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CHEMISTRY

13-16



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Preface

This book is intended to cover the requirements of the various chemistry examinations usually taken at 16+. It is written as a three-year course and takes into account that those starting the course will probably have studied some chemistry in the two years before. For many reasons, the depth of coverage and choice of topics in this period varies a great deal. In an effort to overcome this problem, the first three chapters of this book cover the basic work which it is hoped would have been covered before starting the course at 13. This has been treated mainly as revision but in sufficient depth for someone who has not studied a particular topic to learn the essential background. The main course starts at the beginning of Chapter 4 and from this point a rather different and more detailed treatment is given.

Experiments are not usually described in detail since teachers like to introduce them in their own way. Moreover, it is essential that any student of chemistry must carry out much of this practical work himself, for which reading about experiments is a poor substitute. Experience has shown, however, that many pupils need help in interpreting experimental data and this has been taken into account.

Included at various points in each chapter is a summary of basic information and a list of words or terms to remember; these must be studied carefully. Although modern chemistry involves interpretation of information there is still a need for factual learning.

A number of questions, given in each case at the end of the chapters, help to develop the concepts in each chapter. In some cases they can be answered from the book alone, but more often you will have to use the principles learned to solve a problem. In other cases you may have to develop investigation methods or find out something with the help of books and other publications.

R.W.T.

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The Purification of Substances

1.1 WHAT IS CHEMISTRY?

Ten different dictionaries will probably give ten different definitions of the word “chemistry”. Most of them will, however, suggest that it is the science of substances or materials and few chemists would argue with this statement except perhaps to say that it does not go far enough.

The drawback is that there are millions of different substances in the world; even an apparently simple name like wood covers hundreds of different materials. For example, wood from the cork tree has very different properties from those of wood from the oak. The chemist would be faced with an impossible task if he had to examine each one of these substances separately, and so he tries to break materials up into simpler ones. He can then study the behaviour of each of these in turn. His first task, therefore, is to prepare **pure substances**.

At this stage it is difficult to decide whether a given substance is pure or not and opinions would often vary. If some water, scooped up from close to the sea bed, is examined it usually looks rather murky and may even have a fish or two in it. The fish can easily be removed with the aid of a net, but the remaining liquid will look much the same. To remove the tiny particles of sand etc. left, a fine sieve could be used. If the liquid passing through were poured into a glass tumbler it would probably look clear and colourless, but a small sip would soon show that it is not pure. We obviously cannot rely on our eyes to decide about the purity of a liquid and a chemist would not survive long if he went round tasting all his specimens!

Laboratory chemicals are usually regarded as pure, but even so, some of these are more pure than others. The water from the tap is pure enough for most purposes but needs to be purified further for much scientific work. It might then satisfy most chemists, but it would probably not be suitable for making solutions to be

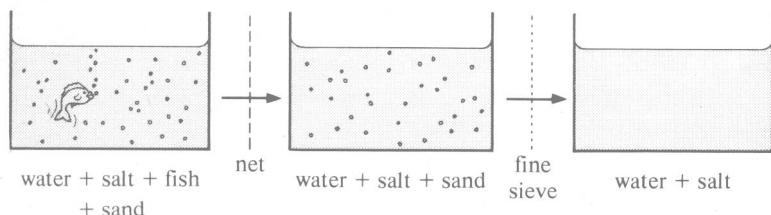


Fig. 1.1 The principles of filtration.

injected into the human body. In most school laboratories there are chemicals which have been specially purchased for certain experiments where a particularly high degree of accuracy is required, but even these may contain small traces of impurities, as you can see from catalogue descriptions or from labels (Fig. 1.2).

The purification of substances can be expensive as you can see from Fig. 1.3, and so there is little point in using a highly purified sample of a substance when the cheaper, less pure, variety may serve the purpose just as well. The term “pure” can obviously mean different things to different people. At this stage of the book it is used to describe a substance which has nothing else mixed with it. There are many widely used methods of purifying chemicals in the laboratory and some of these will now be described.

Sodium carbonate, hydrated (analytical reagent)	
Maximum limits of impurities	
Arsenic	0.0002%
Chloride	0.2%
Iron	0.005%
Lead	0.001%
Sulphate	0.25%

Fig. 1.2

This label, from a bottle of hydrated sodium carbonate, shows that even highly purified chemicals still contain impurities.

Prices of samples of sodium chloride	(£/kg)
Sodium chloride (analytical quality)	1.13
Sodium chloride (pure)	0.75
Sodium chloride (technical quality)	0.30

Fig. 1.3 The difference in costs of three samples of sodium chloride shows that purification is expensive. There is no need to use a very pure chemical when a cheaper, less pure one will do as well.

1.2 FILTRATION

It should be realised from the processes outlined in Fig. 1.1 for “purifying” a sample of sea-water that the smaller the objects to be removed the smaller must be the holes in the sieve. The net which removed the fish did not retain the sand, and presumably the particles of salt which could be tasted in the water at the end had passed through the fine sieve. Even under a microscope these cannot be seen and so they must be very tiny indeed.

A mixture in which the particles of the different substances present are so small that they cannot be distinguished from one another is known as a **solution**. One of the substances in a solution is normally a liquid, and this is known as the **solvent**. The solid dissolving is said to be **soluble** and is called a **solute**. If fine sand is shaken with water, the mixture formed is cloudy and it is easy to distinguish the solid particles from those of the liquid. This type of mixture is called a **suspension** and a solid which does not dissolve in a particular solvent is said to be **insoluble** in that solvent.

In a solution, the solute particles usually differ little in size from those of the solvent and so it would be virtually impossible to produce a sieve fine enough to trap the solute particles while allowing those of the solvent to pass through. In a suspension the solid particles are much larger and so a suitable sieve could remove them. This process is known as **filtration** and, in the laboratory, the “sieve” generally used is a **filter paper**. This is made from a coarse paper in which there are relatively large spaces between the fibres. The liquid particles can pass through these spaces but the solid particles in the suspension are held back. The grade of filter paper normally used in schools will deal with most suspensions but if the solid particles are very fine it may be

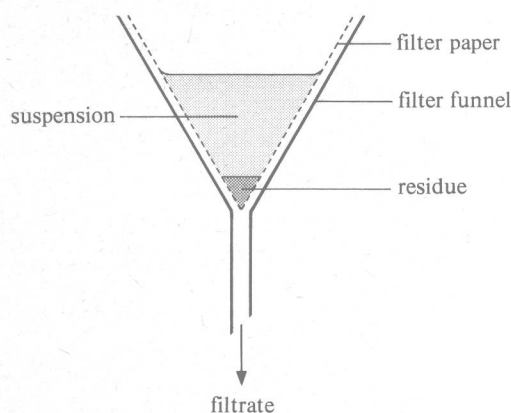


Fig. 1.4 Filtration in the laboratory.

necessary to use another type in which the fibres are more closely packed.

To speed up the filtration process, especially when large quantities of liquid are involved, a **Buchner flask** and **funnel** (Fig. 1.5) are used. The pressure in the flask is reduced by attaching it to a small pump. The liquid is forced rapidly through the filter paper since the pressure below the paper is now less than the atmospheric pressure. This paper is supported on a perforated shelf in the funnel to stop it collapsing. In industry large rotating filters work on a similar principle (Fig. 1.6).

The liquid passing through the paper is known as the **filtrate** and the substance retained is called the **residue**. Some of the liquid may adhere to the residue and, if it is a solution, would contaminate the solid. It can be removed by pouring water over the residue in the funnel, a process known as washing. The residue can then be dried.

1.3 CENTRIFUGING

If a tube containing a suspension is swung round rapidly the contents are forced outwards. The denser particles of undissolved solid pack tightly in the bottom of the tube. This process is known as **centrifuging** and is carried out in the laboratory by means of a **centrifuge** (Fig. 1.7). After centrifuging, the liquid can be separated from the solid either by pouring it off carefully – **decanting** – or by withdrawing it using a **teat-pipette**. The more rapidly the centrifuge spins the more effective is the operation but high-speed centrifuges need to be balanced carefully or they vibrate wildly. Centrifuge heads are usually made to hold two or four test-tubes, and so, if one or three tubes are used for the suspension, the odd one must be counterbalanced by a similar tube of water placed opposite.

Centrifuges are always fitted with covers and these should never be raised while the head is rotating in case a tube has broken, when glass splinters may be thrown out. It is also important to use strong tubes of the right size. The time of centrifuging varies greatly according to the speed of the centrifuge and the sizes of the solid particles. The centrifuge has one big advantage over the use of filter papers; it offers a means of handling corrosive liquids which would attack the paper.

1.4 EVAPORATION

Solutes cannot be removed from solutions by filtration or centrifuging and this type of mixture is usually separated by **evaporation**. This makes use of the fact that the solvent normally boils at a lower temperature than the solute. The solution is placed in a shallow porcelain dish known as an **evaporating basin** and heated gently over a wire gauze to spread the heat more evenly (Fig. 1.8(a)). When very slow and careful evaporation is required the basin can be heated over a steam-bath (Fig. 1.8(b)).

Although this is not a practical textbook a few suggestions on

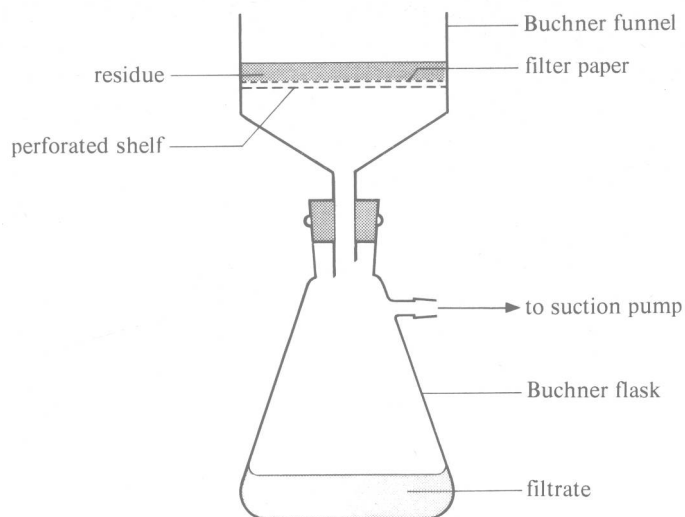


Fig. 1.5 A more rapid method of filtering.

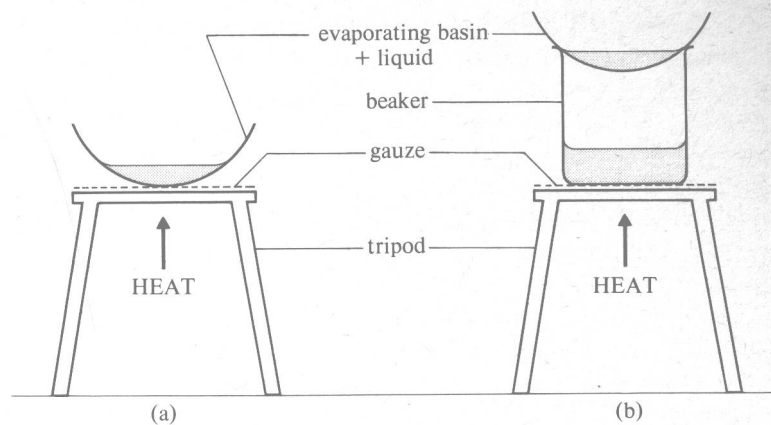


Fig. 1.8 (a) Direct evaporation. (b) Evaporation over a steam bath.

Fig. 1.6 (right)
Filtration on the industrial scale
using large rotary filters.

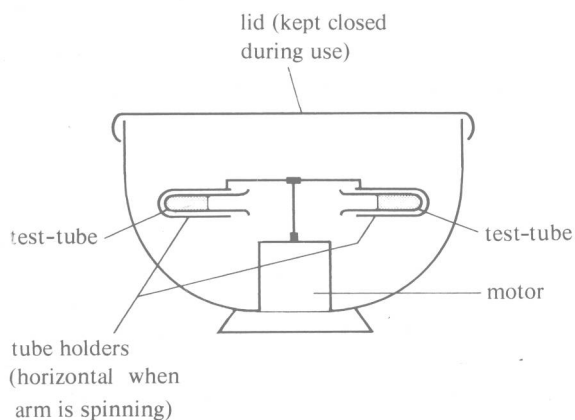
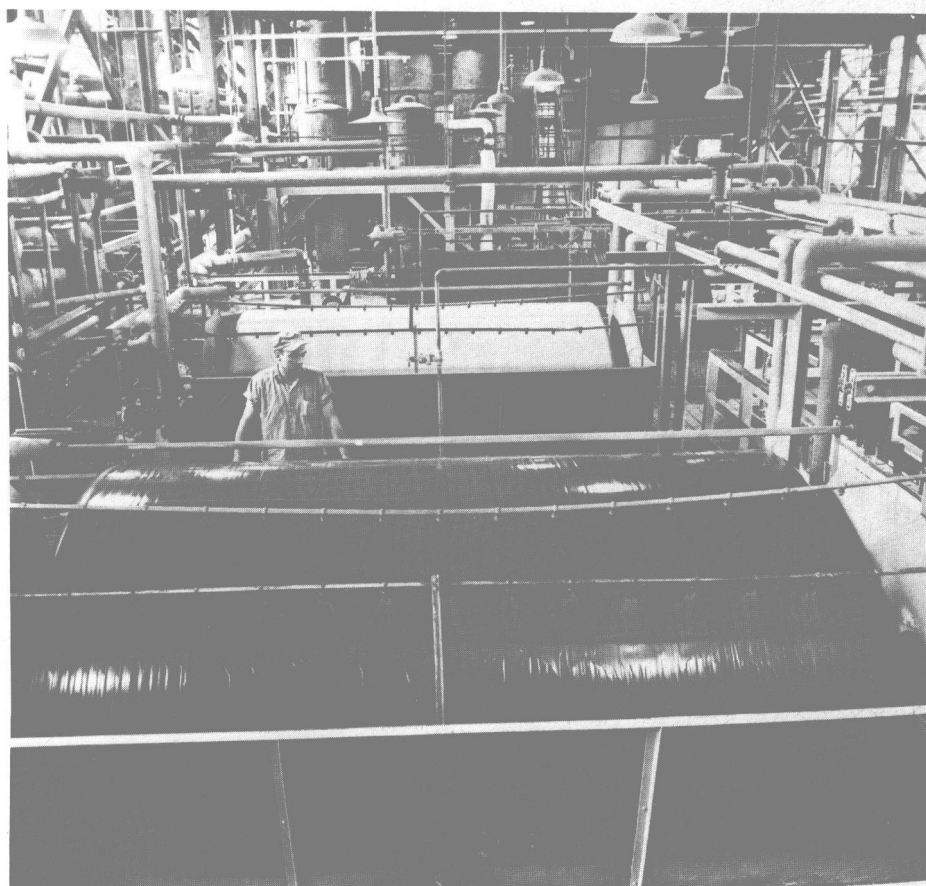


Fig. 1.7 A simplified diagram of a laboratory centrifuge showing the test-tubes in the spinning position.

using the **Bunsen burner** may be useful. Most schools will now be supplied with **natural gas** since this has now largely replaced manufactured **town gas** (Section 2.10). There is little difference in operating the two types, but natural gas burners have finer jets and are a little more difficult to light. When the air-hole is closed the resulting, so-called **luminous**, flame is yellow and gives off soot which makes it unsuitable for heating apparatus. Since this flame is easily visible it is safer to leave the burner with the air-hole closed when you are not actually using the flame, especially in bright sunlight. Moreover, this flame can be turned down low and so is more economical. The flame produced when the air-hole is wide open is the hottest but the flame you will use most is one about four centimetres high obtained with the air-hole about half open (Fig. 1.9).

When evaporating, if the solvent is **flammable**, i.e. it catches fire easily, a naked flame must not be placed directly under the gauze and basin. Another snag of evaporation is that the solvent is "lost" to the air. If the solvent is flammable, expensive, or has a harmful vapour, simple evaporation must be replaced by the more elaborate process of **distillation** (Section 1.6) in most cases.

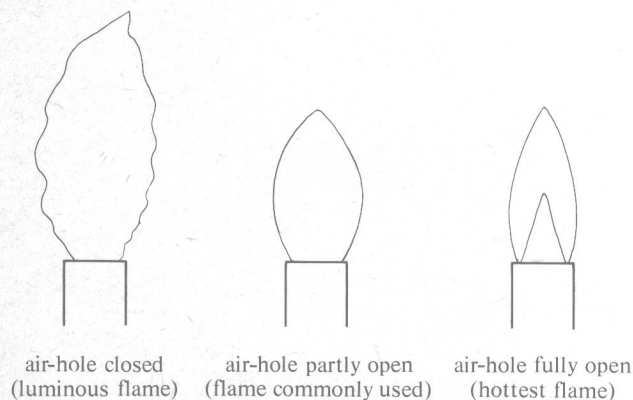
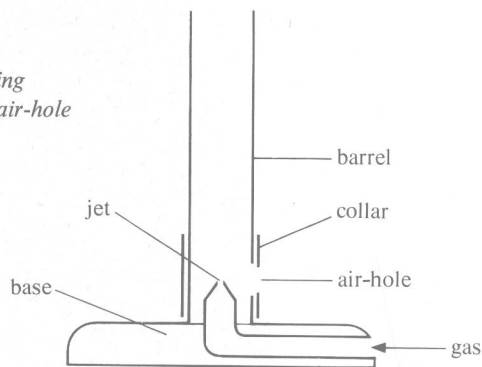


Fig. 1.9
The Bunsen burner showing the effect of opening the air-hole on the flame produced.



1.5 CRYSTALLISATION

If the solution has been concentrated, but not evaporated to dryness, and then allowed to cool, the solid will usually appear as crystals; this process is known as **crystallisation**. **Crystals** are solid particles having flat faces and regular shapes (Fig. 1.10).

Fig. 1.11 shows crystals of pure nickel sulphate. Examine them to see the regularity of shape.

Different solutes dissolve to different extents in the same solvent. The maximum number of grams of a solute which can be dissolved in 100 g of a solvent at a specified temperature is known as the **solubility** (Fig. 1.12). A solution which can dissolve no more solute at that temperature is said to be **saturated**.

Crystallisation can be used to separate mixtures of solids which have different solubilities and the process is then called **fractional crystallisation**.

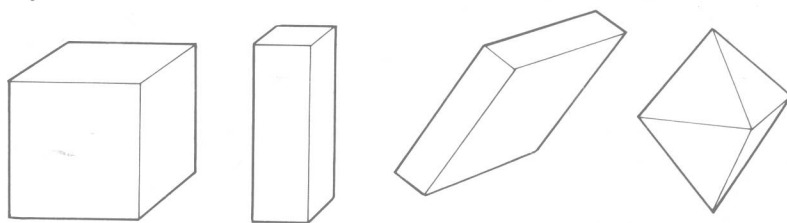


Fig. 1.10 Some simple crystal shapes.



Fig. 1.11
Crystals of pure nickel sulphate.

Solubilities of substances (g/100 g at 25°C)	
Calcium carbonate	0.0013
Calcium hydroxide	0.40
Copper sulphate	34.5
Magnesium sulphate	41.2
Silicon dioxide	insoluble
Sodium chloride	36.0
Sodium hydroxide	42.0
Sodium nitrate	91.8

Fig. 1.12
Solubilities of some common solids at 25°C.

1.6 DISTILLATION

It was explained in Section 1.4 that evaporation is unsuitable in certain situations and then a method of “catching” the vapour and changing it back to the liquid (**condensation**) is required. The complete process is called distillation.

distillation = evaporation + condensation

The apparatus shown in Fig. 1.13 is one commonly used in the laboratory, but there are many possible modifications. For example, the flask could be electrically heated and, in some cases, the water-cooled tube, called a **condenser**, may be replaced by a single tube known as an air condenser.

In a distillation apparatus it is important that the liquid boils evenly; uneven heating may cause the liquid in the flask to jump about and this **bumping** as it is called can become violent and dangerous. To avoid it, a few small glass beads or pieces of broken porcelain are placed in the liquid. Small bubbles of vapour flow from these, so keeping the liquid well stirred and helping it to boil evenly.

When distilling, the thermometer is arranged so that its bulb is level with the outlet from the flask so measuring the temperature of the vapour leaving the flask. As the vapour passes through the condenser it loses some of its heat to the surrounding water (or air) and eventually its temperature drops to the point where it condenses.

If tap-water is placed in the distilling flask and boiled, any dissolved solids remain and the steam given off condenses to reasonably pure water, i.e. distilled water.

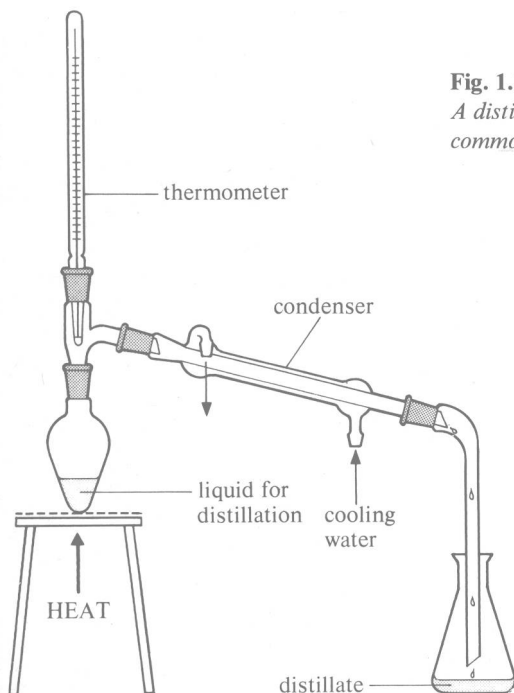


Fig. 1.13
A distillation apparatus commonly used in the laboratory.

1.7 FRACTIONAL DISTILLATION

As described above, distillation is used for obtaining the solvent from a solution, but a slightly modified technique can be used to separate mixtures of liquids which have different boiling points. Two liquids which dissolve in one another are said to be **miscible** whereas ones which do not mix but separate into distinct layers are known as **immiscible** liquids.

A mixture of three miscible liquids (*A* b.p. 50°C, *B* b.p. 100°C and *C* b.p. 150°C) is distilled, using gentle and careful heating. When the temperature of the mixture reaches approximately 50°C liquid *A* boils and its vapour passes over to be condensed and collected in a suitable **receiver**. While *A* is boiling away the temperature remains unchanged, provided too hot a flame is not used, but as soon as *A* has evaporated the temperature begins to rise. At 100°C liquid *B* distils over and can be collected in another receiver. Finally, liquid *C* can be distilled and collected. Thus the three liquids have been collected as three separate fractions; the process is called **fractional distillation**.

In practice, it is difficult to separate a mixture of liquids whose boiling points are close together. To make the separation more efficient a **fractionating column** is included in the apparatus (Fig. 1.14).

Many types of column are used in the laboratory but they

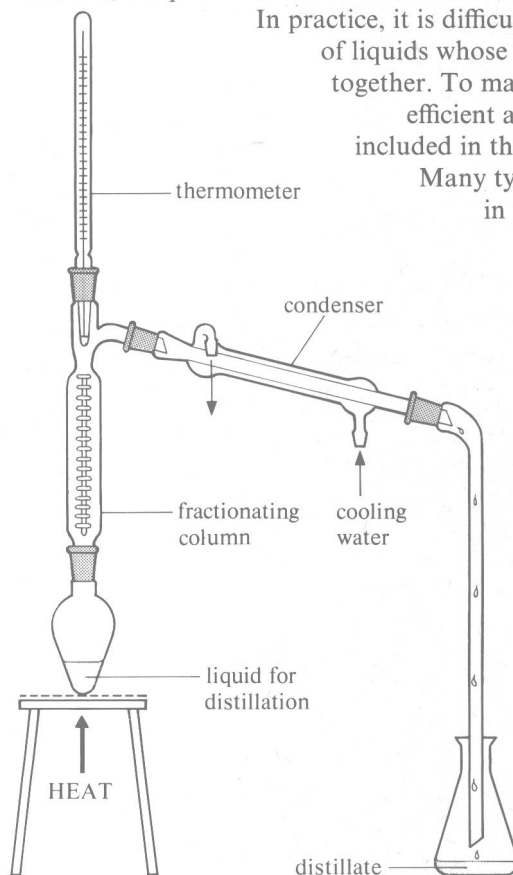


Fig. 1.14
A more efficient distillation apparatus using a fractionating column.

all consist basically of a long glass tube, usually packed with glass beads, pieces of glass tube or rod, etc. to provide large surfaces on which the vapours can temporarily condense. The hotter vapour passing up the tube then comes in contact with the liquid already condensed and evaporates it again. The column gradually warms up, but the lower end is always the hotter and the temperature falls gradually with increasing height. In effect, a whole series of “mini-distillations” is occurring in the column so bringing about a more efficient separation. The longer the column the more efficient is the separation and, in the separation of crude petroleum into the large number of miscible liquids present, columns may be thirty or more metres high (Fig. 1.15(a) and (b)).

1.8 SEPARATION OF IMMISCIBLE LIQUIDS

If two or more liquids which do not mix are placed in a test-tube they settle out in layers. Such a mixture can easily be separated by pouring it into a **separating funnel** (Fig. 1.16) and leaving it to stand. Eventually the lowest (and densest) layer can be run out using the tap provided and the other layers run off in turn.

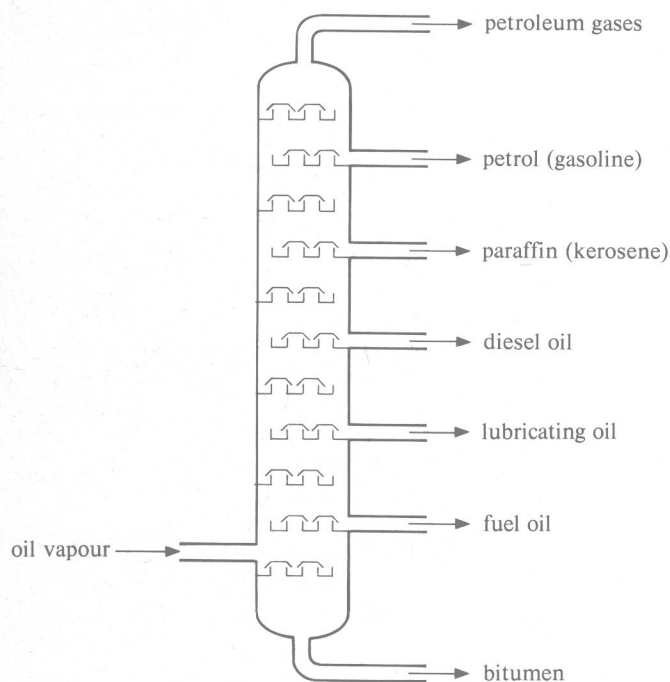


Fig. 1.15 (a) A simplified diagram showing the fractional distillation of petroleum. (b) (right) The distillation columns in an oil refinery may be thirty metres high.

1.9 CHROMATOGRAPHY

Chromatography is a technique originally developed to separate mixtures of coloured substances. A solution containing two coloured solutes (X and Y) is allowed to run down a column packed with a white absorbent powder such as chalk or aluminium oxide; the coloured substances adhere to the powder (Fig. 1.17(a)). When more solvent is poured in the top of the column the solutes gradually move down separating as individual bands (Fig. 1.17(b)).

Column chromatography is still used, but the technique you are most likely to use is known as paper chromatography. The simplest way to carry this out is to take a disc of filter paper and place a drop of the test solution in the centre. The paper is then carefully dried. A thin strip is then cut as shown in Fig. 1.18 to act as a “wick”. The paper is then laid on a beaker with the wick dipping in a suitable solvent. This may be water, ethanol, propanone or some more complex mixture. The ethanol and propanone are both flammable but are suitable for separating substances which do not dissolve in water.

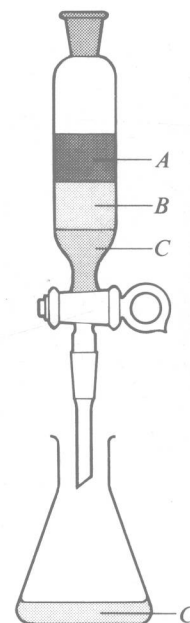


Fig. 1.16 (right) The separating of three immiscible liquids using a separating funnel.

The solvent gradually soaks up the wick until it reaches the main disc where it then spreads out towards the edge of the paper carrying the solutes with it. These are deposited in ring-shaped bands (Fig. 1.18). Fig. 1.19(a), (b) and (c) shows a better method of carrying out paper chromatography which is capable of separating more complex mixtures and also identifying the individual substances present. In the examples shown the mixture is made up of the three substances X, Y, and Z.

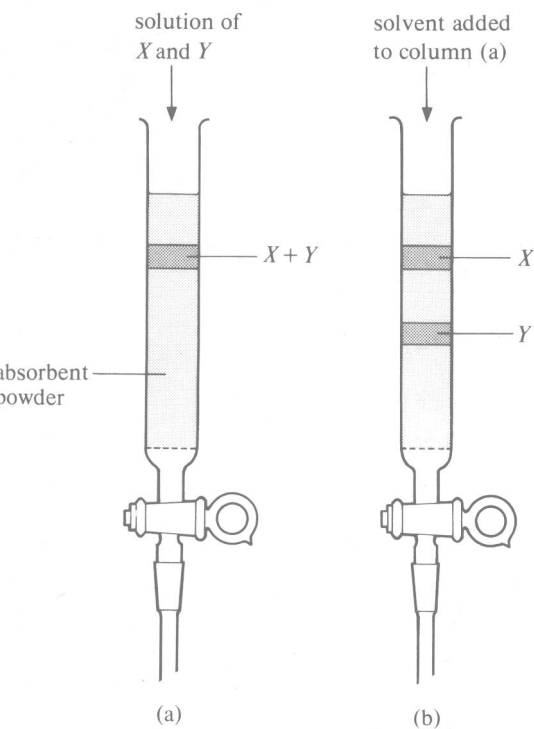


Fig. 1.17 Column chromatography. (a) A solution containing substances X and Y is added to the column. (b) More solvent is added and the layers of X and Y separate.

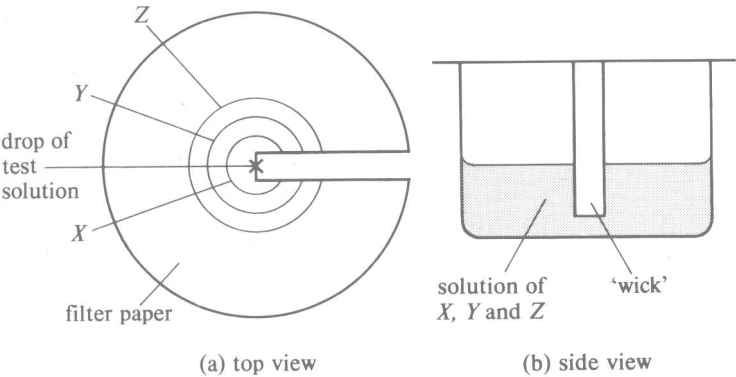


Fig. 1.18 A simple way of separating substances X, Y and Z by chromatography on a filter paper.

The technique can also be used for non-coloured substances, but in this case it is necessary to develop or locate the spots to make them visible. You will be studying this technique in greater detail later in the book (Section 11.5(b)).

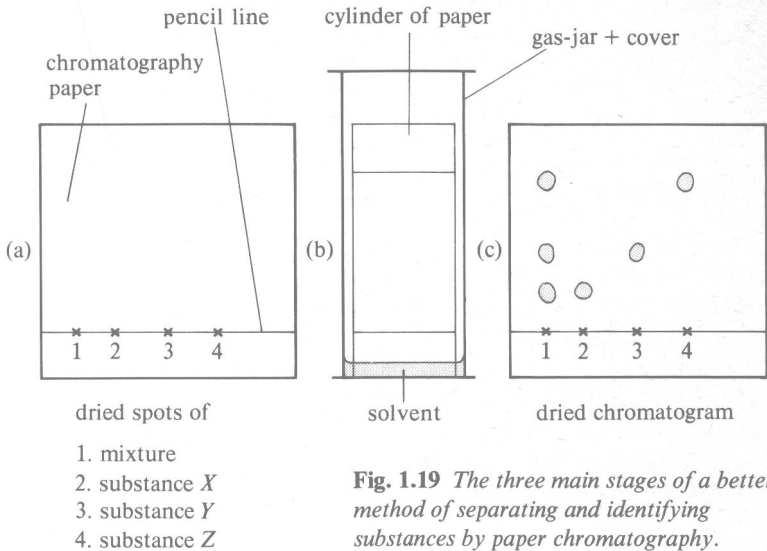


Fig. 1.19 The three main stages of a better method of separating and identifying substances by paper chromatography.

SUMMARY

The table below gives a summary of separating techniques

Type of mixture	Technique suitable
Suspension	Filtration <i>or</i> Centrifuging + decantation
Solution containing a single solute	Evaporation <i>or</i> Crystallisation (to obtain solute) <i>or</i> Distillation (to obtain solvent)
Solution containing two or more solutes	Fractional crystallisation <i>or</i> Chromatography (for small quantities) <i>or</i> Distillation (to obtain solvent)
Two or more miscible liquids	Fractional distillation
Two or more immiscible liquids	Use separating funnel

Words to remember

Buchner flask/funnel	fractionating column
bumping	immiscible
Bunsen burner	insoluble
centrifuge	luminous
centrifuging	miscible
chromatography	natural gas
condensation	pure substance
condenser	receiver
crystal	residue
crystallisation	saturated
decanting	separating funnel
distillation	solubility
evaporating basin	soluble
evaporation	solute
filter paper	solution
filtrate	solvent
filtration	suspension
flammable	teat-pipette
fractional crystallisation	town gas
fractional distillation	

Study questions

1. Although filtration is widely used in industry, filter paper is rarely used on a large scale. Suggest two reasons why this is so. What materials might be suitable for commercial filtration?
2. In Section 1.2 water was separated from sand and pebbles by filtration. Describe briefly another way in which this could be done. What advantages and disadvantages might your proposed method have?
3. Describe briefly three examples of the use of filters in everyday life. These need not necessarily be chemical applications.
4. Name three common substances (not necessarily pure) which are insoluble in water. What evidence can you produce for your choices?
5. Try to find out the name of one liquid used in dry-cleaning. Why is this liquid used in preference to water? If you were given a sample of this liquid (which has a dangerous vapour) after it had been used for cleaning, describe how you could purify it so that it could be reused.
6. Two common ways of drying washing are (a) hanging clothes on a line and (b) using a spin-dryer. Why is the spin-drier more reliable in its use than the clothes line? Which technique is more likely to produce the drier clothes? Explain your answers and state any assumptions you make.
7. Describe how you could obtain some dry, reasonably pure salt crystals from a sample of sandy sea-water?
8. The following table gives the solubilities of two substances, *M* and *N*, in the same solvent at different temperatures. Make use of this information to devise a method of obtaining some reasonably pure crystals of *M* from a mixture containing equal masses of *M* and *N*.

Substance	<i>M</i>	<i>N</i>
Solubility (g/100 g water at 25°C)	20	100
Solubility (g/100 g water at 100°C)	110	110

9. Describe how you could obtain samples of *D*, *E* and *F* from a bottle containing all three. *D* and *E* are immiscible liquids and *D* is denser than *E*. *F* is soluble in *D* but not in *E*.
10. Describe a method by which you could show that a specimen of blue ink contains at least two different dyes. How could the method be modified to distinguish the dyes in the ink from the dye in a felt-tipped pen; this ink was found to be unaffected by water.

2

The Air and Combustion

2.1 HEATING SUBSTANCES IN AIR

When samples of some metals are heated in air their masses increase. This indicates that the metal is combining with something else and this could come from:

1. the flame providing the heat
2. the container in which the metal is heated
3. the air surrounding the metal.

If the metal is heated in a test-tube it is most unlikely that the hot gases from the Bunsen flame would come in contact with the metal and so it appears that the first possibility can be ruled out.

If the metal is heated separately in glass test-tubes, porcelain crucibles and platinum crucibles, the result is always the same. It is improbable that these widely differing materials would all affect the metal in the same way and so the second suggestion is unlikely.

This leaves the third proposal which can be tested by heating pieces of metal, such as copper or zinc, in sealed test-tubes from which the air has been removed by a pump. In neither case does the metal change colour, whereas, when heated in air, the copper goes black and the zinc yellow. Their masses increase, but when heated in the absence of air, they do not. The possibility that these metals reacted with the air or some part of it is worth consideration.

2.2 COMPOSITION OF THE AIR

This proposal can be tested further by heating a piece of copper in a known volume of air and finding whether any of the air is used up by measuring the volume of gas left.

A convenient way of doing this is to use the apparatus in Fig. 2.1. Slowly, 100 cm³ of air is pushed between the gas syringes over

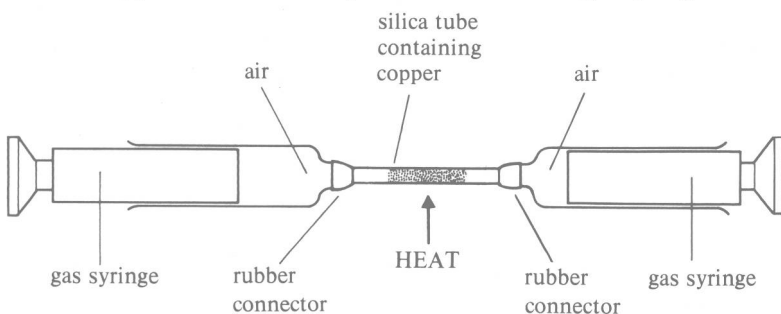


Fig. 2.1 Heating copper in a measured volume of air.

the hot copper. The process is repeated until no further change of volume occurs. On cooling, the volume of gas left is measured; a result of about 80 cm³ is usual, indicating that some of the air has been used up.

There are two possible reasons why no further change occurs. Either all of the copper has reacted, or all of that part of the air which combines with the copper has been used up. If the contents of the silica tube are examined, some of the copper still looks shiny and appears to be unchanged so that the first possibility is eliminated. It seems, therefore, that the part of the air which has been used up must be different from the remainder. This can be confirmed by forcing some of it gently over a burning splint, which goes out.

2.3 OXYGEN

In 1774, the French scientist Lavoisier had carried out a more elaborate version of this experiment in which he heated some mercury in air contained in the apparatus shown in Fig. 2.2. He found that the mercury became coated with a reddish powder which he called, "calx of mercury", and that the volume of air in the apparatus decreased. At this time that part of the air which reacted with the mercury was known as "active air" and the unchanged portion as "inactive air". At about the same time, the

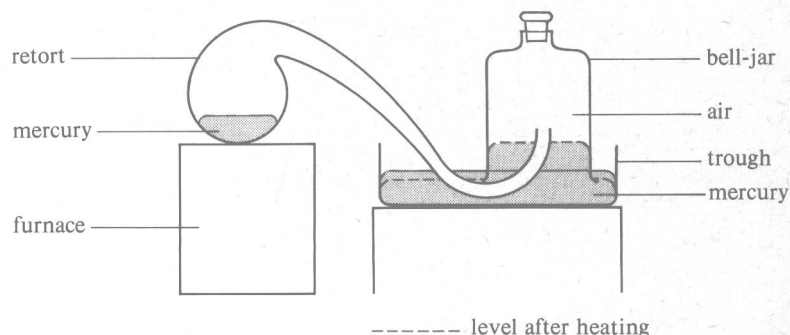


Fig. 2.2 Lavoisier's experiment investigating the composition of air.

British scientist Joseph Priestley prepared a gas similar to the “active air” by heating “calx of mercury” and **collecting the gas over water** – a method still used today and illustrated in Fig. 2.3(a) and (b).

The gas was named oxygen and experiments of the type described confirmed that approximately one-fifth of the air is oxygen, the remaining four-fifths being the “inactive air”. It was later shown that this “inactive air” is mainly nitrogen, but it also contains small quantities of other gases. Fig. 2.4 shows more accurately the composition of dry air.

2.4 THE LABORATORY PREPARATION OF OXYGEN

The substance originally called “calx of mercury” is now known as mercury oxide and it is perhaps not surprising that this and the oxides of some other metals produce oxygen gas when heated. This can be done by heating the oxide or other suitable substance and collecting the oxygen either over water (Fig. 2.3) or in a gas syringe (Fig. 2.5). A glowing splint placed in a jar of oxygen or in the stream of gas ejected from the syringe flares up violently. This is a simple test for oxygen but it does not work if the gas is mixed with large quantities of less reactive gases or combined with other elements.

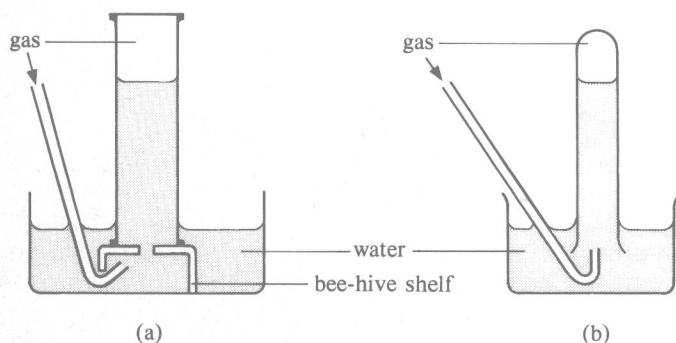


Fig. 2.3 Two ways of collecting a gas “over water”.

Composition of dry air (%)	
nitrogen	78.09
oxygen	20.95
argon	0.93
carbon dioxide	0.03
neon	0.0018
helium	0.0005
krypton	0.0001
xenon	0.00001

Fig. 2.4
The composition of dry air.

In practice, mercury oxide is rarely used to prepare oxygen since it is very expensive and mercury vapour is highly poisonous. Better alternatives are heating potassium manganate(VII) – which is not an oxide – or dropping a solution of hydrogen peroxide on to manganese oxide in which case no heat is required (Fig. 2.6).

2.5 THE COMMERCIAL PRODUCTION OF OXYGEN

These laboratory methods of preparing oxygen would all be far too expensive to use on a large scale and for commercial purposes oxygen is obtained by the **fractional distillation of liquid air**.

First the air has to be changed to a liquid which involves cooling it to about -200°C . Carbon dioxide and water are removed first since these would freeze solid and block the pipes. The remaining air is then compressed at 200 times atmospheric pressure. Compressing the air heats it (as in a vigorously used bicycle pump) and so it is restored to its original temperature by passing it through water-cooled pipes. The compressed air is then passed through a valve, where it is allowed to expand suddenly and this produces a considerable drop in temperature. The process is repeated until

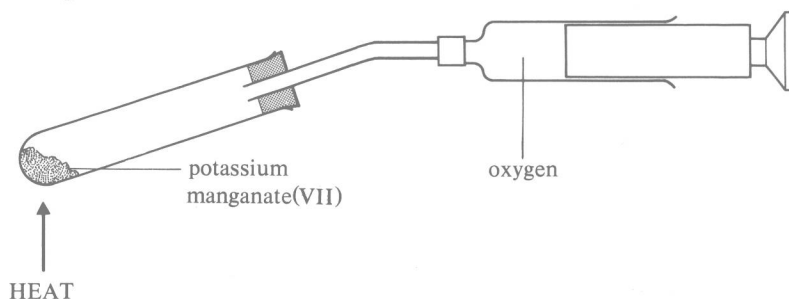


Fig. 2.5 Preparing and collecting oxygen by heating potassium manganate(VII).

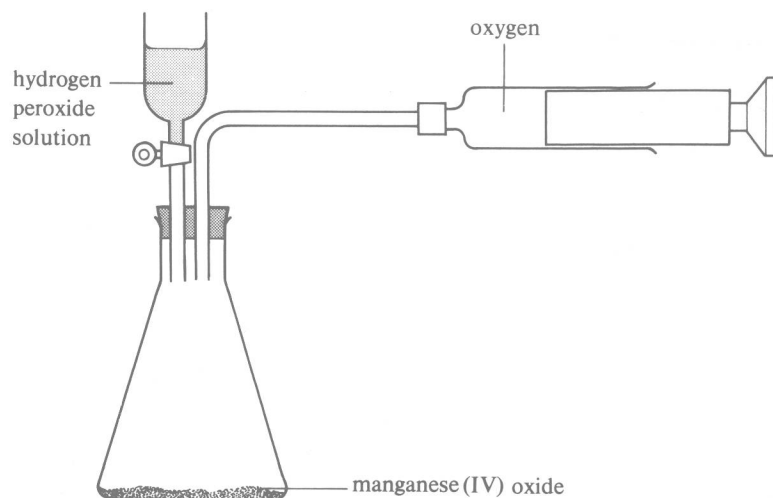


Fig. 2.6 The preparation of oxygen from hydrogen peroxide.

the air liquefies. The expanding air is used to drive an engine which in turn produces electricity. This can then work the pumps so reducing waste of valuable energy. The liquid air is passed to a fractionating column where the lower boiling nitrogen (b.p. -196°C) becomes a gas and the oxygen, which remains as a liquid, can be tapped off. The other substances in liquid air all boil at different temperatures and so it is possible to obtain other gases such as argon during the process (Fig. 2.7(a)).

As with other industrial processes described in this book, this is little more than an outline of principles, but when reading about methods of production you should always look for ways in which costs are kept as low as possible, e.g. by using the cheapest possible materials, making use of other substances produced – known as by-products – and by reducing waste of energy to a minimum.

The main uses of oxygen are in the production of steel (Section 17.2), in oxyacetylene burners, for welding and cutting metals, and for breathing aids in hospitals, high-flying aircraft and for divers. From these main uses it is clear that oxygen is used because substances burn vigorously in it and it is essential for respiration in animals and plants.

Most of the argon formed is used in filling electric light bulbs and the nitrogen for producing ammonia and other important chemicals (Sections 16.5 and 16.6).

2.6 COMBUSTION

A chemical reaction in which a substance combines with oxygen is an example of **oxidation** and the original substance is said to be oxidised. Later in the book (Sections 6.12 and 19.5) the concept of oxidation will be discussed in detail. An oxidation reaction in which enough heat is produced to keep the reaction going without further energy being supplied from outside is called a **combustion**. It can be seen that a substance needs another substance with

which to react if combustion is to occur; the other substance is said to **support combustion** of the first one. Oxygen is an excellent supporter of combustion, but it is by no means the only one.

SUMMARY

1. Substances heated in air combine with approximately 21% of it to produce oxides. This part of the air is oxygen, the remainder being a mixture largely of nitrogen together with a little carbon dioxide and the noble gases, particularly argon.
2. Oxygen is a reactive gas commercially produced by the fractional distillation of liquid air. When a substance combines with oxygen it is said to be oxidised; an oxidation reaction in which sufficient heat is produced to maintain it is called a combustion and the substances produced are known as the products of combustion.

2.7 THE LABORATORY PREPARATION OF HYDROGEN

Another gas which was discovered about the same time as oxygen is hydrogen. If a flame is applied to the open end of a test-tube containing this gas there is usually a small, but noisy, high-pitched explosion. If a lighted taper is plunged upwards into

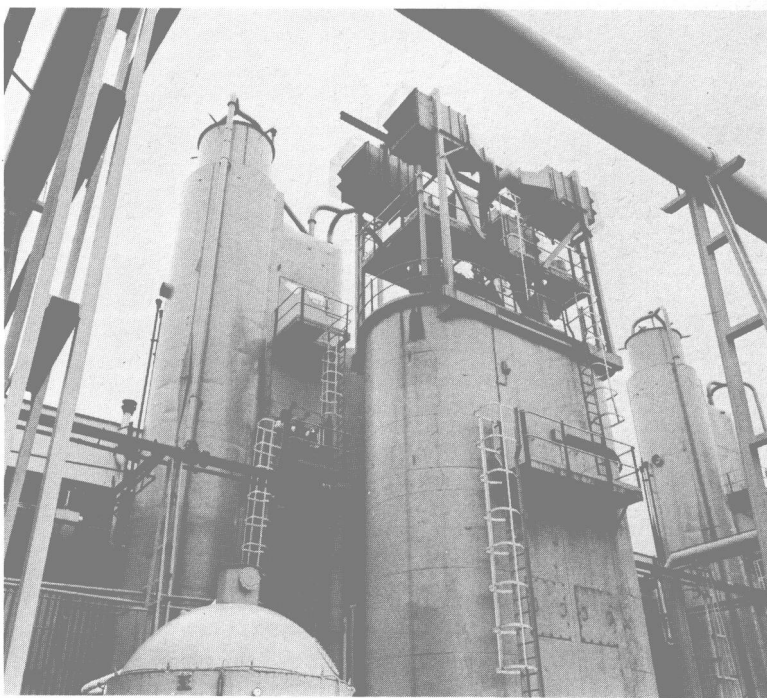
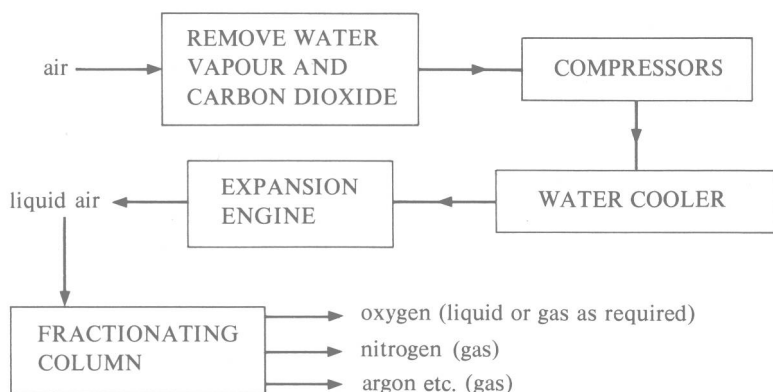


Fig. 2.7 The production of oxygen from the air on a large scale. (a) The main stages in the process. (b) (right) A general view of part of a British Oxygen Company plant.