blue then collects in the second test tube. This is nitric acid.

Place one measure of Copper Turnings (No. KI09) in a test tube and add a few drops of nitric acid. The metal dissolves to form a blue solution, and red fumes are seen in the test tube. These fumes consist of a compound of nitrogen and oxygen known to the chemist as nitrogen peroxide.

The blue liquid contains copper nitrate, and a green solid is left when it is evaporated to dryness. Heat this further and it turns black, more red fumes being given off. The black residue is copper oxide, formed by the decomposition of copper nitrate.

Similar changes take place when nitric acid acts on Zinc (No. KI34) and zinc nitrate

becomes zinc oxide when heated. As we know from previous experiments, this oxide is yellow when hot, but always becomes white again when allowed to cool.

Heat four measures of Lead Nitrate (No. KI13) in a small dry test tube. The crystals break up with a crackling noise and brown fumes are seen in the tube. The residue is yellow and is lead oxide.

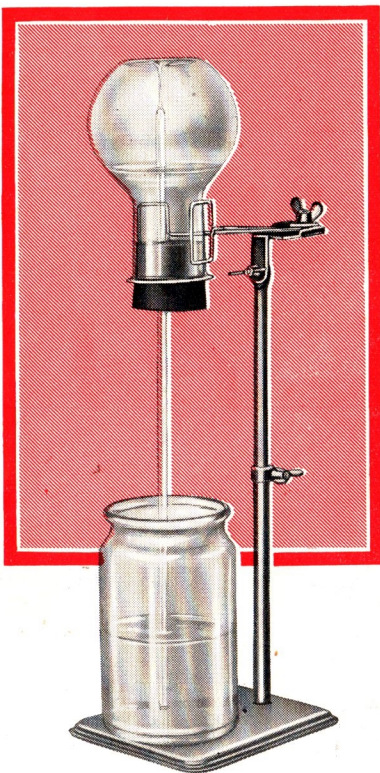


Fig. 48. An ammonia fountain at work. Red Litmus solution in the jar becomes blue when it enters the flask.

10. SULPHUR—THE VOLCANIC ELEMENT

Sulphur or brimstone is one of the most interesting of the elements. Its second name means the burning stone, and it is found in the craters of volcanoes. The gases rising from a volcano in eruption contain fumes of burning sulphur, and this fact, together with the ease with which the element takes fire, the deadly pallor of its pale blue flame, and the choking smell of the fumes produced, have given Sulphur sinister associations as the chief ingredient of supernatural fires! In reality it is a very useful element, and is employed in making matches and gunpowder, and in the rubber and other industries.

Sulphur is a yellow solid, but readily melts when heated, and the liquid boils at a temperature of 450°C.

Like carbon, this element is capable of assuming disguises, and some of the different forms it takes are best made by heating it. Place six measures of Sulphur (No. KI31) in a small dry test tube and heat slowly. The element melts to form a pale yellow mobile liquid, and on further heating becomes darker and darker in colour until eventually it is deep reddish brown. At this stage it is

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Fig. 49. Passing ammonia gas into water in order to form a solution.

syrupe, and does not fall out even when the tube is inverted. Continue to heat, and the liquid, still dark red in colour, again becomes mobile. Pour it into water in the evaporating dish or a saucer. Note that during the experiment part of the Sulphur has been changed into a vapour, which has cooled in the upper part of the test tube, on the sides of which a pale yellow powder is visible.

The form of sulphur obtained by pouring the hot liquid into water is very curious, for it may be pulled and stretched like a piece of indiarubber. It loses its elasticity in a few hours, changing automatically into a hard brittle mass that is seen to consist of tiny eight-sided crystals when examined by means of a powerful magnifying glass or a microscope.

A Novel Printing Process

An interesting experiment with sulphur is reproducing pictures and type from newspapers, for in certain conditions printers' ink adheres to it readily. Surround the portion of a picture that is to be reproduced with a ring of cardboard, which is best obtained by removing the bottom from a small cardboard box. Heat Sulphur in a dry test tube until a pale yellow liquid is obtained, and pour this into the ring (Fig. 51). On cooling remove the

cardboard, and the picture will be found to be transferred to the lower side of the Sulphur medallion formed within it. The picture is of course reversed, but otherwise is a replica of that in the newspaper.

The Fumes of Burning Sulphur

Burn one measure of Sulphur on an old spoon and cautiously smell the gas formed. The gas is sulphur dioxide, and its peculiar choking smell enables it to be recognised easily. Sulphur dioxide is more easily obtained from Sodium Bisulphite, a salt of sulphurous acid, the weak acid formed by dissolving the gas in water. Place two measures each of Sodium Bisulphite (No. KI26) and Sodium Bisulphate (No. KI25) in a dry test tube. Add a few drops of water and a gas is given off. Smell cautiously in order to verify that this is sulphur dioxide.

Dilute sulphuric or hydrochloric acid, or even tartaric acid or vinegar, give more satisfactory results than Sodium Bisulphate (No. KI25), and in order to prepare the gas in larger quantities for further attractive experiments a quarter of the Sodium Bisulphite

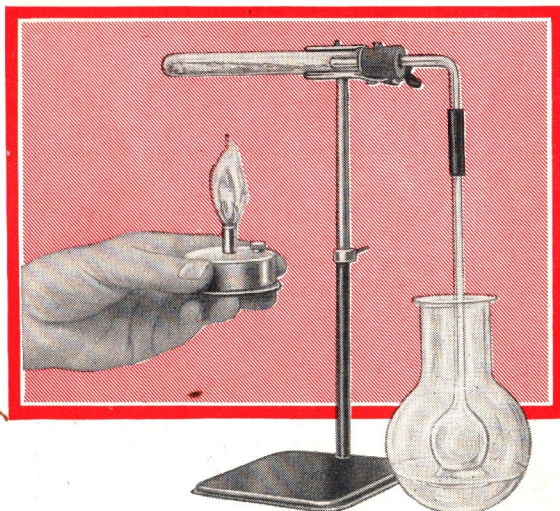


Fig. 50. Preparing a solution of a gas that is easily soluble in water. The use of the thistle funnel prevents the water from rising into the test tube when the production of gas ceases.

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(No. KI26) remaining is placed in the gas generating apparatus and covered with water. Pouring a few drops of an acid, preferably dilute hydrochloric acid, down the thistle funnel immediately produces a violent effervescence. In the following experiments the gas is not collected but is passed through different solutions by means of the large right angle delivery tube, and only a few drops of the acid should be added at a time in order to prevent the action from becoming too violent. In all experiments of this kind a steady stream of the gas produced is to be preferred to a vigorous effervescence.

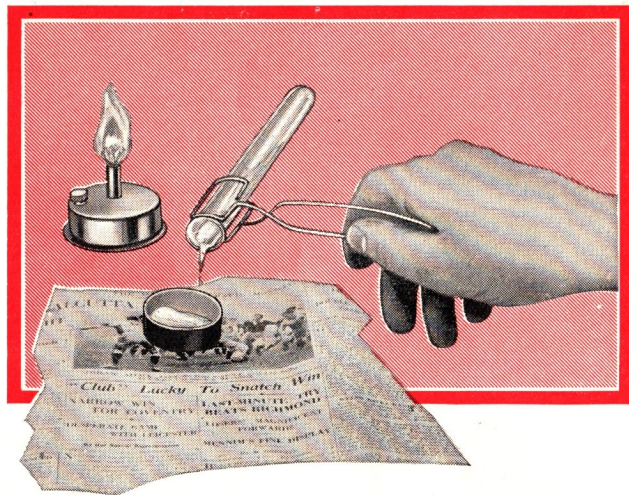


Fig. 51. Reproducing a newspaper illustration on a sulphur medallion.

Bleaching with Sulphur Dioxide

The most remarkable property of sulphur dioxide is its bleaching action. A Litmus paper left in the solution formed by passing the gas through water becomes white. Similarly water containing a few drops of Condy's Fluid, or that has been made red by adding Logwood solution, is decolorised when sulphur dioxide is bubbled through it. Petals of roses and other coloured flowers also are bleached by steeping them in the solution of the gas in water.

The World's Most Important Acid

In a previous experiment we prepared sulphurous acid by dissolving sulphur dioxide in water. This may be changed into a far more important chemical, known as sulphuric acid, by adding oxygen to it. This cannot be done directly, but the necessary oxygen

may be obtained by passing sulphur dioxide from the gas generating apparatus into a test tube half full of hydrogen peroxide, which may be bought from a chemist. Allow the gas to bubble through the liquid for about five minutes, and then test the solution with blue Litmus paper. This is turned red, showing the presence of an acid.

The new acid is stronger than sulphurous acid. It is one of the world's most important chemicals, and is certainly its most important acid, for it is used in an astonishing range of industrial

operations, from the manufacture of dyestuffs, explosives and other chemicals, to the production of fertilisers and even of foodstuffs.

Add a solution of washing soda drop by drop to the solution of sulphuric acid prepared in the last experiment until no more gas is produced. Then evaporate the solution to dryness. The white powder left is sodium sulphate.

Why Silver Egg Spoons Become Stained

Heat a mixture of equal parts of the Sodium Bisulphate (No. KI25), and crushed washing soda, on the charcoal block, holding this so that the material is in the inner portion of the blowpipe flame. Crush the dark-coloured mass obtained, and place a little of it on a silver coin. Dip a glass tube with one end drawn out to a jet into vinegar, tartaric acid, or Sodium Bisulphate solution, close the upper end with the finger and lift it out of the

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liquid. A tube of this kind is known as a pipette or dropping tube, and its contents fall out drop by drop on releasing the pressure of the finger. Allow the acid liquid to fall on the substance placed on the coin (Fig. 52). Tiny bubbles become visible, and an unpleasant smell is noticed. In addition, a black stain that cannot be washed off with water is formed on the coin. Repeat the experiment with the sodium sulphate prepared in the previous experiment.

The gas with the unpleasant smell is hydrogen sulphide, or sulphuretted hydrogen as it is more usually named, and it is a compound of hydrogen and sulphur that is also produced when eggs become bad. In our experiment the coin is stained because a film of silver sulphide, a black insoluble compound of silver and sulphur, is formed on it. The production of this compound accounts for the black stain sometimes seen on silver egg spoons, especially when they have been used for eggs of doubtful quality!

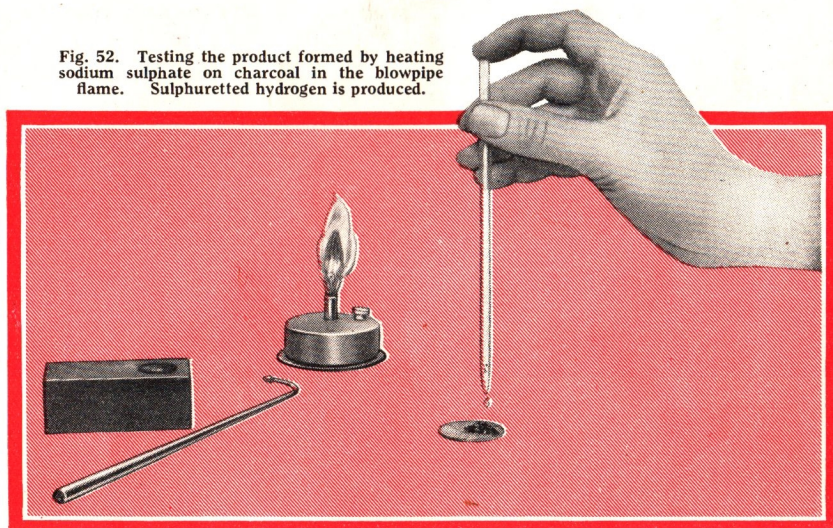
Sulphuretted hydrogen acts like a weak acid to form salts called sulphides, and in this experiment was displaced from sodium sulphide, which was produced by heating Sodium Bisulphate on charcoal. Sulphides may be prepared in other ways. Place two measures each of Calcium Oxide (No. KI03) and Sulphur (No. KI31) in a test tube half filled with water and boil for a few minutes. Pour on a silver coin a few drops of the dark yellow liquid obtained, and the coin becomes coated with a black

film of silver sulphide. The Calcium Oxide and Sulphur have combined on boiling to form calcium sulphide, and this has acted on the sulphur to form the black stain.

A Test for Sulphuretted Hydrogen

We cannot always use a silver coin to test for sulphuretted hydrogen, but a similar stain is formed on a strip of blotting paper that has been dipped in a solution of one measure of Lead Nitrate (No. KI13) in a test tube one quarter filled with water. Add a few drops of dilute hydrochloric acid to the remainder of the calcium sulphide solution, and hold a Lead Nitrate test paper in the gas given off. It immediately becomes coated with black lead sulphide, which has a peculiar metallic glitter that enables it to be recognised without difficulty when making tests of this kind.

Fig. 52. Testing the product formed by heating sodium sulphate on charcoal in the blowpipe flame. Sulphuretted hydrogen is produced.



Certain sulphides are formed by the direct union of sulphur with a metal. Iron sulphide is one of these. Place half the quantity of Iron Filings (No. KI12) contained in the Outfit in a small tin lid, and hold this in a pair of tongs over a glowing fire for a few minutes. Then lift the tin lid off the coals, and sprinkle over it in successive small quantities about half the remaining Sulphur (No. KI31).

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Fig. 53. Preparing a solution of sulphuretted hydrogen.

Part of the sulphur burns with an intense blue flame, but the rest melts and combines with the Iron Filings to form iron sulphide. Re-heat the Iron Filings at intervals during this experiment, and finally hold the tin lid on the fire until no more blue flames are seen, showing that the excess of sulphur has been burned away.

Allow the tin lid and its contents to cool and then place one measure of the black residue in a test tube, cover with one measure of Sodium Bisulphate (No. K125), and add a few drops of water. The gas then given off has the familiar smell of sulphuretted hydrogen and darkens Lead Nitrate test paper. Keep the remainder of the iron sulphide for a later experiment.

In order to prepare a solution of sulphuretted hydrogen for experimental purposes, place the iron sulphide prepared in a previous experiment in the gas generating apparatus, to which is attached the large right angle delivery tube, dipping into a test tube half filled with water. Pour water down the thistle funnel in order to cover the foot of the tube, and then add dilute hydrochloric acid, or if this is not available, a strong solution of Sodium Bisulphate (No. K125). The sulphuretted hydrogen produced bubbles through the water and dissolves in it. In order to obtain a strong solution have a small bored cork on the delivery tube, as shown in Fig. 53, and occasionally disconnect

the generating apparatus. Close the test tube by means of the cork and shake in order to bring the gas that has not been dissolved into closer contact with the water.

With a glass rod place drops of the solution on blue Litmus paper, which is turned red, and Lead Nitrate paper, on which appears the well known glittering black stain.

This solution of sulphuretted hydrogen may be employed in precipitating sulphides, most of which are insoluble in water. Dissolve one measure of Nickel Ammonium Sulphate (No. K120) in a test tube one quarter full of water and add a little sulphuretted hydrogen solution. A thick black precipitate of nickel sulphide is formed.

Repeat the experiment with Iron Alum (No. 111), Manganese Sulphate (No. K119), Cobalt Chloride (No. K105), Copper Sulphate (No. K108) and Lead Nitrate (No. K113), and note the colours of the precipitates.

An interesting alternative way of preparing insoluble sulphides is by passing the gas through solutions prepared in the manner already explained. Take care to wash the delivery tube thoroughly after each preparation. Try also the effect of passing the gas through solutions of these chemicals to which a few drops of dilute hydrochloric acid have been added, and note which sulphides are not precipitated in these conditions.

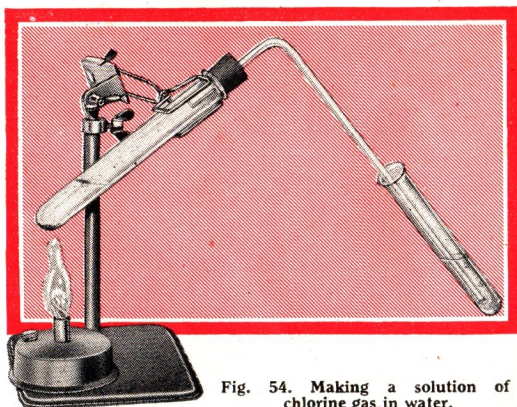


Fig. 54. Making a solution of chlorine gas in water.

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11. A GAS THAT BLEACHES

Heat two measures of common salt with two measures of Sodium Bisulphate (No. KI25) in a small dry test tube, across the mouth of which a moist blue Litmus paper is placed. A fuming gas is given off which turns the Litmus paper red, showing that it is acid. The gas fumes more strongly when brought into contact with moisture by breathing gently across the mouth of the test tube, and thicker white fumes are formed when a glass rod previously dipped in ammonia solution is placed in it.

The gas is hydrochloric acid, which is usually employed in solution in water, a form that we have used in previous experiments. Heat further quantities of common salt and Sodium Bisulphate in a test tube fitted with a small bored cork and the large right angle delivery tube dipping only just below the surface of water in another test tube. The hydrochloric acid gas produced readily dissolves, forming a solution that is acid to Litmus and other indicators.

Pour half of the solution obtained into a clean test tube and drop into it a piece of Magnesium Ribbon (No. KI16) about an inch in length. Hydrogen is given off, and as in previous experiments this gas gives a slight explosion when ignited. Test the remainder of the solution with Granulated Zinc (No. KI34) and Iron Filings (No. KI12), warming slightly if necessary. In each case hydrogen is produced.

Mix two measures each of common salt and Sodium Bisulphate (No. KI25) with two measures of Potassium Nitrate (No. KI24), and heat the mixture in a test tube. By holding the tube against a background of white paper, it will be seen that a greenish gas is produced. Smell this cautiously, and then introduce a piece of moist blue Litmus paper. The paper turns white, a change that has not been observed with any chemical previously tested in these

experiments. The green gas is chlorine. It has a suffocating odour, is neither acid nor alkaline, but bleaches Litmus paper.

Manganese Dioxide (No. KI18) is another



Fig. 55. Absorbing Chlorine in Calcium Oxide.

chemical that liberates chlorine from hydrochloric acid. Heat three measures of Manganese Dioxide, six measures of common salt, and six measures of Sodium Bisulphate, placed in a test tube fitted with a small bored cork and the large right angle delivery tube. Cover with water, add a few fragments of broken plant pot—a device that prevents bumping in the liquid—and heat cautiously. The chlorine may be passed through water (Fig. 54) or into a dry test tube containing six measures of Calcium Oxide (No. KI03). When passing it over Calcium Oxide, cork the test tube at intervals and shake the Calcium Oxide with the gas, which is absorbed (Fig. 55). Keep the product, which is known as bleaching powder, for a later experiment.

An alternative is to heat the dry mixture, the test tube in this case being supported in a horizontal position. The gas may be passed through the wide-necked flask containing six measures of Calcium Oxide, and thence through water, in which it dissolves (Fig. 56). The yellowish solution prepared in this or in the preceding experiment bleaches Litmus paper, red blotting paper and pieces of cloth dipped in it. Cork the tube in order

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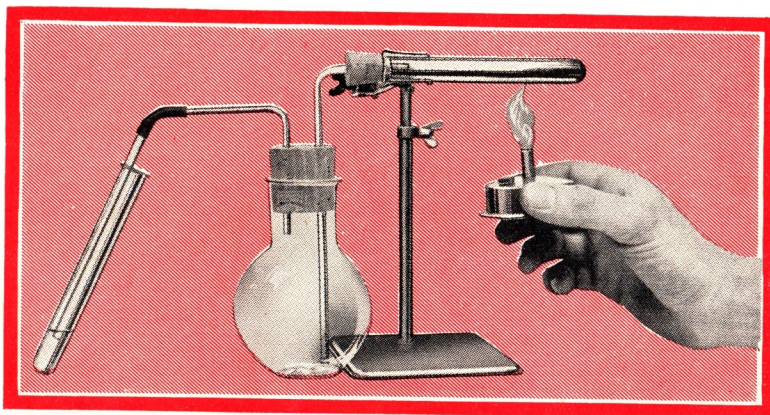


Fig. 56. Passing chlorine over Calcium Oxide in order to form bleaching powder. Unused chlorine is dissolved in water in the test tube.

to preserve the liquid for further experiments.

Fill a dry test tube with chlorine and cork it after dropping into it a strip of printed paper that has been soaked in ordinary writing ink. The writing ink is bleached almost immediately, but the printers' ink is not affected. The reason for this difference is that ordinary ink contains vegetable colouring matter, while printers' ink is largely composed of lampblack, which is an impure form of carbon and is not affected by chlorine.

Chlorine is used on a large scale in industry for bleaching. For this purpose it is made by heating pyrolusite, or manganese dioxide, with strong hydrochloric acid, and absorbed in slaked lime, forming bleaching powder, a white powder that gives up its chlorine readily when dilute acid is added to it.

Materials to be bleached are therefore first soaked in a vat of dilute acid and then in a second vat containing bleaching powder solution. The chlorine is produced on the material to be bleached, for the acid and the bleaching powder come into contact there, and thus the gas is used most economically.

A sample of bleaching powder has been prepared in a previous experiment. Dissolve this in a tube half full of water and prepare a second solution by using an equal quantity of

Tartaric Acid (No. KI33). Soak pieces of coloured cloth or the coloured petals of a flower in the acid solution for about five minutes, and then transfer them to the bleaching powder solution. There they are slowly bleached by the chlorine and in time become white.

A Chemical Family

Chlorine is one of an interesting group of four elements that form a chemical family, for in spite of difference in appearance, they behave similarly from a chemical point of view. The members of this family are fluorine, chlorine, bromine and iodine. Fluorine is a yellow gas that is so active chemically that few metals can withstand its attack. Bromine is a deep red liquid that gives off irritating vapours. Iodine is a dark grey crystalline solid that becomes a violet vapour on heating, and is a useful antiseptic, a solution of it in alcohol forming tincture of iodine, which is used for dressing cuts and wounds owing to its power of destroying germs.

Add a few drops of the chlorine solution previously prepared to a test tube containing Potassium Iodide solution (No. KI23) to a depth of an inch. The colourless liquid becomes reddish brown, because the chlorine displaces iodine from the solution.

Iodine undergoes a very interesting change with starch. Mix two measures of crushed starch with a few drops of cold water into a paste. Heat half a test tube full of water to boiling point and add to it the starch paste. This gives a clear solution of starch. Allow the solution to cool and add a few drops of the liquid containing iodine. The starch paste becomes deep blue in colour. Warm the solution and the blue colour disappears, to reappear on cooling. The blue substance

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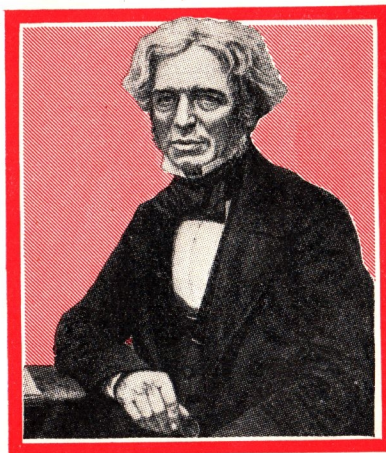
is starch iodide, and its formation is a test for iodine.

A similar result follows the addition of a drop or two of tincture of iodine to starch paste, and this liquid therefore may be used as a test for starch in plants. Cut a potato in two and allow a drop of the tincture to fall from a pipette on to the freshly exposed surface (Fig. 57). As the drop spreads it becomes blue in colour.

A fragment of potato boiled for a minute or two in a test tube containing water to a depth of an inch gives a starch solution that becomes blue on the addition of a drop of iodine solution. Make similar tests with a measure of flour, a similar quantity of bread crumbs, and a few grains of rice. Test also ripe and unripe apples by adding a drop of iodine solution to their freshly cut surfaces. Unripe fruits contain starch, but this is changed into sugar on ripening.

Writing with Electricity

Potassium Iodide solution (No. K123) may be used for electrical writing. For this purpose soak a piece of blotting paper in the solution and lay it on a flat metal plate, such as a Meccano $5\frac{1}{2}'' \times 2\frac{1}{2}''$ Plate, connected by means of copper wire to the negative



Michael Faraday (1791-1867).

terminal of an accumulator or flash-lamp battery (Fig. 58). Then write or draw slowly on the paper with the end of a piece of copper wire

leading from the positive terminal of the source of electricity employed, and the words or design will be traced out. The point of the wire should be blunt, and only the lightest pressure used in order to avoid tearing the wet paper and causing a short circuit.

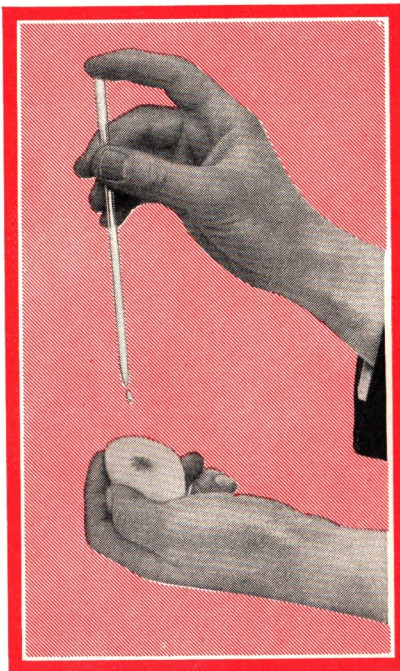


Fig. 57. Testing the freshly cut surface of a potato for starch.

The colour of the letters or lines in this experiment is brown, for the action of the current is to set iodine free, and this gives a brown stain in the presence of Potassium Iodide solution. Blue writing is obtained if the paper is soaked in a mixture of Potassium Iodide solution and thin starch paste, prepared as already explained.

12. ANALYTICAL TESTS

An interesting chemical that is of great importance in the household and in industry is borax, which the chemist calls sodium borate. It is found naturally in many places, notably the bed of a dried up lake in Death Valley, California, and is used in making glazes and enamels, and in the manufacture of soap.

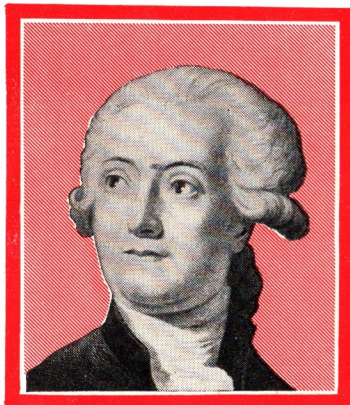
Borax is useful also to the analyst, who employs it for detecting certain metals by the

Kemex Chemical Experiments

colour of the glasses formed when materials containing them are melted with it. Only small beads of glass are made in tests of this kind, which therefore are usually described as borax bead tests.

The beads are made in a loop at the end of a piece of wire, about 3 in. in length, fixed into a short glass tube that serves as a handle. One end of the glass tube is drawn out to a jet. The end of the wire is placed just within the jet, which is then heated until it softens and closes round the wire. On cooling this is firmly held in position, projecting from the glass tubing as a knife blade does from its haft.

After fitting the Nickel Chrome Wire (Part No. K39) into a short length of glass tube in this manner, twist the end into a loop about $\frac{1}{4}$ in. in diameter. Heat the loop in the flame and dip it into a small quantity of borax, placed on a sheet of paper in a convenient position. The borax adheres to the loop, and on heating first swells up as it loses water of crystallisation, and then melts down to form a clear glass bead within the loop. It may be necessary to dip the wire into borax a second time in order to obtain a good bead, and the heating must be continued until a clear colourless glass bead is obtained (Fig. 59).



A. L. Lavoisier (1743-1794).

Touch a minute crystal of Cobalt Chloride (No. K105) with the hot bead and reheat. The bead becomes deep blue in colour owing to the formation of blue cobalt borate. The colour is best seen by using only a very minute quantity of Cobalt Chloride.

The production of a blue borax bead from an unknown chemical helps an analyst to detect the presence of cobalt. Similarly nickel borate glass is reddish brown, while copper, iron and manganese compounds give beads that are green, yellow and pale rose respectively. Test these in turn, cleaning the loop after each test by melting fresh borax on it and shaking off the bead while it is warm.

Flame Tests Used in Analysis

The Nickel Chrome Wire may be used for interesting flame tests that also form part of the regular process of chemical analysis. Clean the wire by dipping it into a test tube containing dilute hydrochloric acid, and then take up on it a small quantity of Strontium Nitrate (No. K.130). Heat the end of the wire strongly in the flame, which is coloured a brilliant carmine red. The colour may be renewed when it fades by dipping the end of the wire in the dilute acid and taking

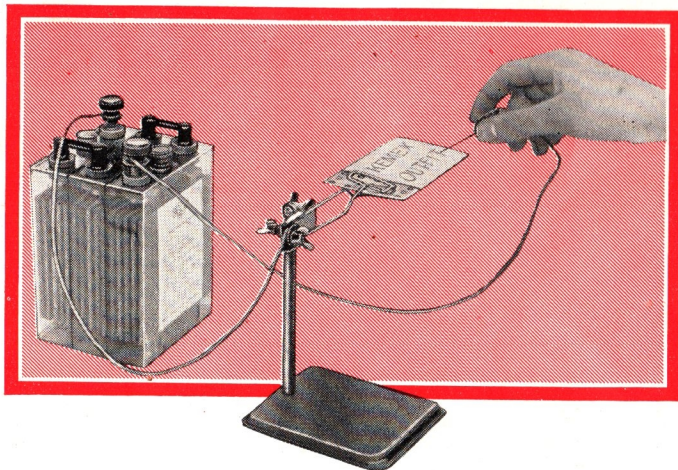


Fig. 58. Writing with electricity.

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up more Strontium Nitrate on the end of it. Test the colours given in this manner by Calcium Oxide (No. K103), Copper Sulphate (No. K108), Lead Nitrate (No. K113) and Potassium Nitrate (No. K124). It will be found that these chemicals give flame colours that are brick red, greenish blue, pale blue and violet respectively. The violet coloration given by Potassium Nitrate is sometimes masked by a yellow coloration, but shows reddish when looked at through a piece of blue glass, for this has the effect of cutting off the yellow rays of light that cause the interference.

The yellow coloration is due to sodium, the metal present in common salt and washing soda. Test common salt in the flame in order to be able to recognise this coloration, the strange light of which gives a ghastly appearance. Sodium compounds are widespread and even dust may give this flame colour.

13. CHEMICAL GARDENS AND DIFFUSION EXPERIMENTS

Nearly half of the crust of the Earth consists of oxygen in a free state in the atmosphere, or in combination with other elements in the oceans and in rocks and soil. The second most abundant element is one that is seldom seen. Its name is silicon, and it is contained in sand, quartz, flint, clay, marl and many other well-known materials. Sand, quartz and flint are forms of silica, a compound of silicon and oxygen. Silica is used in making glass, which is formed by melting sand or flint with limestone and various other chemicals. The kind of glass produced depends on the proportions of its constituents, and glass for special purposes is made by adding other ingredients.

Glass that Dissolves in Water

When clean sand is melted with sodium carbonate, water-glass is formed. This is a

thick glassy-looking syrup that dissolves in water, and is sold in tins for use in preserving eggs. Many interesting experiments may be made with it.

Insoluble silicates are obtained as precipi-

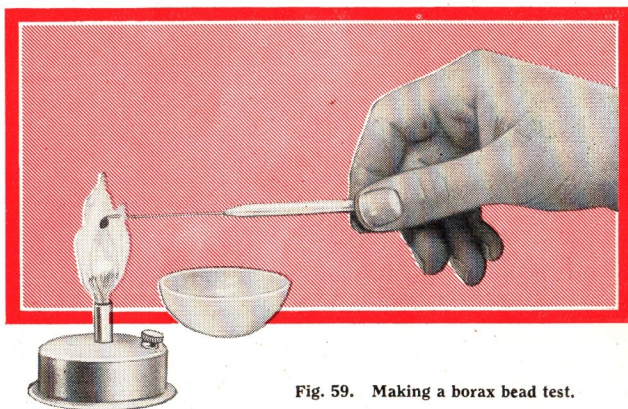


Fig. 59. Making a borax bead test.

tates by adding a solution of water-glass to solutions containing various metals. Dissolve one measure of Cobalt Chloride (No. K105) in a test tube containing water to a depth of an inch, and add a few drops of a solution of water-glass prepared by adding a teaspoonful of the syrup to a test tube half full of water. A beautiful blue precipitate of cobalt silicate is formed. Other coloured silicates are prepared in a similar manner from solutions of the compounds of copper, nickel, manganese, iron, aluminium, strontium and magnesium contained in the Outfit.

Growing Cobalt Silicate Columns

A fascinating way of preparing the insoluble silicates of these metals is to drop crystals of their soluble salts into a solution of water-glass. For this purpose prepare sufficient solution of water-glass to fill a large glass jar, using the syrup in the proportion of one tablespoonful to a tumblerful of water. It is best to use water that has been boiled and allowed to cool. Place the jar in a position in which it will not be disturbed or shaken, and drop into the solution two or three small crystals of Cobalt Chloride (No. K105). A

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slender column of cobalt silicate then grows on each crystal. Usually a small bubble may be seen at the top of each rising column, and in a few hours the growth becomes wider and taller, eventually reaching the surface of the solution. A few very small crystals dropped gently on to the surface of the water-glass solution remain afloat and the growths from these proceed downward to meet those ascending from the heavier crystals at the bottom.

"Plants" in a Chemical Garden

Other chemicals give rise to growths of this kind, and Figs. 60 and 61 show "chemical gardens" produced in this manner by "planting" Copper Sulphate (No. K108), Nickel Ammonium Sulphate (No. K120), Manganese Sulphate (No. K119) and Magnesium Sulphate (No. K117) in addition to Cobalt Chloride. In forming a chemical garden it is best to put in crystals of Copper Sulphate and Nickel Ammonium Sulphate first, for the silicates of copper and nickel do not grow very quickly, and to follow them two days later with one measure each of Manganese Sulphate and Magnesium Sulphate, crystals of Cobalt Chloride being added when the "plants" already in the garden are well developed. The growths



Fig. 60. A chemical garden in which the "plants" have grown from crystals of Magnesium Sulphate and other chemicals.



Fig. 61. Cobalt Chloride growths in a chemical garden.

vary in colour from the white and light brown strands of magnesium and manganese silicates, to the blue and green of the heavier and shorter columns rising from Copper Sulphate and Nickel Ammonium Sulphate; while the remarkably deep blue cobalt "plants" adds to the attraction of the fascinating underwater scene.

Experiments on Diffusion

Twist a sheet of parchment paper, of the kind used by provision merchants and butchers, into the form of a bag, and after wetting it, pour in a dilute solution of Copper Sulphate (No. K108). Then support the bag so that the bottom of it dips into water in a basin. The water on the outside of the parchment bag slowly acquires a blue coloration, showing that the Copper Sulphate is passing through the membrane.

Another form of diffusion apparatus is shown in Fig. 62. The parchment paper is fastened round the wide end of a thistle funnel, which dips just below the surface of water in the wide-necked flask or a basin.

The passage of the Copper Sulphate through the parchment membrane is called diffusion, and is explained by supposing that the small particles in the solution are in continuous

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movement. They do not move very rapidly, and as they are packed fairly tightly together in the liquid they cannot move far in one direction before colliding with other particles and being turned aside. Particles near the edge escape from the mass, however, and make their way through the minute pores of the parchment paper.

A simple experiment that also illustrates the diffusion of Copper Sulphate in solution is carried out by dissolving six measures of Copper Sulphate in a depth of $\frac{1}{2}$ in. of water in a test tube, warming if necessary to form the solution. After cooling, allow water to trickle from a pipette slowly down the side of the test tube, where it forms a colourless layer above the deep blue solution. On standing, the boundary between the two layers becomes indistinct owing to the slow diffusion of the Copper Sulphate particles, and in a day or two the whole of the liquid becomes blue in colour.

On a larger scale the same experiment may be carried out in the wide-necked flask, or a tumbler, and the water is then best added by pouring it gently down a glass rod on to a flat piece of cork floating on the blue solution (Fig. 63). The cork must not be removed after adding the water, as this would mix up the two layers of liquid.

Sugar Separated from Starch

A very convenient form of apparatus for experiments in diffusion may be made from one of the round cardboard boxes used for holding cheese. Cut out the disc portion of both the box and the lid. Stand the cardboard ring thus obtained from the box upside down and

place on it a circular sheet of parchment paper about an inch larger all round. Then put on the remains of the lid so as to hold the paper in position. This gives a shallow drum, with a parchment base that can be floated on water in a basin or a photographic dish of suitable size.

When solutions of such common substances as sugar and salt are placed in the drum they slowly diffuse through the parchment, and their presence in the outer liquid may be recognised by their taste. There are substances that do not pass through a membrane of this kind, however. Starch is one of these. Mix half a teaspoonful of starch into a smooth paste with water in a cup and then fill the cup with boiling water, stirring meanwhile. Add enough sugar to make the liquid noticeably sweet, and after cooling pour some of it into the drum and float this on water (Fig. 64).

The presence of sugar in this liquid is shown by the taste. Test for the starch by adding a drop of the liquid, taken out by means of a glass rod, to a drop of tincture of iodine placed on a clean plate. The deep blue coloration obtained is a proof of the presence of starch. Repeat the test immediately with a drop of the liquid in which the drum is floating, taking care to rinse the glass rod before use. No blue coloration is obtained, showing that no starch is present, and the absence of sugar is proved by tasting. Thus at the beginning of the experiment the liquid in the drum contains both sugar and starch, whereas the liquid on the other side of the drum is water only.

Repeat the tests with the outer liquid a few hours later. The presence of sugar is readily detected,



Fig. 62. Copper Sulphate solution diffusing through parchment paper.

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but no blue coloration is obtained with iodine, showing that no starch has passed through the parchment. Continue the experiment and test the outer liquid repeatedly for starch. No result will be obtained, showing that although sugar passes freely through the membrane, the starch does not. By continuing the experiment and changing the water in the outer vessel frequently, practically the whole of the sugar may be removed from the solution in the drum, leaving the starch behind.

14. EVERYDAY CHEMISTRY

Many attractive tests and experiments with familiar materials may be carried out with the aid of apparatus and chemicals included in Kemex Outfits. For instance the textile fabrics of which our clothing is made may be examined. The chief of these are cotton, wool and silk, and a simple test to enable them to be distinguished consists of heating samples slowly in an old spoon held over a flame. Wool and silk, which are of animal origin, give the acrid smell of burning hair; the vegetable fibre in cotton burns with a similar smell to that of paper and leaves a white ash.

For a further test prepare a solution of caustic soda as explained on page 28. Pour half of this into a test tube or the evaporating dish, drop into it a small piece of woollen cloth and boil. The cloth slowly dissolves. When the experiment is repeated with a small piece of cotton the fabric remains unchanged.

This test enables natural silk to be distinguished from artificial silk. The former

is dissolved, while artificial silk, which is made from wood, is not affected in this manner by caustic soda solution.

Cotton cloths often contain various substances employed as filling or weighting agents. One of these is starch, and this may be detected by boiling a piece of cloth in water in a test tube and adding one or two drops of tincture of iodine. The formation of blue starch iodide in the solution indicates the presence of starch in the fabric.



Fig. 63. Pouring water on a solution of Copper Sulphate without disturbing the lower liquid.

Simple Dyeing Experiments

Dyes have been used for centuries to give an attractive appearance to fabrics. The earliest dyes to be employed were extracted from plants, but to-day the dyer has at his command an immense range of colouring matters manufactured from coal tar products, and chemists have discovered how to employ these in order to give colours that do not fade.

Logwood is an example of a natural dye. It is obtained by boiling Logwood (No. K115) with water in the proportion of two measures to half a test tube full of water. The solution formed is deep red in colour. To small portions of this add solutions formed by dissolving one measure each of Iron Alum (No. K111) and Cobalt Chloride (No. K105) in a little water and solutions of black and red dyes respectively are obtained.

A piece of light coloured cloth may be dyed dark blue by steeping it in a solution of two measures of Sodium Ferrocyanide (No. K128) in a test tube half full of water. After drying, transfer the cloth to a second test tube

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containing two measures of Iron Alum (No. K111) dissolved in a similar quantity of water. It becomes dark blue in colour and retains its colour when dried after steeping for a few minutes. The name of this dye is Prussian Blue.

Cloth may be dyed light blue instead of dark blue as in the above experiment by substituting Ferrous Ammonium Sulphate for Iron Alum.

Make a new solution of Iron Alum (No. K111) and steep two small pieces of woollen fabric in it. Allow these to become dry. Then place one in a solution of Sodium Thiocyanate (No. K129) and the other in a solution of Tannic Acid (No. K132), in each case adding two measures of the chemical to half a test tube full of water when preparing the solution. Colours form on the cloth itself, and on drying the pieces are dyed red and black respectively.

Logwood (No. K115) and Tannic Acid (No. K132) may be employed to dye silk black. Place a small piece of silk for a few minutes in a solution of Iron Alum (No. K111) and allow it to dry. Next steep it in a solution of one measure of Tannic Acid in half a test tube full of water and allow it to dry before boiling it in solution containing two measures of Logwood in a similar quantity of water. After cooling, take out the fabric and wash it. It is then found to be dyed black.

Making Soap from Lard or Butter

Soap is another everyday material with which the chemist is largely concerned, for his tests ensure the purity of the product and its suitability for various purposes.

Soap is made by boiling fats with caustic soda or similar materials. Prepare a solution of caustic soda, add to it a piece of lard or butter about the size of a bean, and boil the liquid for a few minutes, shaking the tube gently all the time in order to prevent the liquid from bumping and spurting out of the tube.

When the butter has dissolved in the hot alkaline liquid, allow the mixture to cool, and heat again for two minutes after adding four measures of common salt. Then pour the mixture into the evaporating dish and as it cools the soap separates out as a solid layer on the surface.

Well-made soap should contain no free alkali, for this is harmful to the skin and also to fabrics that are washed with it. Test various samples of soap by cutting a fresh surface and allowing one or two drops of Phenolphthalein Solution (No. K121) to fall on it. The soap that has been made in the above experiment may give a pink coloration when tested in this manner, owing to the presence of free alkali.

Coloured Writing Inks

A black writing ink may be made by mixing solutions of Tannic Acid (No. K132) and Iron Alum (No. K111), prepared by dissolving one measure of each chemical in one third of a test tube full of water. Thicken the black liquid obtained, which contains iron tannate, by adding a few drops of gum and write with it, using a steel pen. For blue ink substitute Sodium Ferrocyanide (No. K128) for the Tannic Acid.

To make a red ink, boil four measures of Logwood (No. K115) for five minutes in a third of a test tube full of water. Filter and dissolve in the red liquid one measure each of Aluminium Sulphate (No. K100) and Sodium Bisulphate (No. K125).

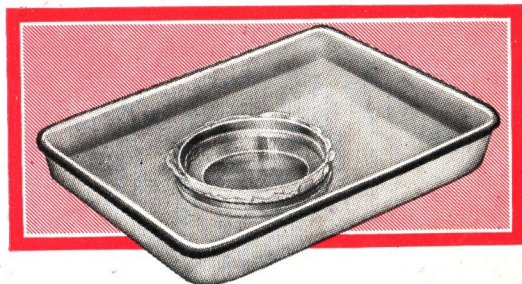


Fig. 64. Separating sugar from starch by means of diffusion through parchment paper.

KEMEX

List of Parts and Contents of Outfits

No.	Description	Quantities in Outfits		
		No. 1	No. 2	No. 3
K1	Test Tube, 5" x 1/2"	4	4	6
K2	" " Heat Resisting, 4" x 1/2"	2	2	2
K3	" " Stand	1	1	1
K4	" " Holder	1	1	1
K5	" " Brush	1	1	1
K6	Funnel	1	1	1
K7	Filter Paper, 3 1/4" diameter	12	50	100
K8	Evaporating Dish	1	1	1
K9	Gauze Square	1	1	1
K10	Flask, Wide-necked	1	1	1
K11	Thistle Funnel	1	1	1
K12	Right Angle Delivery Tube, Small	1	1	1
K13	" " " " Large	1	1	1
K14	Double Angle Delivery Tube	1	1	1
K15	Glass Tube, 12"	1	1	2
K16	" " Rod	1	1	1
K17	Cork, Large, Double Bore	1	1	1
K18	" " Small	1	1	4
K19	" " Bored	2	3	4
K20	Blowpipe	1	1	1
K21	Charcoal Block	1	1	1
K22	Spirit Lamp, Complete	1	1	1
K23	" " Stopper	1	1	1
K24	" " Wick Holder	1	1	1
K25	" " Wick	1	1	1
K26	Universal Stand, Complete	1	1	1
K27	" " Base	1	1	1
K28	" " Pillar	1	1	1
K29	" " Ring	1	1	1
K30	" " Pillar Extension	1	1	1
K31	" " Clamp	1	1	1
K32	" " Top Bracket	1	1	1
K33	" " Wing Nut	1	1	1
K34	" " Washer	1	1	1
K35	Evaporating Stand	1	1	1
K36	Scoop	1	1	1
K37	Rubber Connection Tube	1	1	1
K38	Asbestos Fibre	1	1	1
K39	Nickel Chrome Wire	1	1	1

No.	Description	Quantities in Outfits		
		No. 1	No. 2	No. 3
K40	Instruction Manual, No. 1	1	1	1
K41	" " " " No. 2-3	1	1	1
K42	Universal Stand, Wing Screw	1	1	1
K100	Aluminium Sulphate	1	1	1
K101	Ammonium Chloride	1	1	1
K102	Calcium Carbonate (Marble)	1	1	1
K103	Calcium Oxide (Lime)	1	1	1
K104	Charcoal	1	1	1
K105	Cobalt Chloride	1	1	1
K106	Congo Red	1	1	1
K107	Copper Oxide	1	1	1
K108	Copper Sulphate	1	1	1
K109	Copper Turnings	1	1	1
K110	Ferrous Ammonium Sulphate	1	1	1
K111	Iron Alum	1	1	1
K112	Iron Filings	1	1	1
K113	Lead Nitrate	1	1	1
K114	Litmus	1	1	1
K115	Logwood	1	1	1
K116	Magnesium Ribbon	1	1	1
K117	Magnesium Sulphate	1	1	1
K118	Manganese Dioxide	1	1	1
K119	Manganese Sulphate	1	1	1
K120	Nickel Ammonium Sulphate	1	1	1
K121	Phenolphthalein Solution	1	1	1
K122	Potassium Chlorate	1	1	1
K123	Potassium Iodide Solution	1	1	1
K124	Potassium Nitrate	1	1	1
K125	Sodium Bisulphate	1	1	1
K126	Sodium Bisulphite	1	1	1
K127	Sodium Borate (Borax)	1	1	1
K128	Sodium Ferrocyanide	1	1	1
K129	Sodium Thiocyanate	1	1	1
K130	Strontium Nitrate	1	1	1
K131	Sulphur	1	1	1
K132	Tannic Acid (Tannin)	1	1	1
K133	Tartaric Acid	1	1	1
K134	Zinc, Granulated	1	1	1

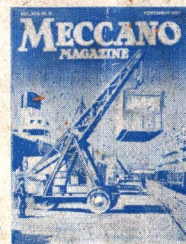
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