

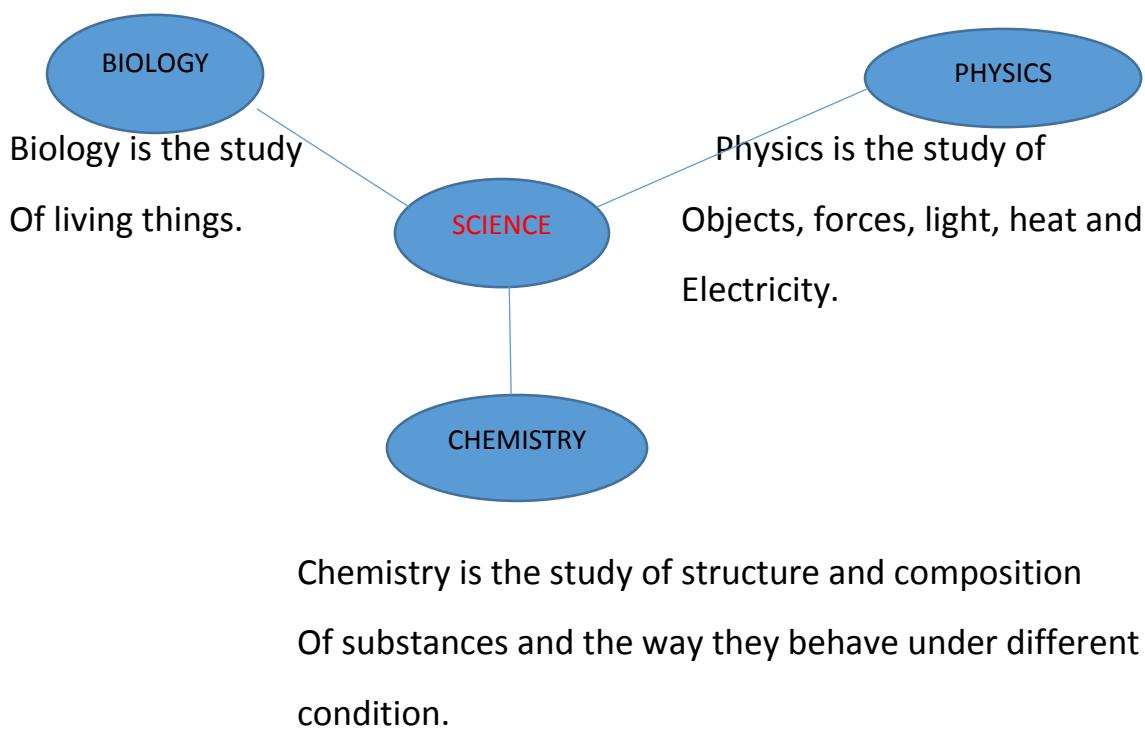
## UNIT 1

## INTRODUCTION TO CHEMISTRY

### Meaning of Chemistry

Science is the systematic study of things in nature, the changes that they are involved in and the reasons for these changes. Our world is full of wonders of science and technology. As you continue to study science, you will learn more about these. You might even become a great scientist yourself and come up with your own interesting and useful discoveries.

At primary school level, science is learnt as a general subject. At secondary level, the study of science is split into three main subjects. These are Physics, Biology and Chemistry (Fig. 1.1). The application from these subjects is studied in related subjects such as Agriculture, Home Science and others.



*Fig. 1.1 Branches of science*

You have already learnt some aspects of Chemistry in primary school Science. In secondary school, we will build upon this knowledge in the various topics that are in this book and other Chemistry books.

## Branches of Chemistry

Chemistry is further sub-divided into six branches as shown in Fig. 1.2

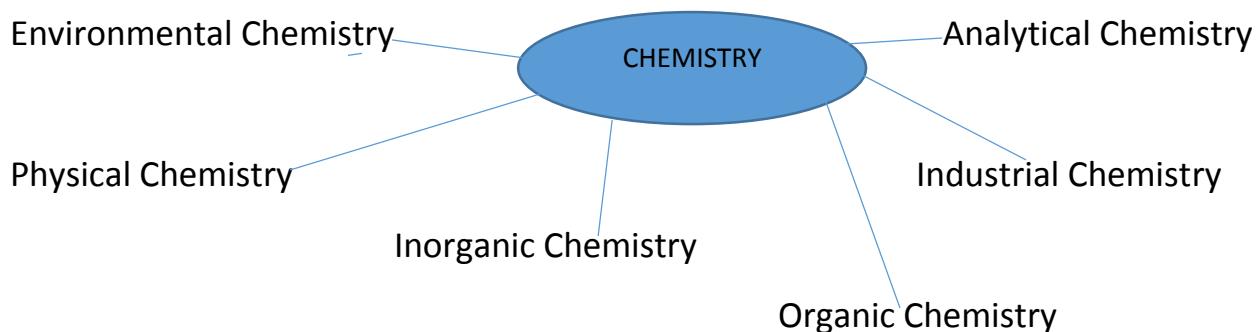


Fig. 1.2 Branches of Chemistry

Table 1.1 gives the description of each of the branches.

Table 1.1 Description of branches of Chemistry

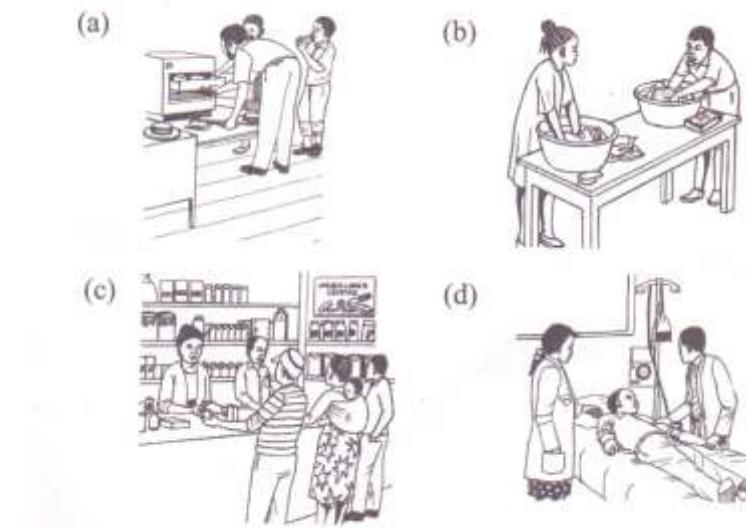
Branch of Chemistry	Description	Example
Analytical Chemistry	The study of the separation, identification and quantification of the chemical components of natural and artificial materials	Food processing industry
Industrial Chemistry	The study of the application of physical and chemical processes towards the change of raw material into beneficial products.	Manufacture of painkillers like Panadol.
Organic Chemistry	The study of compounds of carbon apart from the oxides and carbonates.	Distillation of crude oil.
Inorganic Chemistry	The study of properties of materials from non-biological origins	Fertilizers' composition

Physical Chemistry	The study of how chemical compounds and their constituents react and interact with each other	Separation of coloured substance.
Environmental Chemistry	The study of chemical process occurring in the environment which are impacted by human activities.	Pollution of air, soil and water.

In secondary school we will learn the elements of each one of these Chemistry branches at different depths and coverage in topics of this course. It is important to know that each one of these branches is useful in everyday life.

### Importance of Chemistry in everyday life

Chemistry is very important in everyday life. Fig 1.3 shows some applications of Chemistry in everyday life.



*Fig. 1.3 Application of Chemistry*

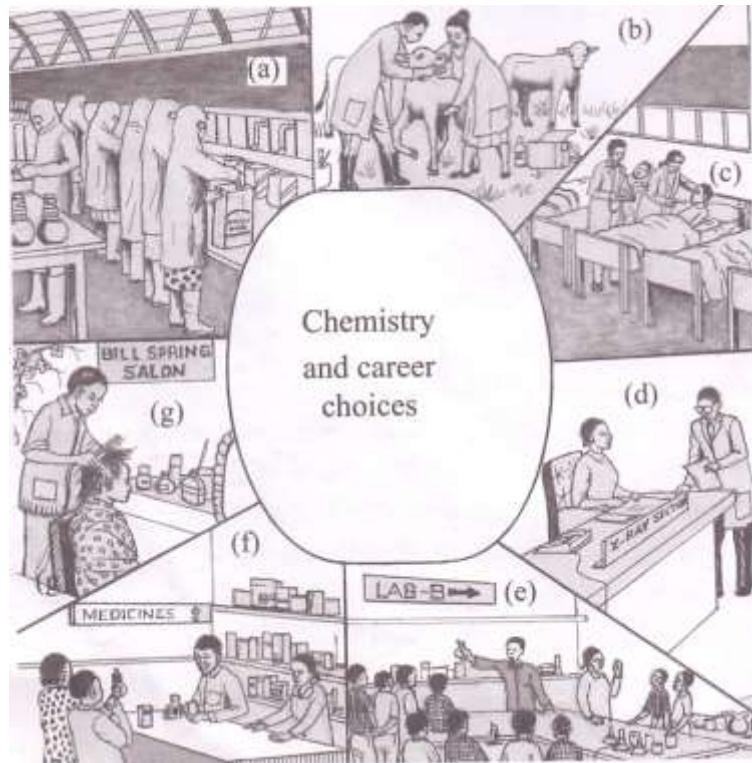
- a. When baking powder is added to doughnut or bread and the mixtures baked in a hot oven, the rising of the bread can be explained in Chemistry.
- b. In the manufacture of soap and other detergents Chemistry is applied. When you wash clothes using soap, the cleaning action is explained in Chemistry.
- c. Chemistry is applied in manufacture of drugs such as Panadol. Such drugs are used in homes and hospitals.
- d. Chemistry is applied in cooking especially in mixing of ingredients while making a cup of tea and preparation of nsima.

Other application of Chemistry include:

- The industrial manufacture of plastics, glass, cement, fabrics like nylon and polyester, soaps, insecticides, fertilizer, industrial food processing and others require sound knowledge of Chemistry.
- The processing of crude oil.
- Discovery, testing and release of new drugs into the market.

### **Chemistry and career choices**

We have already seen that Chemistry has a lot to offer in our day-to-day lives. There are many careers in Chemistry to pursue after leaving school, college or university. Some of these careers are shown in Fig 1.4a, b, c, d, e, f, and g.



*Fig 1.4 Chemistry and career choices*

- a) Chemical engineers
- b) Veterinary officers
- c) Medical nurses
- d) Medical doctors
- e) Chemistry teachers
- f) Pharmacists
- g) Beauticians

Other career include geochemist, biochemists, food chemist, quality control personnel, environmental chemist, photochemist, assurance personnel among others.

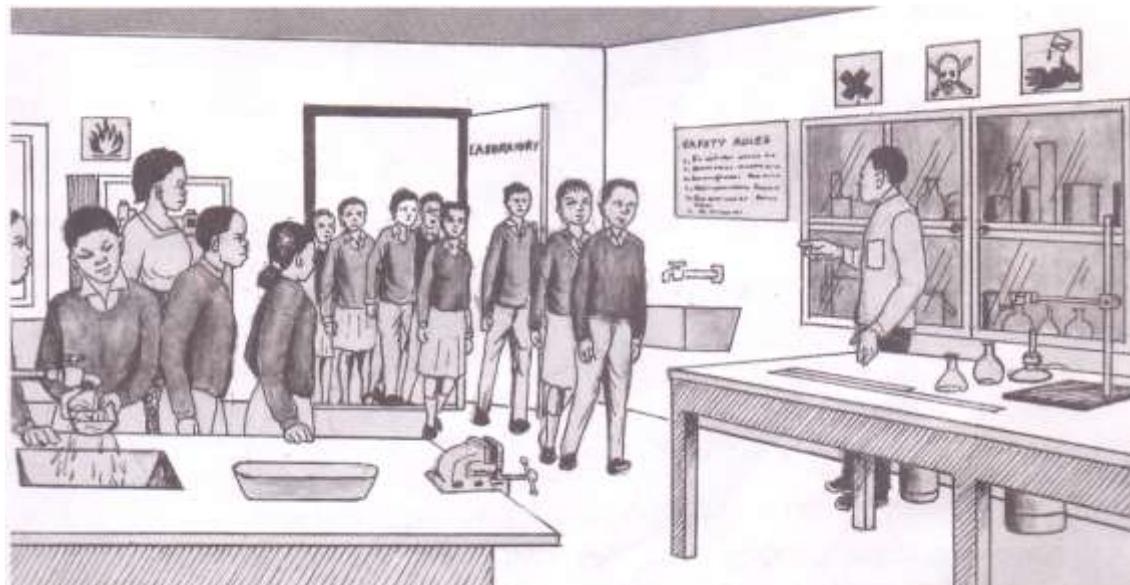
All careers in Chemistry contribute to the welfare of the society besides offering employment to individuals. What is the importance of each of the careers in fig. 1.4 to our society?

### **Chemistry laboratory and laboratory safety rules**

Chemistry experiments are done in order to learn about changes in substances. The experiments can be carried out in special rooms, known as laboratories.

However, experiments in Chemistry can also be carried out in other rooms or spaces not necessarily designed as being special. These include your classrooms and open spaces. Three elements; space, apparatus and chemical must be present in a laboratory.

A laboratory is costly to build and maintain. In addition, chemicals and apparatus are very expensive. Some of them are even harmful or poisonous to the users and the environment. Therefore, careful storage and usage of the laboratory, apparatus and chemicals are needed. Symbols or warning signs have been developed to indicate the dangers associated with the laboratory, apparatus and chemical. Study Fig. 1.5 which shows some of these symbols.



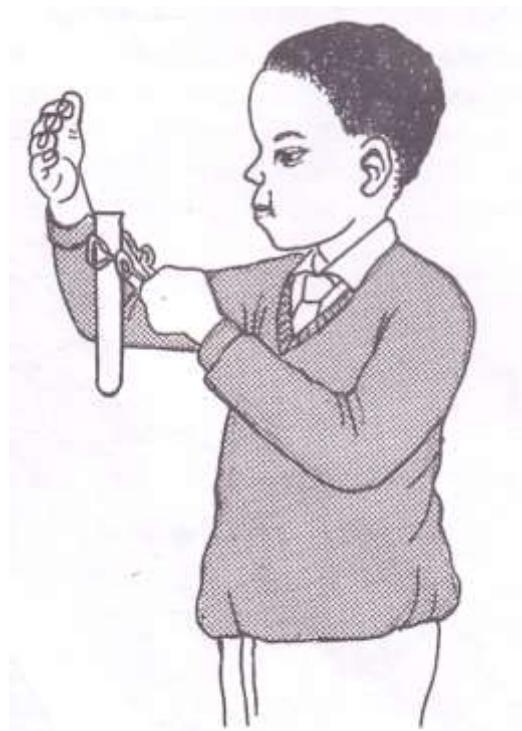
*Fig. 1.5 Students entering into the laboratory in an orderly manner*

### Laboratory safety rules

We have already pointed out that there are dangerous and poisonous chemicals stored in the laboratory. Also, there are glassware, cutting instruments and even explosives materials like gas cylinders in the laboratory. Therefore, observing safety while in the laboratory is very important. There are a number of laboratory safety rules to observe in order to stay safe while in the laboratory. They include:

- Do not enter the laboratory without the teacher's permission.
- Enter the laboratory in an orderly manner
- Do not rush as you enter the laboratory or scramble for the front bench.

- Avoid unnecessary movement in the laboratory
- Do not do unauthorized experiments.
- Do not taste any of the chemicals, even those that you think are harmless.
- Do not temper with electric, gas or water fittings. Always turn off water taps and gas taps when not in use.
- Do not smell gases directly. Hold the gas source about 15-20 cm from the nose and waft the gas towards the nose with your palm, then sniff carefully, Fig 1.6



*Fig. 1.6 Wafting a gas towards the nose*

- never interfere with other student's experiments left in progress in the laboratory
- Never boil a liquid while facing the test tube towards yourself, friends or books in case of spouting.
- Keep naked flames away from inflammable substances, especially volatile liquids or solutions.
- If chemicals get into contact with your clothes, skin, eyes or mouth, wash the affected area with plenty of water immediately.

- In case of a burn or a cut, report the accident to the teacher or laboratory technician immediately.
- Read the label on all reagents before you use them.
- Do not put hot materials into the waste containers.
- All solid waste should be properly disposed of. Do not pour insoluble materials into the sinks.
- Read and follow instructions carefully before you perform any experiment.
- All experiments that emit poisonous gases should be performed in a fume chamber, or open space.
- After finishing with a reagent bottle, replace the stopper and store it in its proper place.
- Always clean all the apparatus and bench tops after use. Wash your hands well before leaving the laboratory.
- Do not move anything out of the laboratory unless authorized.

**Caution:** we often use glassware in experiment in Chemistry. Broken glassware can cause deep cuts. It is very important to prevent accidents in the laboratory.

When performing experiments you must protect yourself by:

- Wearing a laboratory coat to protect your clothes from dirt and chemicals.
- Wearing plastic goggles to protect your eyes from solid particles or splashing liquids.
- Wearing plastic gloves to protect your hands from corrosive chemicals.
- Holding hot objects with holders for example tongs or improvised holders like folded paper.
- Tying long hair at the back.
- Preparing irritating gases in the fume chamber or open space.

### **Common laboratory apparatus**

Apparatus refers to the set of equipment used by chemists to do their work. Apparatus may be made of metal, wood, plastic or glass. Most apparatus used in chemistry experiments are made of glass. Can you suggest reasons for this?

It is important to know the apparatus to choose and use when performing a particular experiment. You also need to be aware of the purpose and accuracy of the common laboratory experiment.

Laboratory apparatus can be put into various categories as follows:

- Measuring apparatus
- Heating apparatus
- Other apparatus

### a) Measuring apparatus

In Chemistry, different apparatus are used to measure volume, mass, temperature and time.

#### 1. Apparatus for measuring volume

Examples include:

##### I. Measuring Cylinders

They are used for measuring approximate volumes of liquids or solutions. They have different capacities. Measuring cylinders may be made of glass or plastic.

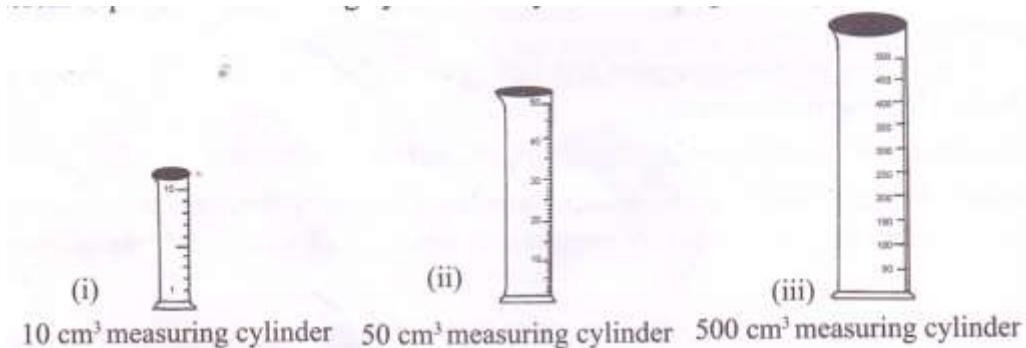


Fig. 1.7 Measuring cylinders of different capacities

##### II. Burette

A burette is used for measuring accurate or exact small volumes of liquids or solutions during chemical analysis.

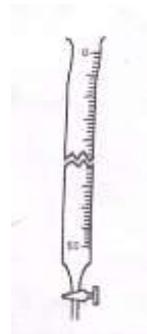


Fig. 1.8 Burette

### III. Pipette

A pipette is used for transferring or measuring small exact quantities of liquids or solutions during chemical analysis.

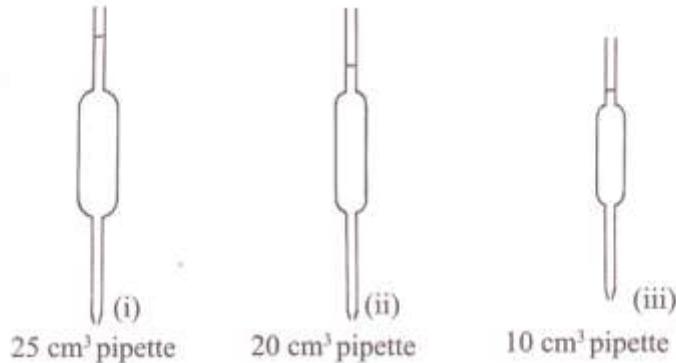


Fig. 1.9 Pipette of varying capacities

### IV. Volumetric flask

A volumetric flask is used to prepare accurate volumes of liquids or solutions. They have different capacities.

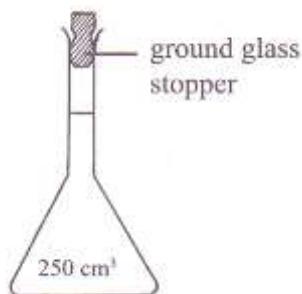
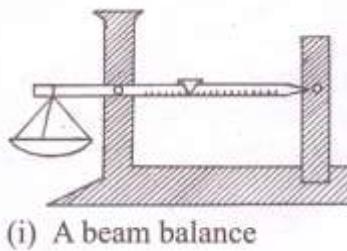


Fig. 1.10 250 cm<sup>3</sup> Volumetric flask

## **2. Apparatus for measuring mass**

A beam balance or electric balance is used to measure accurate masses of substances.



(i) A beam balance



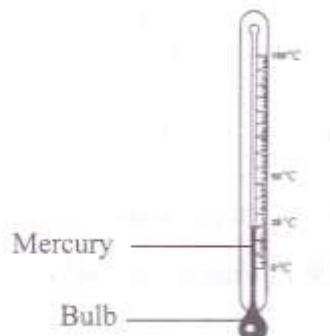
(ii) Electronic balance

*Fig. 1.11 Balance*

Balances measure mass of substances in grams (g) or kilograms (kg)

## **3. Apparatus for measuring temperature**

A thermometer is used to measure temperature. Examples of thermometers are mercury and clinical thermometer.



*Fig. 1.12 Mercury thermometer*

## **4. Apparatus for measuring time**

A stop watch or clock is used to measure time.



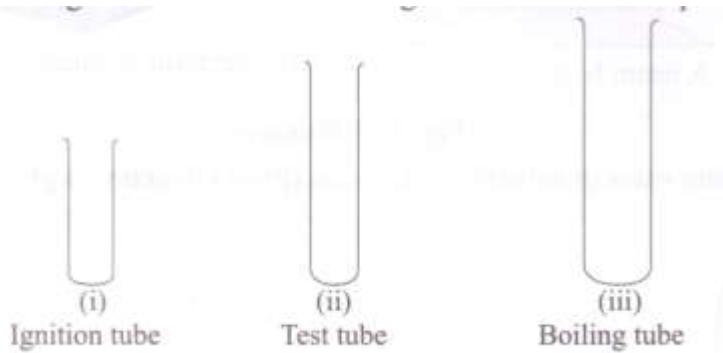
*Fig. 1.13 Stop watch*

**b) Other apparatus**

These are mainly made of glass and are also known as **glassware**. Examples include

**i. Tubes**

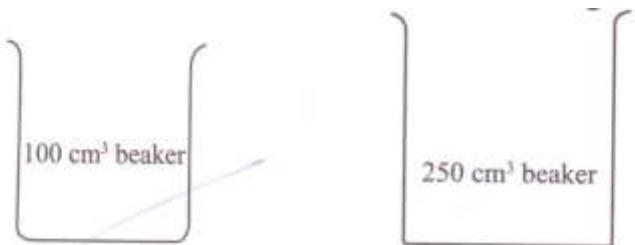
There are different kinds of tubes depending on their use. Ignition tubes are used for decomposition of compounds by strong heating, ordinary test tubes are for heating substances and boiling tubes are used for boiling small amounts of liquids or solutions.



*Fig. 1.14 different types of tubes*

**ii. Beakers**

They are lipped glass or plastic vessels of various capacities. Beakers have many uses such as boiling liquids or solutions (if they are made of glass) and holding the liquid or solution.



*Fig. 1.15 Beakers of different capacities*

### iii. Flasks

There are various types of flasks in fig. 1.16. They are usually used to hold liquids or solutions when heating and in reactions where heat may be liberated.



*Fig. 1.16 Different types of flasks*

### iv. Gas jars

Gas jars are used to collect gases that are prepared in the laboratory. They are usually covered with glass covers to prevent a light gas from escaping.



*Fig 1.17 A gas jar*

#### v. Evaporating dish

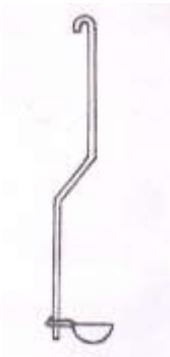
Evaporating dishes are used to evaporate water from solutions in order to recover the dissolved solid substance (solute). They are usually made of porcelain.



*Fig. 1.18 Evaporating dishes of different shapes.*

#### vi. Deflagrating spoon

A deflagrating spoon is used to hold substances when burning them in air or in gas jars.



*Fig. 1.19 Deflagrating spoon*

#### vii. Spatula

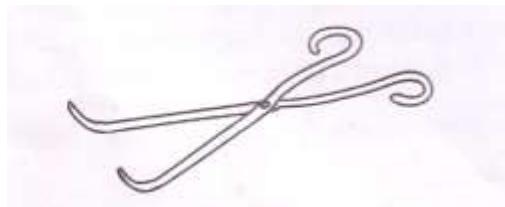
A spatula is used to scoop small quantities of solid chemicals from a container



*Fig. 1.20 Spatula*

#### viii. Tongs

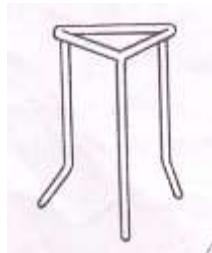
Tongs are used to hold hot crucible or crucible covers.



*Fig. 1.21 Tongs*

**ix. Tripod stand**

A tripod stand is used to support apparatus like beakers when boiling or heating a liquid.



*Fig. 1.22 Tripod stand*

**x. Wire gauze**

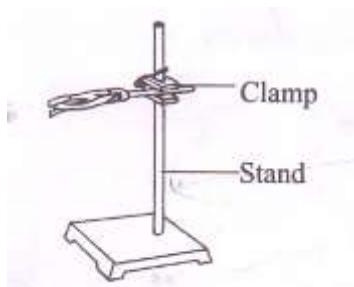
It is usually placed on the tripod stand to hold beakers when boiling liquids.



*Fig. 1.23 Wire gauze*

**xi. Stand and clamp**

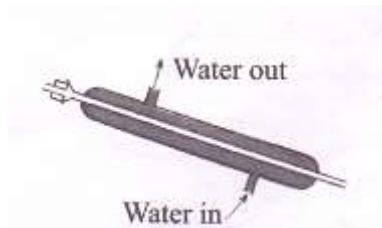
A stand is used to mount apparatus, whereas a clamp is used to hold burettes and other apparatus when doing an experiment.



*Fig. 1.24 Complete stand and clamp*

### xii. Liebig condenser

Cools the vapour as it passes the centre of the condenser during distillation when separating miscible liquids.



*Fig. 1.25 Liebig condenser*

### xiii. Funnels

There are different types of funnels:

#### a. Thistle funnel

Used to add liquids into flasks at once:



*Fig 1.26 Thistle funnel*

#### b. Dropping funnel

Used to add liquids or solutions into flasks in small quantities.

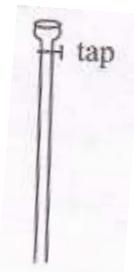


Fig. 1.27 Dropping funnel

**c. Separating funnel**

Used to separate immiscible liquids.

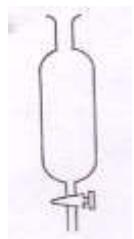


Fig 1.28 Separating funnel

**d. Filter funnel**

Used to pour liquids into containers with small mouths and also filter solutions containing undissolved solids. A filter paper is placed inside the funnel to hold undissolved material.



Fig. 1.29 Filter funnel

**e. Dropper or Teat pipette**

Used to deliver a little liquid when required in drops in to another container.

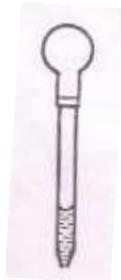


Fig. 1.30 Dropper or Teat pipette

### c) Heating apparatus

There are various heating apparatus that can be used in the laboratory. Examples include: spirit lamp, candle, gas or kerosene stove, electric heater and the Bunsen burner.

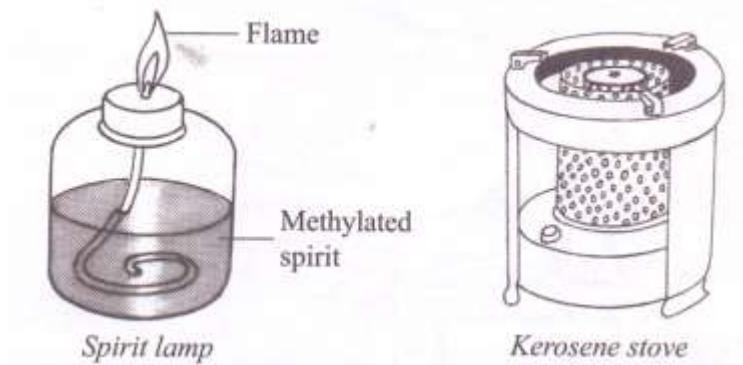
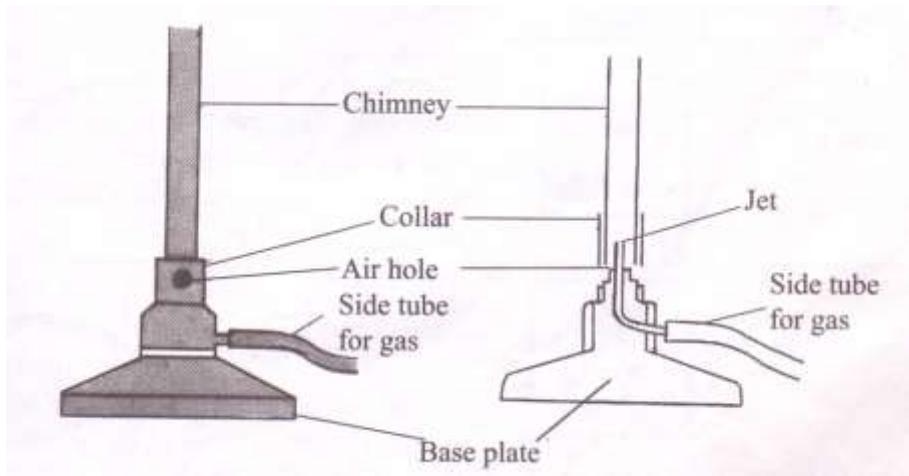


Fig. 1.31 some heating apparatus

The most important and widely used heating apparatus in the laboratory is the Bunsen Burner.

### The Bunsen Burner

The Bunsen burner was invented in 1854 by a German chemist called Robert Wilhelm Bunsen from where it derives its name.



*Fig. 1.32 parts of Bunsen burner*

### Hazard symbols and their meanings

There are hazard symbols associated with the dangers in handling laboratory chemicals and equipment. Refer to appendix I for some of these symbols.

Hazard	Symbols	Meaning
<b>Explosive</b>		Explosive substances can detonate any time. Chemicals bearing the explosive sign should be handled carefully and according to instructions.
<b>Toxic</b>		Toxic substances are poisonous. They can cause death. So if any chemical pours on your skin accidentally wash it with a lot of water. Watch out for toxic substance sign
<b>Oxidizing</b>		There are some substances which can explode in the presence of an oxidizing agent

<b>Corrosive</b>		There are some substances which can burn your skin. They corrode bench tops, metals and floors. Pour water on any chemical that you may pour on the benches accidentally to make it less corrosive
<b>Flammable</b>		There are some chemicals that evaporate easily and catch fire very easily. They should never be brought near open flames. Once they are removed from the bottle the bottle must be stoppered immediately
<b>Irritant</b>		Irritant or harmful substances make you sick or endanger your health. They can endanger your health if they come in contact with your skin or eyes for too long. People who are asthmatic should be extremely careful to avoid inhaling such substances.

### Common units of measurement and their symbols

You are familiar with some physical quantities such as: length, mass, time, density, temperature and volume. Different books depending on when the book was written have expressed the quantities in different units. For instance, old books expressed length in yards, instead of metre, temperature in degrees Fahrenheit instead of Celsius or Kelvin, volume in cubic feet instead of cubing metre among others. This is no longer the case.

### The SI Unit

In modern scientific literature an international system of units has been adopted. It is known as the SI system of units (**SYSTÈME INTERNATIONAL D'UNITES**)

It is accepted worldwide and the units are used to express quantities. Each quantity has an SI unit and each unit has a symbol. For example, the quantity mass has the unit of either Kilogram symbol (kg) or with gram with symbol (g).

### **Representation of physical quantities**

Physical quantities are represented in symbols and units. Physical quantities are meaningless if they have no units.

### **Basic and derived units of measurement**

In Chemistry, there are basic acceptable International system of units (SI) used for measurements. These units are increasingly used in schools, colleges and universities. The four basics SI units as shown in table 1.2

**Table 1.2 Basic SI units of various quantities**

<b>Physical quantity (measurement)</b>	<b>Name of SI unit</b>	<b>Symbol of SI unit</b>
Length	Metre	m
Mass	Kilogram	Kg
Time	Second	s
Temperature	Kelvin	K

The SI rules demand that the symbol should not be followed by full stop nor should they be plural of the mass of substance. For example, ten kilograms is abbreviated as 10 kg not 10 kg. or 10 kgs.

There are however, units that are not SI and are used in books for example, centimeters (cm) or millimeters (mm) for length, grams (g) for mass, degrees centigrade ( $^{\circ}\text{C}$ ) for temperature and cubic centimeters ( $\text{cm}^3$ ) or litres for volume among others.

### **Derived units**

Other physical quantities obtained by combining (dividing or multiply) one or more of the basic physical quantities are called derived physical quantities. They derived physical quantities have their derived units as shown in table 1.3

**Table 1.3 Some SI derived units expressed in terms of base units**

Derived quantity	Derived SI unit name	Symbol
Area	square metre	$m^2$
Density	kilogram per cubic metre	$kg/m^3$
Amount of substance (concentration)	moles per cubic metre	$mol/m^3$

### SI prefixes

Sometimes working with SI units gives challenges because you may be dealing with very large numbers or very small values. For example, one mole of a substance contains approximately  $6.0 \times 10^{23}$  particles of a substance. Imagine counting  $6.0 \times 10^{23}$  particles. How many days would it take? Others for example length of a bond between atoms in a molecule would be as small as  $10^{-10}$  meter. In order to accommodate such extreme values, multiple of units in powers of ten are indicated by means of agreed prefixes. Table 1.4 gives some of the prefixes and their sizes as multiples of ten.

**Table 1.4 Some prefixes and their sizes as multiples of ten**

Multiple	Prefix
$10^{-6}$	Micro
$10^{-9}$	Nano
$10^6$	Mega
$10^9$	giga

See other multiple of units as shown in Appendix II

### Measuring physical quantities

Measurements are made in our everyday lives; in school laboratories, shops, farms and many other places. The measurements taken include mass and volume which are taken to find out the quantity of matter present in a substance.

Knowledge of quantities of measurements is very important in sciences and other subjects. In industrial preparation of substances, definite quantities of reactants have to be measured to produce a desired quality and quantity of a substance. This helps us to avoid use of excess substances, to know the cost of substances used and finally know the quantity to be produced.

Soap is a substance which is used every day. It is manufactured using vegetable oil for example palm oil and another substance called potassium hydroxide.

Measurement of volume of these reactants to get quality soap is necessary.

Measurement of other quantities accurately for example temperature is also important.

### **Scientific methods of investigation**

We have stated what Chemistry is, its role in society and some careers that one can pursue after successfully studying Chemistry. But we must now try to answer the question “How do we study Chemistry?” people who studied Chemistry in the 18<sup>th</sup> century had very few apparatus to use.

They were kind of speculative philosophers whose main aims were:

- a. Transmutation (change natural) of base metals like copper, iron and silver to gold.
- b. To discover universal cure of diseases.
- c. To discover means of prolonging life indefinitely.

It is important to note that early experiments were done in a systematic manner. Today, with modern equipment and technology we study Chemistry with organized modern scientific methods. These modern methods are tested and verified in the laboratory. The knowledge gained during testing or experimentation in the laboratory is used later in scientific research, for example in discovery of new drugs and medicines. This forms the basis of.

The following steps should be carried out when carrying out an investigation:

**1. State the question.**

What is it that you want to find out?

The question may refer to the explanation of a situation or a specific observation you made. The problem you want to investigate may be for example, why do we sleep? Why is the sky blue? Or How can I design a drug for a particular disease for example cancer or HIV?

**2. Research on the topic.**

Read to find out what chemists have investigated and learned about the topic.

**3. State hypothesis**

Predict using available information, what you think will happen on testing or in the experiment. This guess must be stated and proved right or wrong by experimentation.

**4. Testing the hypothesis**

Design an experiment or method for testing if the hypothesis is true or false.

**5. Analyze the results**

This involves examining the data collected and determining what the results of the experiment show. Do the results support the hypothesis?

**6. Draw a conclusion**

If the results of the experiments come out as predicted, the hypothesis will be accepted as being correct. However, if the results do not come out as predicted, the hypothesis will be rejected as false. This means that another hypothesis is required. A new question can be asked on the same topic and the investigation repeated.

**Formulation of theories and laws**

Using scientific inquiry, theories and laws can be formulated, through experiments. If the results of an experiment are corrected and consistent they can be used to come up with a law. For example,

From the results, you notice that mass divided by volume gives a constant.

The two quantities; mass and volume are related.

The relationship between mass and volume is called density. This relationship can be used to formulate a law. For example:

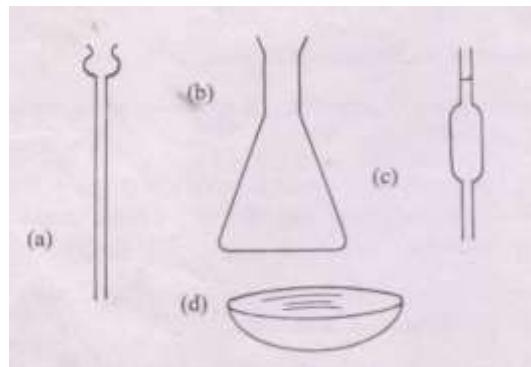
$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

**Revision Exercise 1**

1. Why are most of the apparatus in Chemistry laboratories made of glass?
2. Name two apparatus used for accurate measurement of volume

3. Matter occupies ..... and has .....

4. Name the following apparatus



5. Write down the names of four common units of measurements and their symbols.

b). What physical quantities do the following SI symbols represent?

- 1) s
- 2) k
- 3) kg
- 4) mol

6. Differentiate between basic and derived units of measurements.

7. Name four instruments used to measure physical quantities

8. (a) What is Chemistry

(b) Name two branches of Chemistry

(c) Write down three application of Chemistry

9. Name four careers one can pursue after studying Chemistry.

10. (a) State four laboratory rules.

(b) Why must students enter the laboratory in an orderly manner?

Standard form (also known as scientific notation) is the way that scientists write numbers that are too big or too small. For example:

In Chemistry we often have to deal with either very large numbers, such as Avogadro's number which is, 602,300,000,000,000,000,000,000, the speed of light 30,000,000,000, centimeters per second or very small numbers, such as the mass of an electron 0.000,000,000,000,000,000,000,911g.

Such very large and very small numbers can be conveniently written as a product of two numbers.

Thus scientific notation is a convenient method of expressing these large and small numbers in exponential form (that is, in powers of ten).

For example;

- a. The Avogadro's constant 602, 300,000,000,000,000,000,000, is expressed in scientific notation as  $6.023 \times 10^{23}$ . The notation  $6.023 \times 10^{23}$  tells us to multiply 6.023 by 10 twenty-three times.  
The first is called the digit term which is a number greater than one but less than ten.  
The second number is called the exponential term and is written as 10 with an exponent (10 raised to a power n where n is a whole number, that is,  $10^n$ ).  
b. The speed of light 30, 000,000,000 cm per sec is expressed in scientific notation as  $3.0 \times 10^{23}$ . The notation  $3.0 \times 10^{23}$  tells us to multiply 3.0 by 10 times.  
c. The mass of an electron 0.000, 000,000,000,000,000,000,911g, is expressed in scientific notation as  $9.11 \times 10^{-28}$  g. the notation  $9.11 \times 10^{-28}$  tell us to divide 9.11 by 10 twenty-eight times.  
d. The co-efficient of expansion of copper 0.0000167 is expressed in scientific notation as  $1.67 \times 10^{-5}$ . The notation  $1.67 \times 10^{-5}$  tells us to divide 1.67 by 10 five times.

Multiply a digit term by a positive power of 10 moves the decimal point that number of places to the right.

Examples of scientific notation with digit terms multiplied by positive powers of 10 written as numbers:

$$1.23 \times 10^4 = 12300$$

$$9.46 \times 10^7 = 94,600,000$$

$$7.29 \times 10^2 = 729$$

Multiplying by  $10^{-n}$  where n is negative, move the decimal point (•) the number of places indicated by the value of n to the left.

Example of scientific notations when the value of n is negative number:

$$2.56 \times 10^{-1} \text{ or } 2.56 \times \frac{1}{10} = 0.256$$

$$6.42 \times 10^{-2} \text{ or } 6.42 \times \frac{1}{10^2} = 0.0642$$

$$5.745 \times 10^{-5} \text{ or } 5.745 \times \frac{1}{10^5} = 0.00005745$$

### Significant figures

A number can have one or more digits. For example 20 has two digits that is 2 and 0; 231 has three digits that is 2, 3 and 1.

A significant figure is a digit in a number that shows how the number is nearer to its exact value. Figure 2.1 is an example used to illustrate a significant figure.

The ruler in figure 2.1 is graduated in centimeters, it is used to determine the length of the graphite rod besides it to three significant figures. One may not agree on the exact value of the third digit which reads as 1.35 cm or 1.36 cm. in other words, the third digit is uncertain but important in that it shows that the graphite rod is slightly longer than 1.3 cm. the two certain digit are known as significant figures.

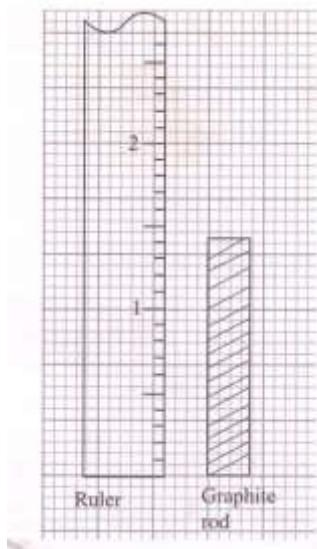


Fig. 2.1 illustrating a significant figure

All figures apart from the last are precise meaning they can be obtained again and again in an experiment. If the uncertainty in the last figure is known, it is included using a plus or minus after the measured value. Our ruler in fig 2.1 permits one to measure the length of the rod to the nearest 0.01 cm, thus the uncertainty is represented by including it in the measurement as shown  $1.35 + 0.01$  cm.

When a number is rounded off, the significant or important figure are not changed.

For example, 14728 rounded off to 3 significant figures will be 14700. In this case 1, 4 and 7 are the 3 significant figures.

The scientific notation (standard numbers) in a measurement eliminates the confusion that sometimes occurs as to whether or not zeros in a number are significant.

#### Example 2.1

Avogadro constant 602, 300,000,000,000,000,000,000 written in standard form as  $6.023 \times 10^{23}$  shows that the number has four significant figures. Thus the zero present in the figure above ate not significant.

### Example 2.2

The co-efficient of expansion of copper 0.0000167 written in standard form as notation  $1.67 \times 10^{-5}$  shows the number has three significant figures. Thus the zeros present in the figure above are not significant.

In any number, the first non-zero digit is the first significant figure. The zeros in the above examples are needed only to locate the decimal point and are, therefore, not significant.

A zero will be counted as significant figure if it comes in between other digits or appear after digit numbers written after occurrence of a decimal point.

### Example 2.3

The figure 1.907cm has 4 significant figures the “0” is counted as a significant figure because it is in between numbers. Written in standard form it will be  $1.907 \times 10^0$ .

### Example 2.4

The figure 0.0090 ha two significant figures, that is; figure ‘9’ and ‘0’ after figure 9. The reason being the zero that appear after digit numbers written after a decimal point are significant. Written in standard form it will be  $9.0 \times 10^{-3}$ .

### Example 2.5

75.000 cm contain 5 significant figures. The zeros are all significant since they are not meant to locate the decimal point. The last zero in this case indicates an estimate made to 0.001 cm

## Significant figure rules

1. Non-zero digits are always significant

### Example 2.6

1976 has 4 significant figures while 1.24 has three

2. All zero between other significant digits are significant

### Example 2.7

10078 has 5 significant figures.

10.02 has 4 significant figures.

3. The number of significant figures is determined starting with the left most non-zero digit

### Example 2.8

0.00789203

The zero on the left hand side are not significant.

The left most non-zero digit of a decimal figure is usually the most significant in this case 7 is the most significant. The zero between digit '2' and '3' is significant

4. The right most digit of a decimal figure is usually the least significant figure.

### Example 2.9

0.00543 the digit '3' is the least significant figure

5. In a figure without a decimal point, the right most non-zero digit is the least significant figure

### Example 2.10

7900, the least significant figure is '9'.

6. Zeros are significant only if they are not being used exclusively to locate the decimal point.

### Expressing the numerical results to the correct number of significant figures

Measured quantities are often used in calculations. For example

- a. In addition and subtraction

When measured quantities are used in addition or subtraction, a figure in the answer will be significant if each number in the problem contributes a significant figure at that particular decimal level.

#### Example 2.11

$$\begin{array}{r} 68.043 \text{ g} \\ + 8.22 \text{ g} \\ + 12.416 \text{ g} \\ \hline 88.679 \text{ g} \end{array}$$

Three is not significant

Six is not significant

'9' in the final answer is not significant

- b. Multiplication and division

When experimental quantities are multiplied or divided, the number of significant figures in the result is the same as that the quantity with the smallest number of significant figures.

#### Example 2.12

$$\text{Number of moles of an element} = \frac{\text{Given mass of an element expressed in grams}}{\text{Relative atomic mass of an element}}$$

The number of moles present in 24.12 g of carbon element whose relative atomic mass is 12 g will be calculated as;

$$\text{Number of moles of carbon} = \frac{24.12 \text{ g}}{12} = 2.01$$

The moles will be written as 2.0 (2 significant figures) because 12 had two and the least number of significant figures, but not as 2.010 which is a 4 significant figure.

c. Exact number

When pure numbers used in a calculation are exact rather than approximate, the accuracy of a calculation is not affected. Pure numbers have an infinite number of significant figures if they do not represent measurements. Pure numbers are easy to identify because they do not have units.

**Example 2.13**

Obtain the average titre volume from three burette readings recorded as 20.3, 20.5 and 20.1 respectively.

$$\frac{20.3+20.5+20.1}{3} = 20.63 = 20.6$$

A three significant figure should be retained in the calculation.

d. Rounding off numbers

1. When the last digit of a figure is greater than 5, increase the last remaining digit by 1.

**Example 2.14**

14.628 to 4 significant figures is 14.63

2. When the last digit is less than 5, it can be dropped, leaving the last remaining digit unchanged.

**Example 2.15**

15.473 to 4 significant figure is 15.47

## **2.2 Application of units of measurements**

The international system of units (SI) are basic units from which all other units of measurements are formed, as products of basic units.

For example

### **(a) Area**

The dimension of area may be derived by multiplying the fundamental dimension of length and width. That is,

$$\text{Area} = l \times w \text{ or}$$

$$\text{Area} = l \times l \text{ (if the length are equal)}$$

Since the SI unit of length is metre, the derived unit of area is therefore square metres as shown;

$$m \times m = m^2$$

The commonly used unit of length in Chemistry is the centimetre, the derived unit of area using centimeter units is square centimetres as shown;

$$cm \times cm = cm^2$$

### **(b) Volume**

Volume is a three dimensional quantity that is used to characterize states of matter.

Volume is obtained by multiplying unit of length by width by height;

$$\text{Volume} = l \times w \times h \text{ or}$$

$$\text{Volume} = l \times l \times l = l^3 \text{ (if all length are equal).}$$

The derived unit of volume is cubic metres  $m \times m \times m = m^3$

### **(c) Density**

Density is a derived quantity that relates the mass of a substance to its volume.

That is density =  $\frac{\text{mass of substance}}{\text{volume}}$

Given that the SI derived unit of volume is cubic metres ( $m^3$ ) the SI derived unit of density is kilogram per cubic metre (kg per  $m^3$ ).

(d) Molarity

Molarity is a derived quantity that relates the number of moles to its volume, that is,

Molarity =  $\frac{\text{number of moles of a substance}}{\text{volume}}$

The derived unit of molarity is moles per cubic metre (mol per  $m^3$ ).

### Accuracy and precision

The term precision and accuracy are widely used in sciences where measurements are made.

Accuracy of a result is the degree to which the experimental value agrees with the true or expected value.

#### For example

In the laboratory, if you measure 20.5 g of a given substance, of which the actual or known mass is 25.0 g, then your measurement is not accurate.

Precision means that the degree of closeness of two or more measurement results in a given experiment.

#### For example

- a. If you obtain a titre volume of a liquid or solution from a burette three times, and get a volume like  $23.4 \text{ cm}^3$  each time, then your measurement is precise. Measurements can be precise but inaccurate. Precise results, are results obtained using sensitive instrument that can measure small changes.

- b. The mass weighed using an electric balance (3.162 g) will be more precise than the mass weighed using a spring balance (3.16 g).
- c. A measuring cylinder with a limited graduation (markings) on it is used to obtain experiment readings, the same error will be obtained during the experiment. The reading may be somewhere in between two markings, you will have to estimate its correct position by the eye. The reading will not be accurate.

Errors make results different each time an experiment is repeated. These random errors are minimized by repeating an experiment and finding the average mean of the results.

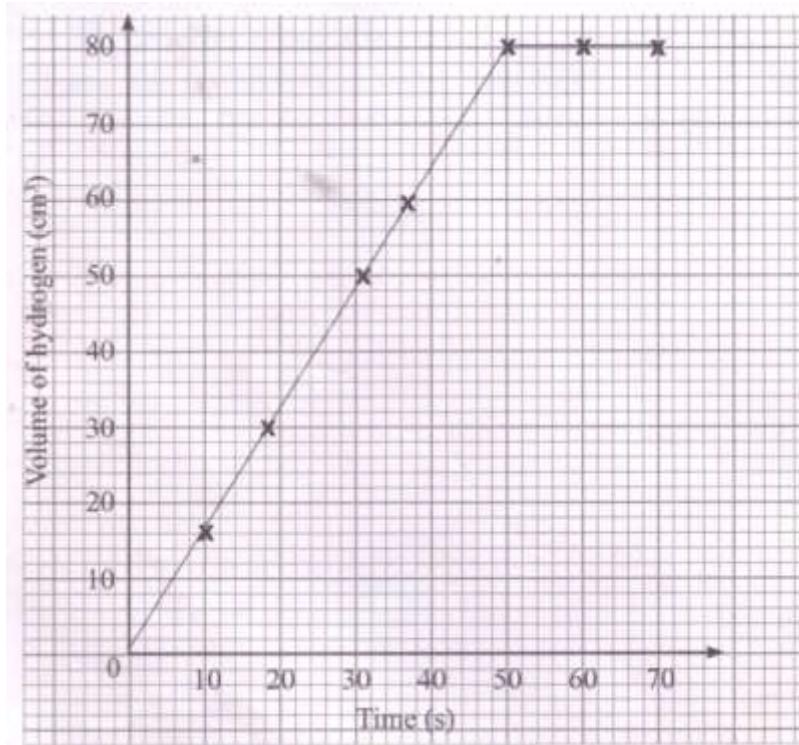
## **Graphs**

Graphs are visual representations of data values measured in an experiment. They can be in form of line graphs or pie charts. Graphs are made using the results got from an experiment, to make it easier to understand and interpret the experimental data. The different types of graphs display data in different ways for different uses.

### **(a) Line graphs**

They are used to compare two sets of continuous data.

For example, fig. 2.2 is a line graph showing volume of hydrogen gas evolved against time taken, when zinc granules react with dilute hydrochloric acid.



*Fig. 2.2 A line graph of volume of hydrogen gas evolved against time taken when zinc granules react with dilute hydrochloric acid*

### How to make a line graph

- Use a simple scale that will make it easier to plot all the points.
- Use the data from the table to determine a convenient scale.
- Draw and label the scale on the vertical axis (y-axis).
- Draw and label the horizontal axis (x-axis).
- Locate the points on the graphs
- Connect the points plotted using line segments.
- Write the title of the line graph.
- When data starts with a large value use a broken scale.

### How to interpret line graphs

The data in a line graph is represented using continuous line continuous line segments. Components on the vertical axis and horizontal axis represent values on the table (variables).

A line graph allows one to predict values of variables that are not in the table through drawing;

- A vertical dotted line from the x-axis to the line on the graph, then another one from the line graph to the y-axis to predict a quantity on the y-axis or
- A horizontal dotted line from the y-axis to the line on the graph, then another one from the line graph to the x-axis to predict a quantity on the x-axis.

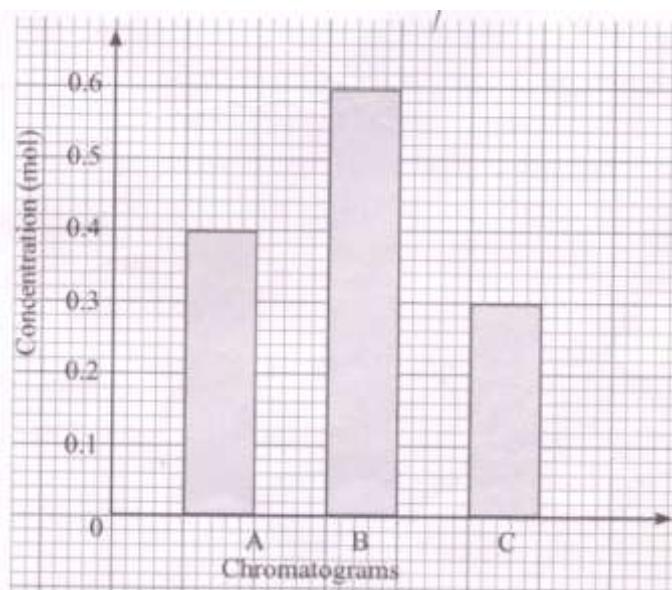
### **Advantages of line graphs**

1. They are good at showing specific values of data, meaning that when one variable is known the other variable can be easily determined.
2. They show trends in data clearly that is, how one variable is affected by the other variable as it increases or decreases.
3. They can be used to predict results of data not yet recorded.

### **(b) Bar graphs**

They are graphs that visually display information using a series of bars or rectangles. Bar graphs enable comparisons between variables, by determining the value of a given variable compared with others. They are used when one of the data sets is an ordered set of data (categorized).

For example, Fig. 2.3 is a bar graph showing concentration of solute in mol of three chromatogram samples.



*Fig. 2.3 a bar graph of concentration of solute of three chromatogram samples*

A bar graph is thus a graphical display of data using bars of different heights.

### **How to construct a bar graph**

- a. Write the title of the graph
- b. Label the axes.
- c. Use the grid lines to create the scale
- d. Use each bar to show quantity of a particular category.

### **How to interpret bar graphs**

Components of the graph namely title, axes and scale of measure are reviewed.

A title summarizes the topic, the axes shows data that has been grouped into categories or various quantities that were measured. Scale of measure determines the length of the bar and gives frequency within the categories. The data is represented using vertical or horizontal bars.

A frequency axis with an actual measurement is used to determine the quantity represented by the bar. The bar lengths are compared using the frequency (length). Note the span of the range and the breaks included within the interval.

### **Advantages of bar graphs**

- a. Bar graphs helps us to see relationships quickly.
- b. They are useful for comparing facts.

### **Types of bar graphs**

#### **(a) Horizontal bar graphs**

Horizontal bar graph contains a horizontal axis known as x-axis, the bars on a horizontal bar graph start at the y-axis and run parallel to the x-axis.

The numerical value (measured values) are recorded below the x-axis while the category names (names of bar) lies on the y-axis. They are suitable when more space is required for writing categories names that are long.

#### **(b) Vertical bar graphs.**

Vertical bar graphs contain a vertical axis known as the y-axis, the bars on a vertical bar graph start at the x-axis and run parallel to the y-axis.

The numerical values (measured values) are recorded on the y-axis, while the category names (names of the bar) lies below the x-axis. They are suitable when the names of categories on a vertical bar graph are short.

### (c) Pie charts

A pie chart is a circular graph divided into sectors that represents numerical proportions. The angle of each sector is proportional to the quantity represented.

#### How to construct a pie chart

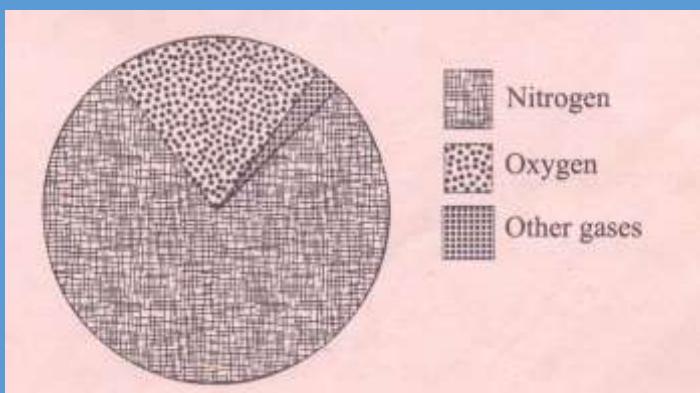
1. Construct all sectors or data points to percentages of the whole data set.
2. Convert the percentages into angle, by multiplying the percentage of a sector by  $360^\circ$  degrees.
3. Make sure the angle calculations are correct by adding all the angles, which should total up to  $360^\circ$  degrees.
4. Draw a circle on a blank sheet of paper, using a compass.
5. Using a ruler draw a horizontal line (radius) from the centre to the right edge of the circle.
6. Measure the angle of the data with a protractor, starting from the radius line, and mark it on the edge of the circle. Use the ruler to draw another radius line to that point, use the new radius as a base line for your next sector and repeat the procedure until you get the last data point. Measure the last angle to verify its value since both lines will already be drawn.
7. Label and shade the sections of the pie chart to highlight the important data.

### Example 2.16

Study the data in the table below and draw a pie chart.

Gas	Percentage composition by volume in air
Nitrogen	78.10%
Oxygen	20.90%
Carbon dioxide	0.03%
Argon	0.90%
Other noble gases	0.07%

### Solution



### Interpreting pie charts

Pie charts are graphs that measure part of variable in relation to the whole.

Interpretation of a pie chart depends on the title, the key and graph itself.

Pie charts are coloured by categories, which are distinguished by use of line or shapes for purposes of correct interpretation.

- The title should be read correctly to avoid misinterpretation
- The key helps to interpret the pie chart by assigning unique colours to each of the variables.
- Sizes of the sectors are compared percentage wise.

## Advantages of pie charts

- a. We can compare one quantity with another easily.
- b. We can compare each quantity with the total (whole).
- c. It is easy to understand and interpret the data.

## Graph data interpretation

Data interpretation problems usually requires two basic steps:

- a. One has to read a graph in order to obtain certain information
- b. The information is manipulated in order to make conclusions.

## Interpreting graphs

To interpret a graph one reads the title, looks at the key, reads the labels and then studies the graph to understand it. The title is important, because it tells what information is being displayed.

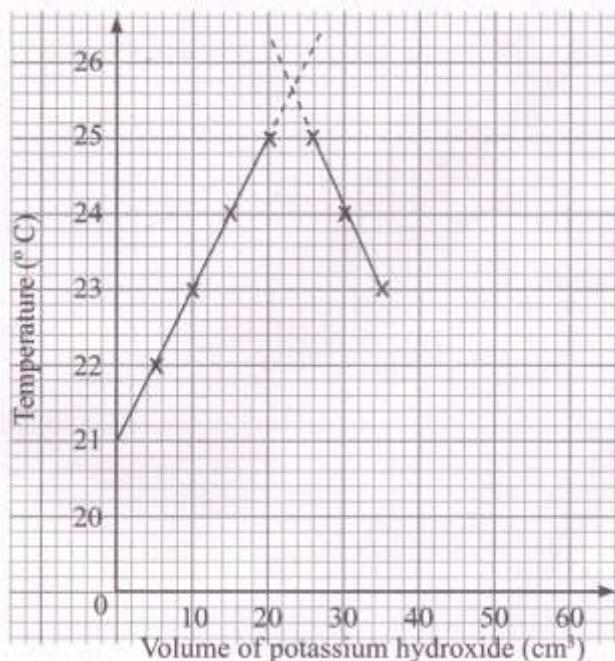
A key will explain symbols and colours used in the graph. In graphs, use of different colours will help us to quickly compare different variables because each variable has a unique and different colour representing it.

Labels of a graph tell us what variables are being displayed. Conclusions are made faster using graphs than using data table or written data.

*Table 2.0 shows a data table from a reaction between different volumes of potassium hydroxide and different volumes of hydrochloric acid and Fig. 2.4 shows a graph representation of the same data.*

**Table 2.0 Data from the reaction between different volumes of potassium hydroxide and different volumes of hydrochloric acid**

Volume of potassium hydroxide	0	5	10	15	20	25	30	35
Volume of hydrochloric acid ( $cm^3$ ) + Volume of potassium hydroxide ( $cm^3$ )	20	25	30	35	40	45	50	55
Temperature ( $^{\circ}C$ )	21.0	22.0	23.0	24.0	25.0	25.0	24.0	23.0



*Fig. 2.4 a graph of temperature against volume produced when different volumes of potassium hydroxide react with different volumes of hydrochloric acid*

From the graph, the highest temperature is 25.5, read where the extrapolated lines meet.

### (a) Extrapolation

This is the process of obtaining new data points outside a known set of data points.

#### Example 2.17

Using the data set of variables in the table below, draw a graph of temperature against time taken for the magnesium ribbon to react completely with dilute hydrochloric acid.

Time (minutes)	0	1	2	3	4	5	6	7	8
Temperature (°C) of hydrochloric acid	21.0	21.0	21.0	X	26.0	25.0	24.0	23.5	22.0
	Temperature of the acid before adding the ribbon	Time Mg is added		Temperature of the acid with time after adding magnesium					

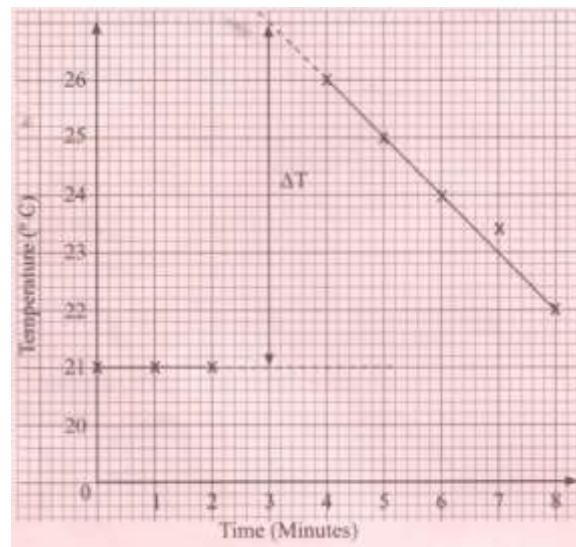


Fig. 2.5 A graph of temperature against time taken for magnesium ribbon to react completely with dilute hydrochloric acid

From the extrapolated graph the highest temperature reads 27.0 °C whereas the lowest temperature reads 21.0 °C

(b) Interpolation

This is an estimation of a value that lies within two known values that are in a sequence of values.

Examples 2.18

Using the data set variables in the table draw a graph of change in volume (in litres) against change in temperature in (kelvin).

Volume( $dm^3$ )	1	2	3	4	5
Temperature (Kelvin)	120	240	360	480	600

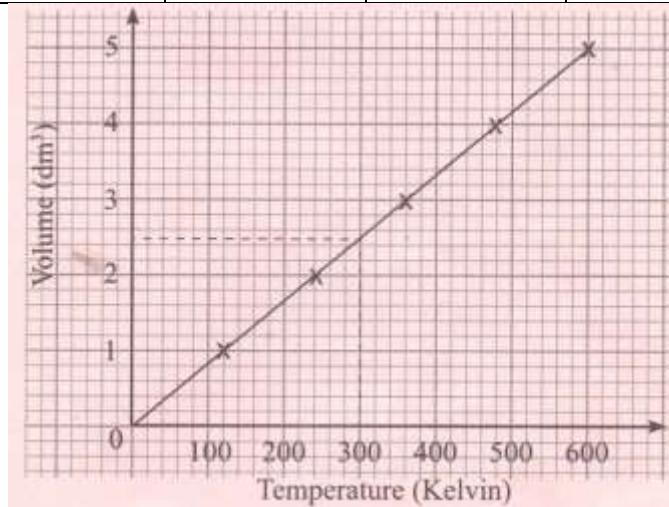


Fig. 2.6 A graph of volume against temperature

Use the above graph to interpolate (predict) the volume occupied by the gas in litres when temperature is 300K. This is done by drawing a vertical dotted line from the x-axis to the straight line of the graph, and then drawing a horizontal dotted to the y-axis. They y-value at this point, reads  $2.5 dm^3$  which is equal to the volume occupied by the gas at 300 K.

## Review Questions

1. Write these numbers in standard form.
  - (a) 0.057
  - (b) 9
2. Write out these numbers in ordinary notation.
  - (a)  $7.46 \times 10^{-2}$
  - (b)  $1.5 \times 10^3$
3. Indicate the number of significant figures in each of the given numbers.
  - (a) 0.008
  - (b) 477.0
  - (c)  $1.8 \times 10^3$
  - (d)  $0.5 \times 10^2$
  - (e) 1800
4. Perform these operations and express the answer to the proper number of significant figures.
  - (a)  $223.46 + 119.32$
  - (b)  $76 \times 0.0000023$
  - (c)  $1.67 \div 52.8$
5. Round off these numbers to 3 significant figures.
  - (a) 348.6
  - (b) 1989
  - (c) 0.0008297
6. Round off these numbers to 1 significant figure.
  - (a) 157
  - (b) 80.77
  - (c) 0.09483

7. Convert,
- 5689 g to kg
  - $770000 \text{ cm}^3$  to  $\text{m}^3$
  - $0.016 \text{ m}^2$  to  $\times \text{cm}^2$
8. The table below shows the temperature at which crystals appeared when 6 g of oxalic acid were dissolved in different volumes of hot water. Use the data to draw a graph of temperature against volume of water.

Volume of water in $\text{cm}^3$	4	6	8	10	12
Temperature at which first crystal appear ( $^\circ\text{C}$ )	62.0	58.0	53.0	48.0	42.0

Use the graph predict (interpolate);

- Temperature at which crystals would first appear if the total volume of water was  $7 \text{ cm}^3$
- Mass of solid that would dissolve in  $100 \text{ cm}^3$  at  $60^\circ\text{C}$ .

## UNIT 3

## COMPOSITION AND CLASSIFICATION OF MATTER

### Matter

In primary school science, we learnt about **matter**. Matter is defined as anything that has mass and occupies space. For example a pencil, book, bicycle, water, paraffin, oxygen and carbon dioxide.

Give two more example of matter.

How can we show that matter occupies space and has mass? Mass is related to weight or heaviness. The heavier the substance is the more mass it has. Space is related to volume. The larger the volume, the more space it occupies.

### **Experiment 3.1**

#### **Aim: To show that matter occupies space**

#### **Apparatus and chemicals**

- Beaker
- Boiling tube
- Water

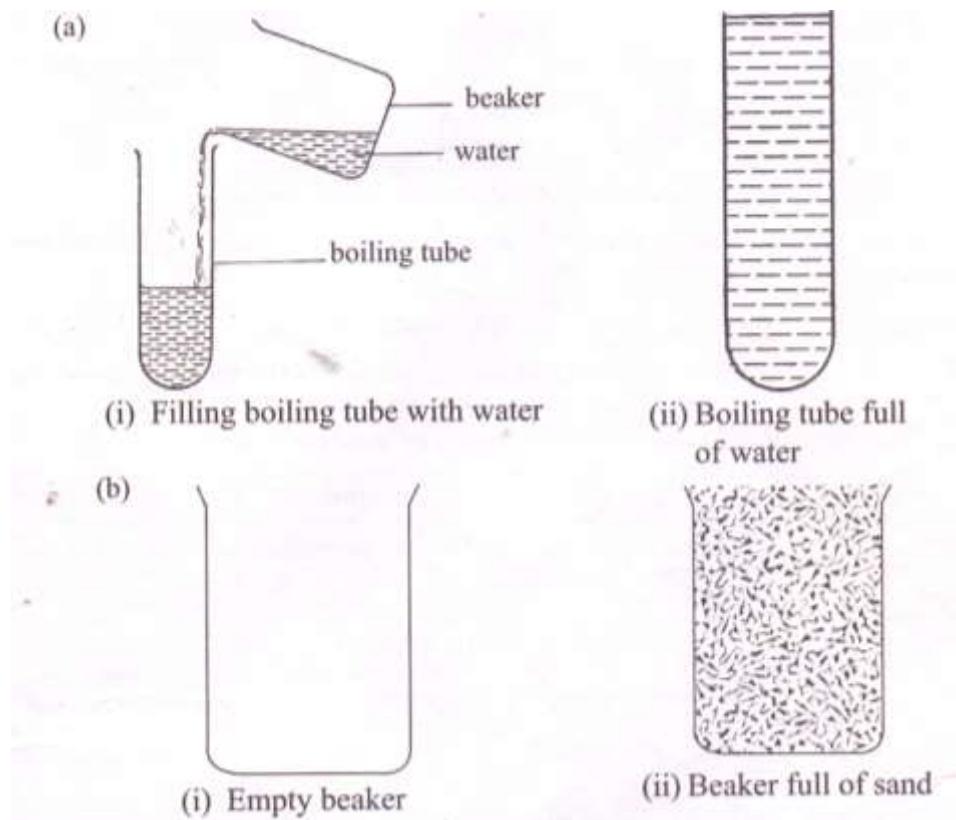
#### **Procedure**

1. Fill the beaker or boiling tube completely with water
2. Add more water.

What happens when you add excess water?

Was your beaker initially empty?

Liquids and solids fill a container. Once the container is completely full, if you add more of the liquid or solid it will spill over. Before the container is filled with a liquid or a solid, it is not as empty as it appears. It is usually full of air. It is only that we cannot see air. The liquid or solid displaces the air in the container! This experiment shows that liquids, solids and gases occupy space.



*Fig. 3.1 (a) and (b) to show that matter occupies space*

Matter exists in three different forms namely:

- Solids – for example stone, 10 tambala coin and wood.
- Liquids – for example water, paraffin and cooking oil.
- Gases – for example air, oxygen, hydrogen and carbon dioxide

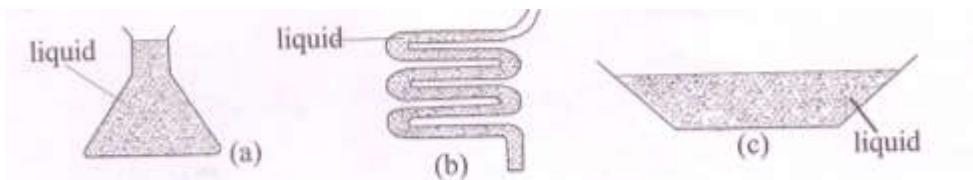
### Properties of matter

What is the main difference between these states of matter? Check your differences with your friend's differences.

In Chemistry we study how matter behaves and how the state of a substance can be changed from one form into another under certain conditions. For instance, the volume of a gas can be reduced by increasing the pressure at a given temperature. A gas can be compressed, liquids can only be slightly compressed but the volume of a solid cannot be changed by pressure.

### Properties of the three states of matter

- A solid has a definite shape and a fixed volume.
- A liquid has no definite shape. It takes up the shape of the container in which it is put, but it has a fixed volume.



- A gas (or vapour) has no definite shape or volume – it takes up all the space in a container.

These properties are physical properties of the states of matter.

The following table shows a summary of the properties of the three states of matter.

**Table 3.1 Summary of differences in properties of states of matter**

Property	Solid	Liquids	Gas
1. Shape	Definite shape	No definite shape	No definite shape
2. Volume	Fixed volume	Fixed volume	No definite volume
3. Density	Very high	high	Low
4. Flow	Does not flow	Flows easily	Flows easily
5. Packing of particles	Closely packed	Some spaces left	Particles are far apart

The three states of matter will respond to certain changes. For example, when temperature is increased, they all expand (increase in volume). The effect is much bigger for a gas (which will occupy a much bigger volume) than a solid or a liquid.

### Experiment 3.2

**Aim:** To show that matter has mass

**Apparatus and chemicals**

- 2 empty tins
- Sand

### ***Discussion***

You will realize that the tin, although initially full of air is lighter than when it is full of other substances like sand or water.



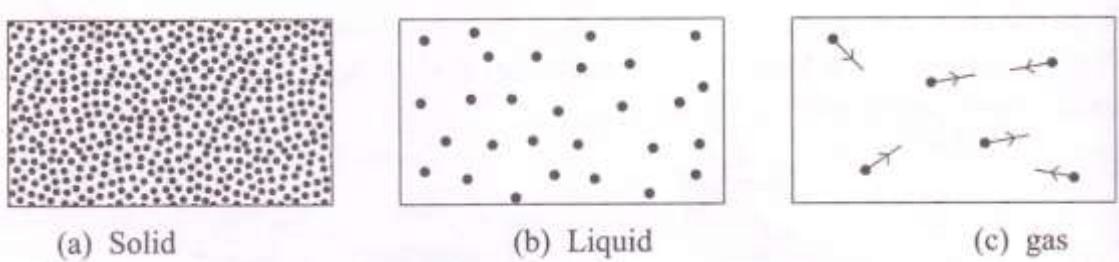
*Fig 3.3 to show that matter has mass*

### **How matter is put together**

Perhaps you wonder why solids keep their own shape but liquids and gases do not. To find out why you have to study and understand how matter is put together.

What happens when you break a piece of chalk into smaller pieces? The smaller pieces you made could be broken into still small pieces that are sometimes called particles. Each particle is too small to be seen.

When many particles group together you are able to see the substance (matter). The particles in solids, liquids and gases are grouped together in different ways. In Fig. 3.4 the little dots represents the particles. It shows the particles might look in a solid, liquid and gas.



*Fig. 3.4 Arrangement of particles in solids, liquids and gas*

The particles in a solid are in fixed positions and are arranged in a regular pattern. In liquid and gas, the particles are not in fixed positions, hence a liquid and a gas flow. Later on you will observe that the particles in a gas are always moving in different direction all the time.

You can measure the mass of an object in units called grams. A gram is a small amount of mas. What is the mass of your pencil in grams?

The mass of larger objects is measured in units called kilograms. A kilogram is 1000 grams. What is the mass of this text book in grams? How many exercise books would you need to have a mass greater than a kilogram? Use a balance and a set of gram weights to find out that air has mass and therefore it is a matter. We can also measure the volume of matter. How?



*Fig. 3.5 measuring the mass of matter*

## The particulate nature of matter

We learnt that matter occupies **space** and has **mass**. We learnt that under different conditions, matter can exist as a solid, liquid or gas. By now you should be familiar with the properties of the three states of matter.

In this sub-unit, we shall learn about what makes up matter. We shall also learn about the behavior of matter under different conditions of temperature.

### Experiment 3.3

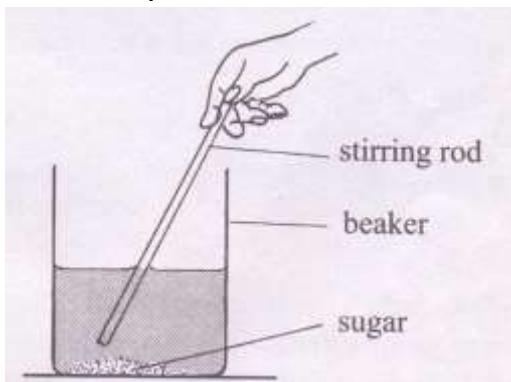
**Aim: To show that matter is made up of tiny particle**

**Apparatus and chemicals:**

- A beaker
- Sucrose (sugar)
- Glass rod
- Water

**Procedure**

1. Put about  $20\text{ cm}^3$  of water in clean beaker.
2. Add 1 tablespoonsful of sugar.
3. Stir the solution
  - What do you observe?



*Fig. 3.6 dissolving a sample of sugar.*

4. Repeat the experiment with coffee instead of sugar

Sugar disappears because it has dissolved in the water. The sugar breaks down to tiny particles which we cannot see. These particles are distributed throughout the solution.

The coffee will also break down to tiny particles which will also be distributed throughout the solution. The solution will be coloured black.

Other coloured substances can be used to show the distribution of particles in water as shown in the experiment below.

### Experiment 3.4

#### Aim: To investigate movement of colour (particles) in a liquid

#### Apparatus and chemicals

- Beaker
- Glass tube
- Potassium permanganate or blue copper (II) sulphate crystals
- Water

#### Procedure

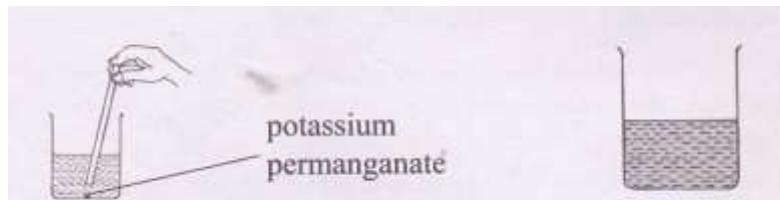


Fig. 3.7 placing a crystal of potassium permanganate in water

1. Put  $30\text{ cm}^3$  of water in a beaker
2. Place a crystal potassium permanganate at the bottom of the beaker with the help of a glass tubing as shown in Fig. 3.7
3. Close the end of the glass tubing with the index finger and remove the glass tubing
4. Leave the beaker undisturbed for an hour.
  - What do you observe?  
The water slowly turns purple.

Both the potassium permanganate crystal and water are made of particles. The purple colour of potassium permanganate spreads out because the particles leave

the crystal and mix with water particles. They are distributed throughout the solution which eventually becomes purple.

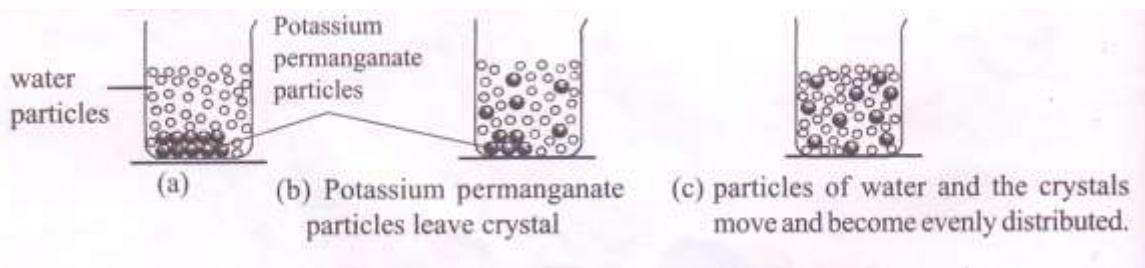


Fig. 3.8 Potassium permanganate particles distribute in water evenly.

The solute particles move throughout the liquid. This experiment shows that matter is made up of tiny particles. The potassium permanganate particles move slowly from the crystal into the water because they keep on colliding with water particles as they move, eventually becoming evenly distributed and colour of water turns purple. This is called diffusion.

The movement of particles also takes place in air. You may have smelled the aromas of various plants or fruits for example pineapples, bananas, rose flowers among others. The same things happens when we smell a nice perfume, or a foul smelling gas like from a rotten egg.

Experiment 3.5 also shows how gas particles spread in the air.

### Experiment 3.5

Aim: To show that gases consist of particles that move

Apparatus and chemicals

- Gas jars
- Bromine
- Glass tubing

### Procedure

1. Using a pipette or glass tubing, place a drop of bromine into a gas jar A as shown in fig. 3.9 (a).
2. Invert another gas jar B over gas jar A. (See fig. 3.9 (b)

3. Wait for a few minutes. What do you observe?

- Why do we have to wait for a few minutes?
- Bromine in a vacuum will spread very fast. Explain why?

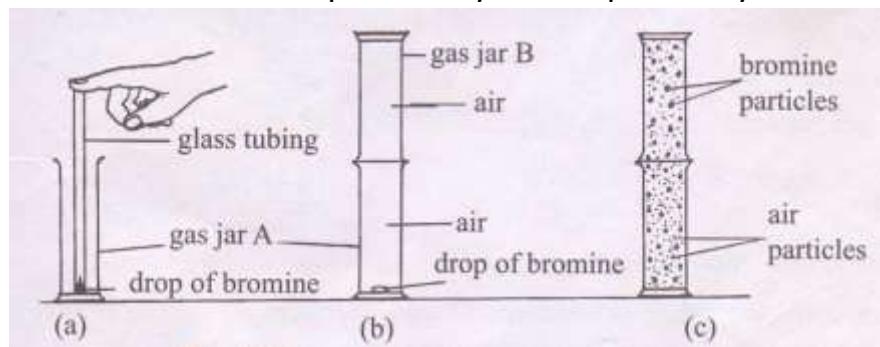


Fig. 3.9 Bromine particles spread in air

Note: Caution should be given to the poisonous nature of bromine gas.

Experiment should be performed in a fume chamber.

Both gas jar will be filled with a red-brown colour. Air is colourless while bromine vapour is red-brown. It is also heavier than air. When the jar of air, (B) is inverted over a jar of bromine, the red-brown colour spreads up into it. After a few minutes, the gas in both jars is red-brown.

The reason for this observation is that both air and bromine are made of tiny particles that are in motion. As they move, these particle collide with one another and with the walls of the jars. These collisions cause the particles to change direction. Eventually they become evenly distributed hence the red-brown colour. This experiment shows that the gases are made up of particles. The particles of bromine diffuse from the lower gas jar where they are more concentrated to the upper jar where they are less concentrated. This is called diffusion. The smaller stable particle of matter is known as an atom. So matter is made of atoms.

### Change of state of matter

Change of state occurs through either.

- Melting or freezing
- Evaporation or condensation

Changes in temperature can cause a substance to change its physical state. Water can be a solid (ice), a liquid (water) and a gas (vapour or steam). Its state can be changed by heating or cooling.

### Experiment 3.6

Aim: to investigate what happens when ice heated to boiling

Apparatus and chemicals

- Beaker
- Tripod stand
- Bunsen burner
- Thermometer
- Wire gauze
- Ice cubed

Procedure

1. Arrange the apparatus as in fig. 3.10
2. Place some ice cubes in a beaker. Place a thermometer in the beaker containing the ice cubes and record the temperature.
3. Warm the beaker and its contents gently as in Fig. 3.10 (b). Note and record the temperature at which the ice changes to liquid water.
4. Continue heating until water changes to vapour as shown in fig. 3.10 ©.
5. Record the temperature at which bubbles appear.

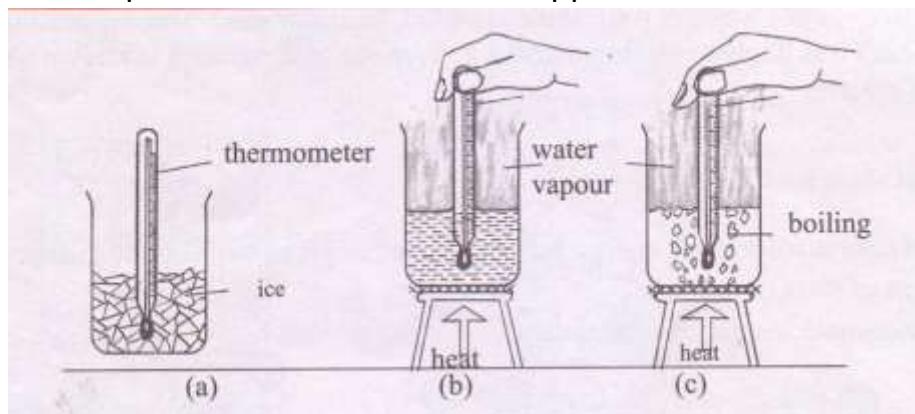


Fig. 3.10 Heating ice

6. Describe your observations as you were heating ice to change into vapour. What conclusion do you draw from this experiment?

When ice is heated, its temperature rises steadily until it all changes to liquid. The change is called melting and this temperature is melting point. It is  $0^{\circ}\text{C}$  for pure water. On further heating, the temperature rises steadily, and some of the water changes to water vapour (or steam). This change is called evaporation. The hotter the water gets, the quicker it evaporates and soon bubbles appear. The change is now known as boiling and this temperature known as boiling point of water. At sea level the boiling point of pure water is  $100^{\circ}\text{C}$ .

The change of state from solid (ice) to liquid (water) and then from liquid to gas (vapour) can be reversed by cooling. On cooling, the gas changes into water. This process is called condensation. On further cooling (to below  $0^{\circ}\text{C}$ ), the liquid changes into solid. The process is called freezing. The freezing point of water is the same as the melting point of ice ( $0^{\circ}\text{C}$ ). All these changes can be represented as in Fig. 3.11

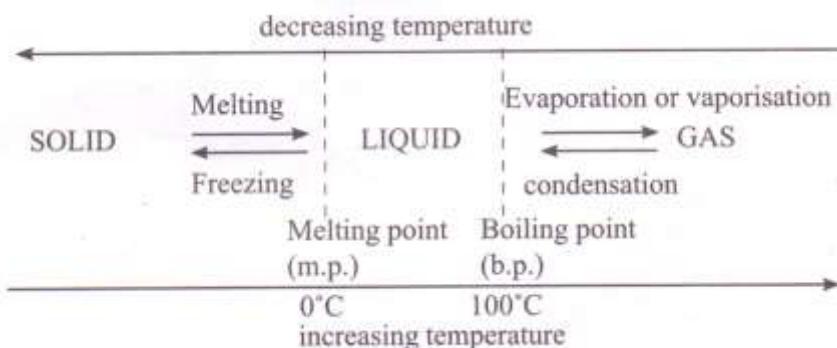


Fig. 3.11 Changes in physical state of water

We can conclude from the changes observed that something happens to cause changes in state of matter. The readings we note on the thermometer enable us to predict that the changes in temperature must, in a way be responsible for determining the state of matter.

But what determines whether a substance is solid, liquid or gas? The explanation of the state in which it is likely to be found under a given set of conditions is known as kinetic theory of matter.

#### Changes of state and kinetic theory

Changes of state involves heat energy being absorbed or given out. Kinetic theory explains the changes of state of matter.

### Intermolecular forces of attraction

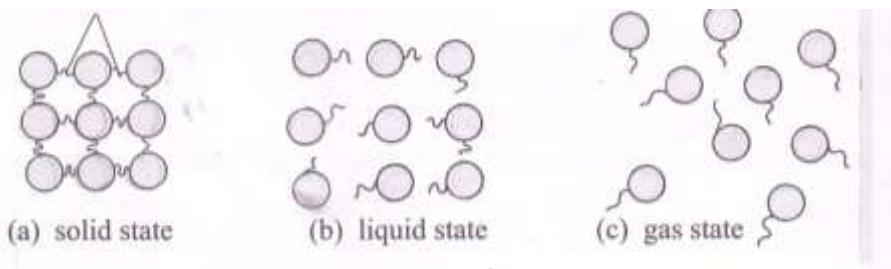


Fig. 3.12 Arrangement of particles in the three states of matter

#### (a) Solids

- Solids have strong forces of attraction between particles.
- Particles are held in fixed positions by the attractive forces.
- When a solid is heated, the energy is gained by the particles. The extra energy makes the particles vibrate more and more, but within the same positions.

Further heating weakens the forces of attraction between the particles even more and the solid changes to liquid. This is melting.

#### (b) Liquids

- Liquids have moderate forces of attraction between particles
- The particles are free to move randomly but tend to stick together. So, all liquids do not have definite shape and fill the bottom of any container Fig 3.13 (a).

#### (c) Gases

- When a liquid is heated, the heat energy goes to the particles and makes them move faster.
- The fast-moving particles at the surface overcome the forces of attraction from the other particles in the liquid and escape into the air. This is evaporation. Fig. 3.13 (b).

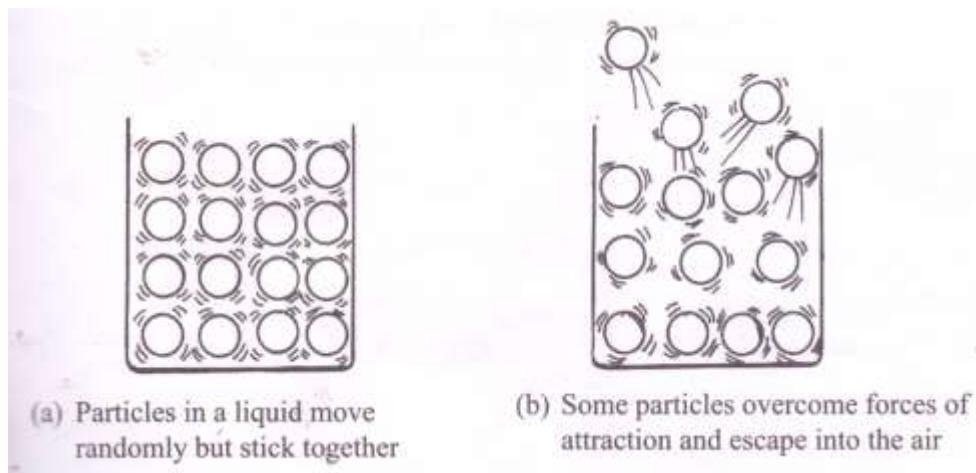


Fig. 3.13 Movement of particles in a liquid

The above changes of state can be summarized as in fig. 3.14

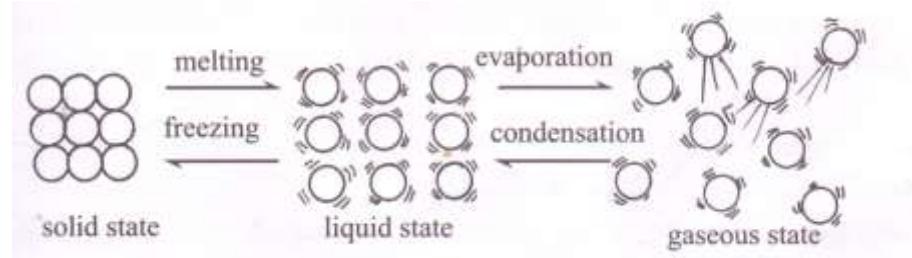


Fig. 3.14 Movement of particles in the three state of matter

- When the liquids get hot enough all the particles acquire enough energy and speed to overcome the forces of attraction between them and most escape from each other to the air.
  - During boiling big bubbles of gas from inside the liquid as the particles break away from each other.
- It is possible to compress gases easily because there is a lot of free space between the particles, unlike in solids and liquids.

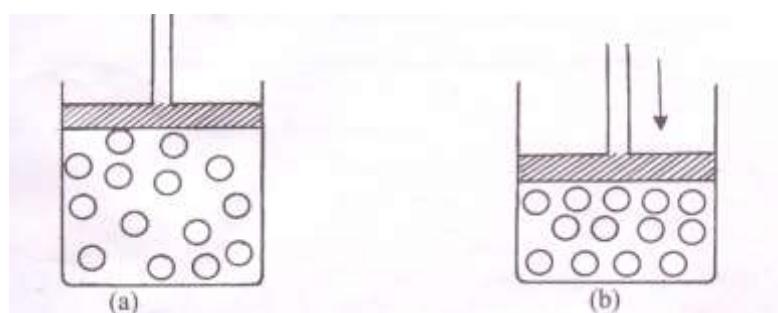


Fig. 3.15 Compressing a gas

During condensation, the kinetic energy of the particles of gas is decreased. The particles come closer together, and the forces of attraction between them increase. A liquid is formed. This process is called condensation.

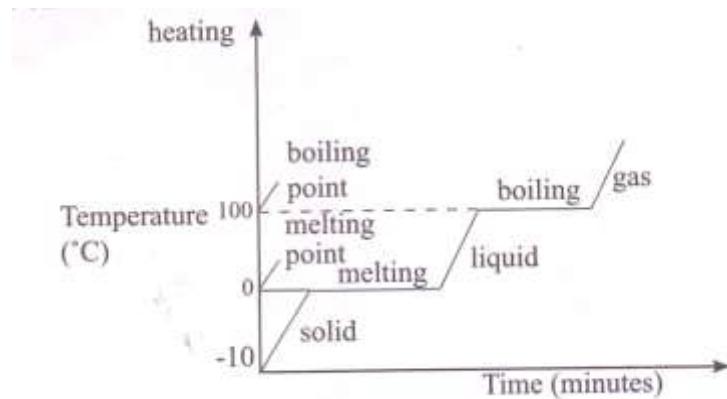


Fig. 3.16 Heating curve showing changes of state

A thermometer is used to measure temperature. Water freezes when the temperature drops to zero degrees Celsius. This can also be written as  $0^{\circ}\text{C}$ . This temperature is called the freezing point of water.

Ice melts when temperature of ice rises to  $0^{\circ}\text{C}$  on a thermometer. So  $0^{\circ}\text{C}$  is both the freezing point and the melting point of water.

This involves decreasing the kinetic energy of liquid particles even more. Particles come even closer together and the forces of attraction between them increase, so that they only vibrate within the same positions. A solid is formed. This process is called **freezing**.

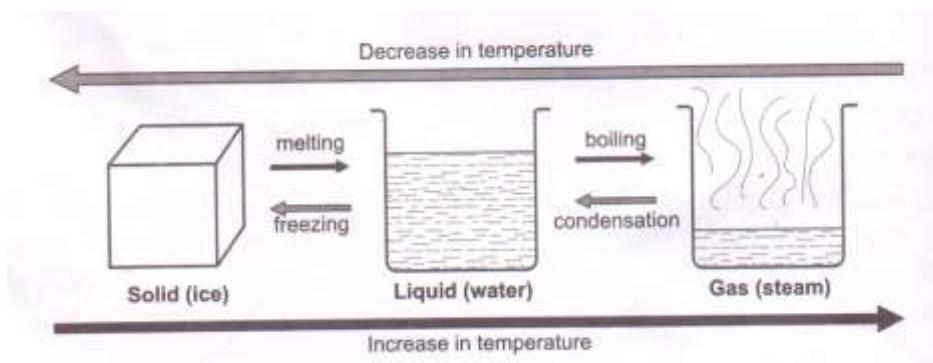


Fig. 3.17 Effect of heat on ice

## Elements, atoms, molecules, compounds and mixtures

### Experiment 3.7

Aim: To distinguish between elements, compounds and mixtures

#### Apparatus and chemicals

- Watch glass
- Splint
- Iron filings
- Dilute hydrochloric acid
- Test-tube
- Magnet
- Sulphur powder

#### Procedure

1. Mix some iron filings and sulphur powder in any proportion and place the mixture in a watch glass.
2. Pass a magnet over the mixture and note what happens.
3. Put half of the mixture in a test tube, add water and shake. Does it dissolve?
4. Allow the mixture to settle. How many layers can you see? Which is the bottom layer?
5. Put the rest of the mixture in a test tube and add a few drops of dilute hydrochloric acid. Apply a burning splint. What do you observe?
6. Now place a mixture of 7 g of iron filings and 4 g of sulphur powder in a test tube. Heat the bottom of the tube.

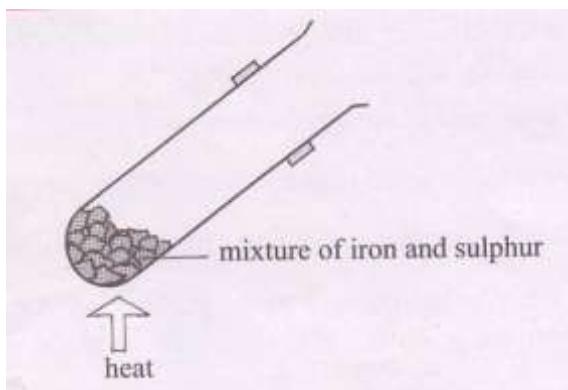


Fig. 3.18 preparing iron (II) sulphide

7. When the mixture starts to glow, remove the flame and note that the glow starts to spread slowly through the mixture without further heating. Allow the product to cool, and then remove it.
8. Note the colour of this product. Pass a magnet over it. Is there any attraction? To half of this product in a test tube, add some water and shake. Is there any sign of separation?
9. Put the rest of the product in a test tube and add a few drops of dilute hydrochloric acid. Smell the gas produced cautiously. Describe the smell.

From the experiment, the observations are as follows:

- Before heating, the iron could be separated from sulphur by sedimentation or using a magnet. These are physical methods. These methods could not separate iron and sulphur after heating.
- Before heating, iron reacted with dilute hydrochloric acid to form a gas (that produces a “pop” sound with a burning splint). Sulphur was not affected. The product of heating iron and sulphur produced a gas with a pungent smell (rotten egg smell).
- The mixture of iron and sulphur was an average of grey (for iron) and yellow (for sulphur) before heating. After heating, the product has a dark grey almost black colour.
- In mixing iron and sulphur no heat was produced. But after heating the mixture, the formation of the new substance produced enough heat resulting in a bright red glow.

## Elements

There are substances that cannot be split (broken down) into simpler substances by any chemical means for example the iron and sulphur which form iron (II) sulphide. The two substances are called elements.

An element can therefore be defined as a pure substance which cannot be split into anything simpler by any chemical process.

There are about 118 known elements. About 90 of these occur naturally on the earth and its atmosphere while the rest have been made in the laboratories. Most elements can be classified as;

- Metals or

- Non-metals.

All metals, except mercury, are solids at room temperature. They are also good conductors of electricity. (Graphite is a non-metal but is a good conductor of electricity).

In the following tables 3.2 is a list of elements that you will come across frequently in your chemistry course.

Table 3.2 some examples of elements

Metals	Non-metals
Sodium	Carbon
Magnesium	Nitrogen
Calcium	Silicon
Zinc	Oxygen
Iron	Sulphur
Lead	Chlorine
Copper	Bromine
Silver	Iodine
Gold	Fluorine
Platinum	Hydrogen
Tin	Neon
Barium	Argon
Beryllium	Boron
Mercury	Phosphorus
Uranium	Helium
Cobalt	
Chromium	
Manganese	
Aluminium	

### (a) Atoms

In Chemistry, these particles can be divided even further until no further division is possible. The smallest particle that cannot be divided further is called an atom of that element.

We have seen that matter is made of tiny particles. These particles can be divided further by chemical means. When they cannot be divided any further without changing the properties of the element, the particles at this stage are called atoms.

An atom is the smallest particle into which an element can be divided without losing the properties of the element.

Single atoms are far too small to be seen, even by the most powerful microscope for example, about four billion iron atoms would fit side on the full stop at the end of this sentence.

An atom of one element is different from an atom of another element. That means that an iron atom is different from a copper atom. A copper atom is different from a zinc atom and so forth.

We can therefore conclude that there are as many atoms as there are elements. There are about 188 different types of elements and therefore there are also 118 different types of atoms.

Atoms do not exist on their own; they usually take part in a chemical reaction.

### (b) Compounds

We have already seen that when substances like iron and sulphur are heated, they combine chemically to form a new substance called iron (II) sulphide. A substance formed by chemically combining two or more elements together is called a compound.

Therefore, we can define a compound as a substance made up of two or more elements combined together by chemical means. For instance, when magnesium (one element), is burnt in air, it combines with oxygen (another element) to form magnesium oxide. Magnesium oxide is therefore a compound of magnesium and oxygen elements.

Other examples of compounds are water, a compound of hydrogen and oxygen; common salt, a compound of sodium and chlorine; copper carbonate a compound of (copper, carbon and oxygen). A compound has different chemical and physical properties from those of the elements of which it is composed of.

Table 3.3 gives examples of compounds and the elements that they are made of. You will learn more compounds in Chemistry later.

Table 3.3 Examples of compounds and the elements that they are made up of

Name of the compound	Elements in the compounds
1. Magnesium oxide (white ash)	Magnesium – grey metal Oxygen – colourless gas
2. Water (liquid at room temperature)	Hydrogen – colourless gas Oxygen – colourless gas
3. Copper carbonate (green powder)	Copper – red-brown metal Carbon – black solid Oxygen – colourless gas
4. Carbon monoxide (colourless gas)	Carbon – black solid Oxygen – colourless gas
5. Copper (II) oxide (black powder)	Copper – red-brown metal Oxygen – colourless
6. Copper (II) nitrate (green solid)	Copper – red-brown metal Nitrogen – colourless gas Oxygen – colourless gas

### Chemical symbols of elements

In Chemistry, names of elements are written in shorthand form known as chemical symbols. This system of writing symbols uses letters taken from the name of the element. This could be the English or Latin name of the element. The symbol of an element consists of one or two letters. The first letter of a chemical symbol must always be a capital letter. The letters should not be joined like in handwriting. They must be printed. These symbols are an international code.

This means that all over the world, symbols are written in the same way no matter how people spell the name of the element.

The names of several elements begin with the same letter. In such cases, the second or third letter of the English or Latin name is included. This is always a small letter. Table 3.4 shows some elements whose chemical symbols take up the first letter of their English name.

Table 3.4 some elements and their symbols

Elements	Chemical symbols
Carbon	C
Hydrogen	H
Nitrogen	N
Oxygen	O
Phosphorus	P
sulphur	S

Table 3.5 shows some elements whose chemical symbols use two letters from the English name of the element.

Table 3.5 some elements and their symbols

Element	Symbol
Calcium	Ca
Cobalt	Co
Chlorine	Cl
Magnesium	Mg
manganese	Mn

Table 3.6 shows elements whose chemical symbols use one or two letters from the Latin name of the element.

Element	Latin name	Chemical symbol
Potassium	Kalium	K
Sodium	Natrium	Na
Iron	Ferrum	Fe
Lead	Plumbum	Pb
Silver	Argentum	Ag
Copper	Cuprum	Cu
Mercury	Hydrargyrum	Hg
Gold	aurum	Au

We do not expect you to learn the Latin names. But you should learn the symbols and read them in English. For instance, Cu is read as Copper.

Note: The symbol of each element represents one atom of that element. For instance:

Mg represents one atom of magnesium.

2Mg represents two atoms of magnesium.

## Molecules

All compounds consists of two or more elements chemically combined together and exist as single particles. The smallest particle of an element or a compound which can normally exist in a free and separate state is called a molecule. This means that there are some elements whose molecules consist of one atom while others have two or more. Molecules of a compound have two or more different atoms.

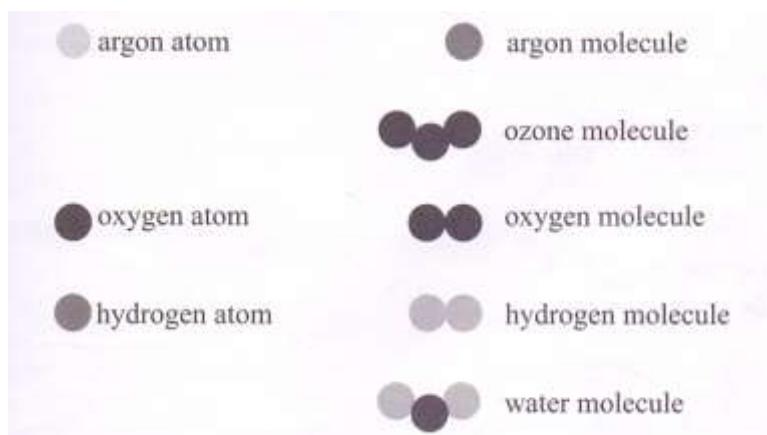


Fig. 3.19 Atoms and molecules

Note: Molecules composed of one atom. Like argon and the other noble gases are called monatomic. Molecules composed of two atoms like oxygen, hydrogen, nitrogen and other are known as diatomic. Molecules composed of three atoms like ozone are known as triatomic.

There are molecules which are made by combining different elements like in water.

Determining type and number of atoms in chemical formulae of compounds

A correct formula of a compound or substance shows the atoms present in compound in their simplest ratio. For example, the formula of copper (II) sulphate is  $CuSO_4$ . It contains 1 atom of copper, 1 atom of sulphur and 4 atoms of oxygen.

We will learn more on chemical formula in unit 6. Below is the table with other compounds and the type and number of atoms they contain.

Table 3.7 some compounds and type and number of atoms they contain

Compound name	Formula	Atoms of elements present
Common salt (sodium chloride)	NaCl	1 Sodium (Na) and 1 Chloride (Cl)
Water	$H_2O$	2 Hydrogen (H) and 1 Oxygen (O)
Calcium	$CaCO_3$ .	1 Calcium (Ca), 1 Carbon (C) and 3 atoms of Oxygen (O)

### Pure substances and mixtures

We say something is pure if it consists of one substance only. We are familiar with clean water. If you filter muddy water the filtrate is clear. Is the filtrate pure water? Is clean water from rivers, streams and well pure?

Often it is difficult to tell by mere appearance whether a substance is pure or not. A substance may appear as one, yet it is a mixture of two or more substances.

Name substances that you know that appear as a single substance but are mixtures. Are the following substances pure?

- Tap water
- Rain water
- Well or spring water
- Alcohol (ethanol)
- Brass or bronze
- Metals – copper, gold, iron
- Common salt (sodium chloride)

### Criteria for purity

We can use the melting or boiling to test the purity of a solid or liquid. A pure substance melts and boils at a definite, a sharp temperature.

## Experiment 3.8

Aim: To determine the melting point of naphthalene

Apparatus and chemicals

- A thermometer
- Stand and clamp
- Burner
- Boiling tube
- A melting point tube
- Naphthalene
- Beaker
- Stirrer
- Water
- Camphor

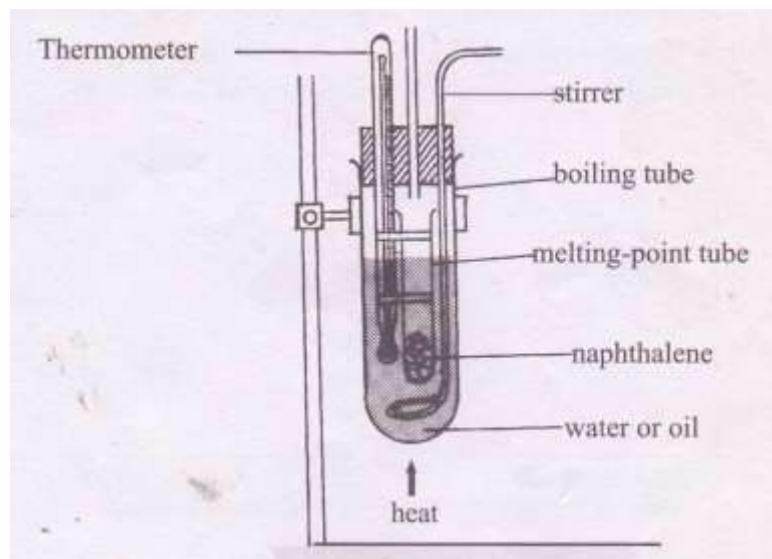


Fig. 3.20 Apparatus for measuring the melting point of a solid

1. Set up the apparatus as in Fig 3.20
2. Using a rubber band, tie the thermometer and the melting-tube together.
3. Place the thermometer and the tube in a water-bath and heat the water gently while stirring continuously.
4. Note and record the temperature at which naphthalene melts.
5. Repeat the experiment using a mixture of naphthalene and camphor (camphor is an impurity in this case).

6. Copy Table 3.9 and fill the empty column

Which of the two has a sharp melting point?

Table 3.8 Temperature results of melting Naphthalene and Naphthalene + Camphor mixture

Substance	Temperature °C
Melting point of Naphthalene	
Melting point of naphthalene + camphor	

What effect does camphor have on melting point of naphthalene?

When naphthalene is melting, the temperature stops rising. It will only rise again when all naphthalene has melted. Naphthalene as a single pure substance melts at 80°C. An impure substance melts over a range of temperature, not at a particular point.

Note: if a solid has a melting point above 100°C, an oil bath is used instead of a water bath. This makes it possible for melting points above 100°C to be measured.

It is possible to follow the temperature of a solid before and after melting. The results can then be used to plot a graph and produce a heating curve as in Fig 3.21.

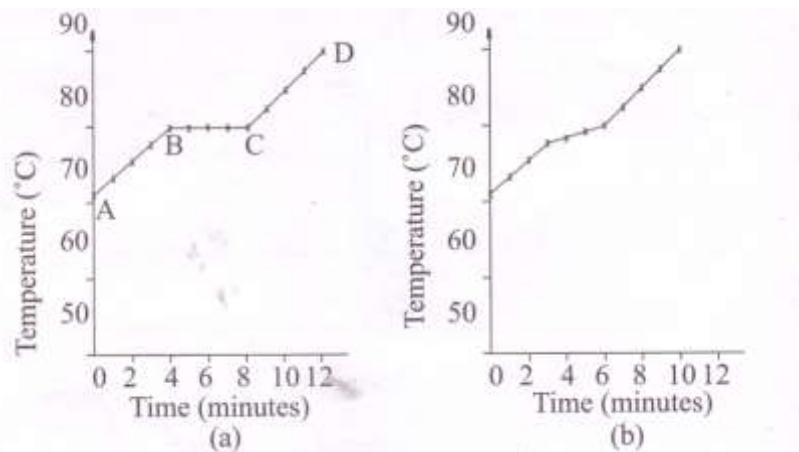


Fig 3.21 (a) The heating curve for naphthalene (b) The heating curve for impure naphthalene

Portion AB: Naphthalene absorbs heat energy and the temperature increases steadily.

Portion BC: Pure naphthalene changes its state from solid to liquid. The temperature stays constant at 80°C until all naphthalene melts. All the heat absorbed is used to change solid naphthalene into liquid therefore no temperature rise.

Portion CD: The liquid absorbs more heat energy. The temperature rises. It stops rising when the boiling point of the liquid is reached.

Note: The heating curve of the impure substance has no horizontal portion because impure substances melt over a range of temperatures.

### Experiment 3.9

Aim: To determine the boiling point of pure water

Apparatus and chemicals

- Boiling tube
- Thermometer
- Burner
- Stopper
- Stand and clamp
- Water

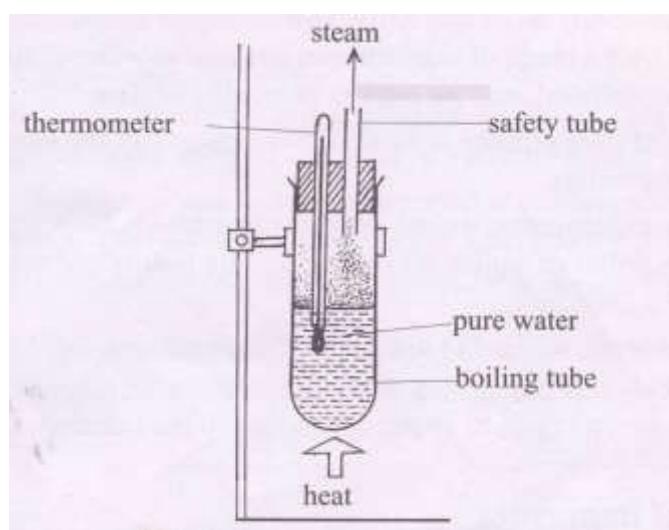


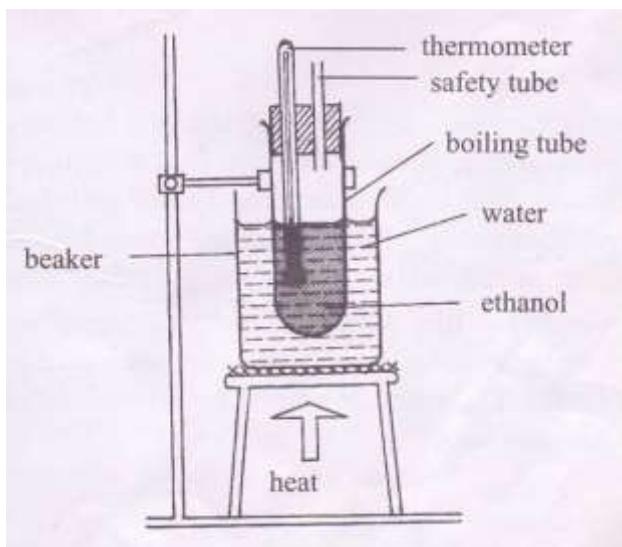
Fig. 3.22 Apparatus for measuring boiling point of pure water.

## Procedure

1. Set up the apparatus as shown in Fig. 3.22
2. Put distilled water in a boiling tube and heat.
3. Note and record the highest steady temperature reached on the thermometer.
4. Repeat the procedure using ordinary water and salty water.

Which one boils at a higher temperature, pure water or tap water? Explain.

The boiling point of pure ethanol can be measured in a similar manner. But ethanol is gently heated to the boiling point using a water bath instead of direct flame because ethanol is a volatile liquid and flammable and therefore can catch fire (see Fig. 3.23).



*Fig. 3.23 Apparatus for measuring boiling point of ethanol*

Ordinary water boils at a slightly higher temperature than distilled water because it contains dissolved substances. Salty water will boil at an even higher temperature because of the salt added. It will boil over a range of temperatures because as water evaporates, the solution becomes more concentrated, and the boiling point slowly rises.

The boiling point of pure ethanol is  $78^{\circ}\text{C}$ . If it contains any impurities, it will boil at a slightly higher temperature.

From the previous experiments, we can see that impurities raise boiling points of liquids and lower melting points of solids. Therefore boiling points and melting points can be used to test a substance.

Another method that can be used to test purity is chromatography. A coloured substance if pure, will give only one ring or spot on a chromatogram. If it contains impurities (other dyes), it will give several rings or spots; depending on the number of dyes present.

### Applications of impurities

The effect of impurities on melting point has been used in:

- a. Extraction of metals
  1. Extraction of aluminium metal – the ore containing this metal must be melted first. It has a high melting point of about  $2000^{\circ}\text{C}$ . An impurity called cryolite is mixed with the ore before melting. This impurity lowers the melting point to  $850^{\circ}\text{C}$ .
  2. Extraction of sodium – calcium chloride is added to lower the melting point from  $1465^{\circ}\text{C}$  to  $600^{\circ}\text{C}$ .
- b. In countries which experience winter, crude salt is poured on the roads to lower the freezing point of water. This prevents ice from forming on the roads. Water freezes at  $0^{\circ}\text{C}$ . With some impurities like salt, it freezes at a lower temperature.

### Mixture

Many of the materials that we see around us are not pure but are mixtures of a number of substances. A pure substance contains only one kind of matter. There are many ways in which pure substances can mix with each other. The air we breathe is not a single pure substance. It is a mixture of gases. Milk is a mixture which contain droplets of fats and proteins. Spread in water. Muddy water contains soil particles spread in water.

In Chemistry and in everyday life, we often need to separate mixtures obtain pure substances. This is because in some cases only a pure substance, and not a mixture will give us the desired results. For instance, you will prefer to add pure salt to your rather than salt that has sand in it.

### Types of mixture

There are two types of mixtures:

1. Homogenous (uniform) mixture is one in which the particles are evenly distributed into each other for example common salt solution, milk among others.
2. Heterogeneous mixture is one in which the particles are not evenly distributed into each other for example mixture of common salt and sand.

A mixture has at least two parts or components. There are different types of mixtures as shown below.

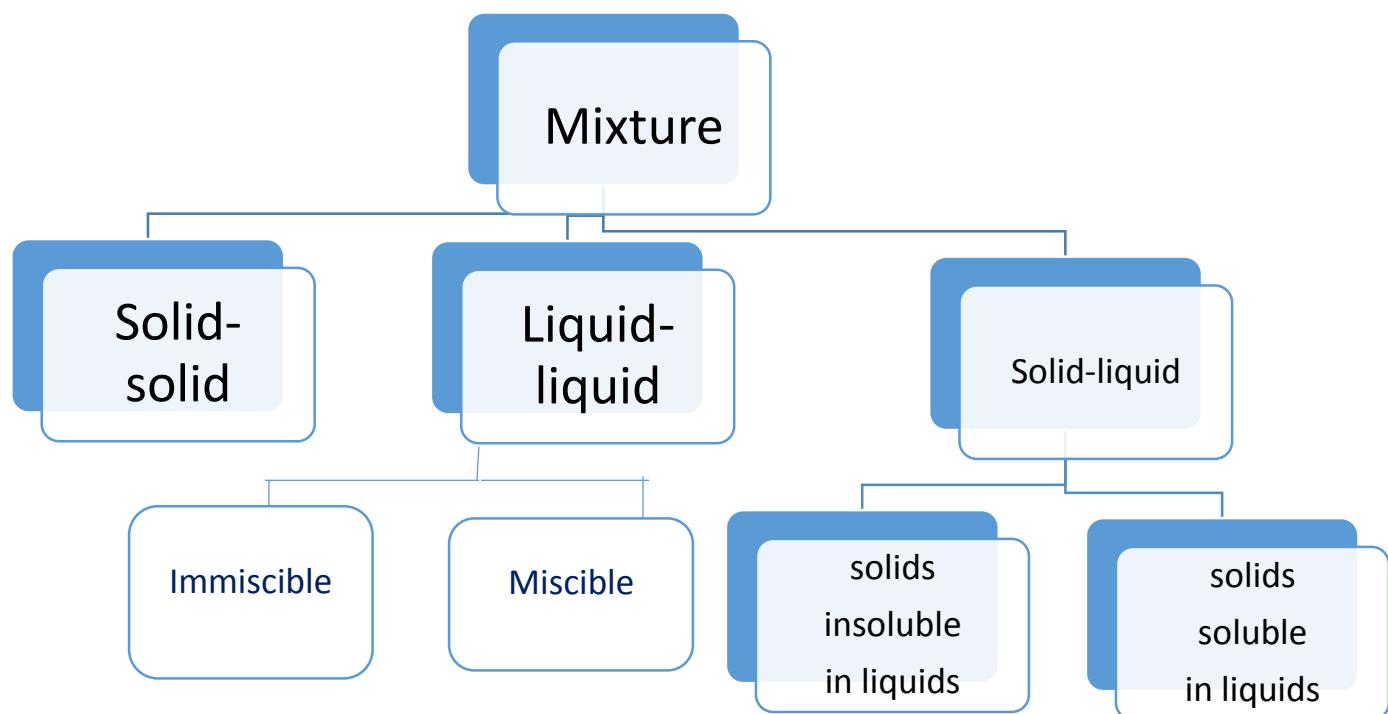


Fig. 3.24 Types of mixtures

(a) Solid – Liquid mixtures

Solid-liquid mixtures can be in solution form or suspension.

(i) Solutions

Experiment 3.10

Aim: To investigation types of solid – liquid mixtures

Apparatus and chemicals

Beakers

- Sugar
- Potassium permanganate

Flour

Stirring rods

Table 3.9 Types of solid - liquid mixtures

No	substance	Water Particles of the solid Have spread in water And cannot be seen	Water particles of the solid have not Spread in water and can b seen
1	Sugar		
2	Sodium Chloride		
3	Flour		
4	Sand		
5	Potassium permanganate		
6	Copper (II) sulphate crystals		

When salt, sugar, potassium permanganate and copper (II) sulphate crystals dissolve on water, the particles of these solids spread throughout the water and

cannot be seen. The mixture formed is called a **solution**. Therefore, a solution is a uniform mixture of solids and liquids. It may be coloured or colourless.

The substance that dissolves is called a **solute**. In the above experiment, salt, sugar, potassium permanganate and copper (II) sulphate crystals are solutes. The liquid in which the solute dissolves is called a **solvent**. Water, in this case, is the solvent. When a substance is dissolved in water the solution is called an **aqueous solution**.

### (ii) Suspensions

In experiment 3.10, we observed that flour does not dissolve in water. Instead, particles of flour spread throughout the water. Some particles eventually settle at the bottom (sediment) of the beaker if it is left to stand for a long time. Particles which have not settled at the bottom form a suspension. They make the mixture cloudy. In muddy water, suspended soil particles make the water appear brown.

Note: do not confuse a suspension with a solution. Particles suspended in liquid can be seen but, in solutions, the particles are invisible.

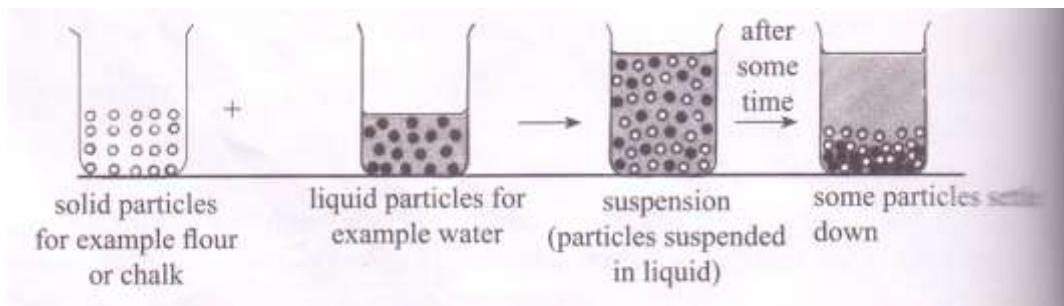


Fig. 3.25 Particles suspended in a liquid

### (b) Liquid – liquid mixtures

Liquid – liquid mixtures can be miscible or immiscible.

#### Experiment 3.11

Aim: To investigate types of liquid-liquid mixtures

Apparatus and chemicals

- Test tubes
- Test tubes rack
- Kerosene

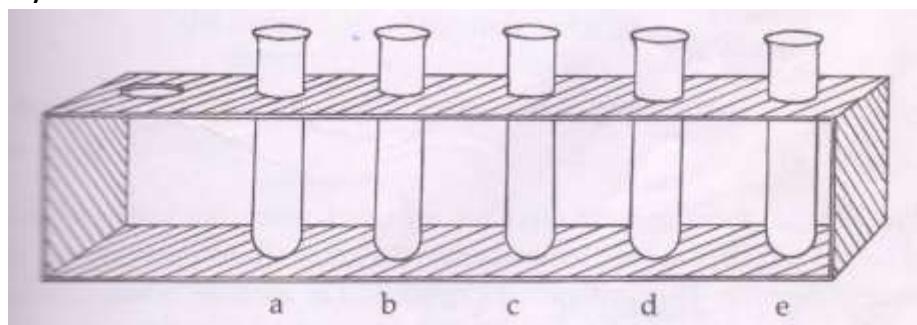
- Oil
- Alcohol (ethanol)
- Water
- Propanone

#### Procedure

Mix a small quantity of each of the following pairs of liquids in a dry test tube and shake well. Place each mixture in a test tube rack as shown in Fig. 3.26.

- (a) Water and ethanol
- (b) Water and oil
- (c) Oil and kerosene
- (d) Ethanol and oil
- (e) Ethanol and propanone

What do you observe?



*Fig. 3.26 Test tubes in a test tube rack*

Copy Table 3.10 in your notebooks and record your observations.

Table 3.10 Types of liquid – liquid mixtures

Mixture	Observation
(a) Water and ethanol	
(b) Water and oil	
(c) oil and kerosene	
(d) ethanol and oil	
(e) ethanol and propanone	

Ethanol mixes completely with water. Kerosene and oil mix completely. In each of the above examples, one is the solute and the other a solvent. In the process of

forming a solution, ethanol and the water particles spread uniformly in each other to form a solution.

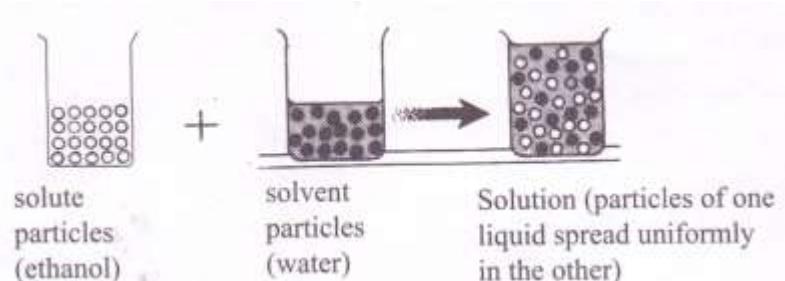


Fig. 3.27 Liquid dissolving in water

Liquids that mix completely to form a solution are said to be miscible. Other examples of miscible liquids are beer, wine, vodka and whisky which are solutions of ethanol in water.

Other liquids do not mix but form layers. Examples include, water and kerosene, water and oil. Such liquids are said to be immiscible.

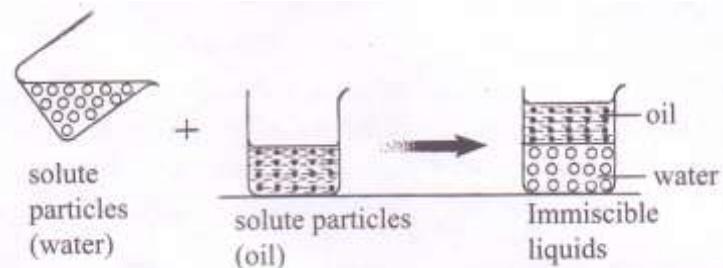


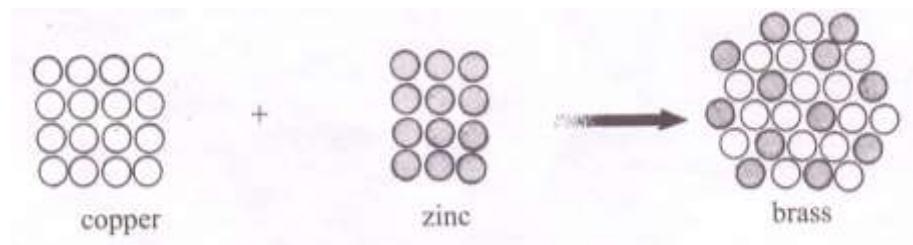
Fig. 28 Immiscible liquids

### (c) solid-solid mixtures

Some solids can also form uniform mixtures with each other. For instance, when some metals are heated together at high temperatures, they melt. Before solidifying, the liquid metals can mix together. This is dissolving one metal in another. When such a mixture solidifies, it forms an alloy which is a uniform mixture of the metals.

Brass is alloy of copper and zinc. Copper and tin form an alloy called bronze.





*Fig. 3.29 Solid – solid mixture*

(d) gas – liquid mixture

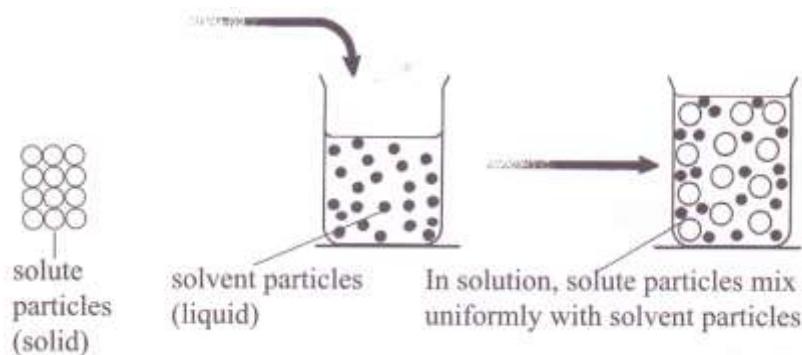
Some common examples are Carbon dioxide dissolved in fizzy drinks for example soda, air (oxygen) dissolved in water is for aquatic animals and plants.

(e) Gas – gas mixtures

Air is a mixture of several gases such as oxygen, nitrogen, carbon dioxide, water vapour and noble gases.

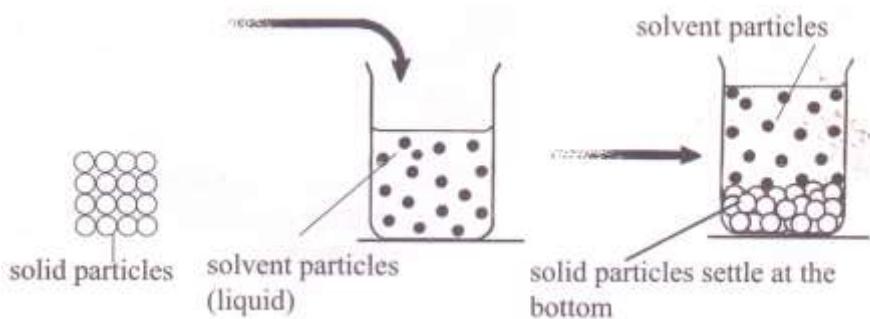
Factors that affect solubility

Water is a common solvent. It is sometimes called universal solvent because it dissolves many substances. Substances that dissolve in a specific liquid are said to be soluble in that particular liquid.



*Fig. 3.30 Dissolving a solid in a solvent*

If they do not dissolve, they are said to be insoluble. Insoluble substances finally settle at the bottom of the liquid.



*Fig. 3.31 Insoluble particles in a liquid*

The extent to which a solid dissolves in a liquid is known as its solubility. Solubility of a solid depends on several factors.

### Experiment 3.12

**Aim:** To investigate the effect of temperature on solubility  
**apparatus and chemical**

- Cold water
- Beaker
- Potassium nitrate

#### Procedure

1. Add about  $100\text{ cm}^3$  of cold water in a beaker.
2. Add potassium nitrate a little at a time and stir.
3. Repeat until no more salt dissolves.
4. Place the beaker with the solution on tripod and heater while stirring. What can you observe?

When you added more salt and it dissolved, it showed the solution was unsaturated. The water dissolved more solids because it did not have enough salt particles. A solution which has dissolved enough solute in the presence of some undissolved solute at a given temperature is called a saturated solution. If a saturated solution is warmed, it usually dissolves more solute, that is, it becomes unsaturated. In other words, some salts are more soluble in hot water or other solvents than they are in the cold solvents.

The amount of substance which dissolves in a specified amount of solvent, for example  $100\text{ cm}^3$  water to form a saturated solution (at any given temperature) is known as its solubility.

Note: Solubility of a solid can be increased by raising the temperature or adding more solvent.

### Methods of separating mixtures

Various methods can be used to separate mixtures of two or more substances. These methods depend on the physical properties of the substances which make up the mixture. These include solubility, density, boiling point, miscibility and others.

#### Separating an insoluble solid from a liquid

An insoluble solid can be separated from a liquid by three separation techniques:

- (a) Decantation
- (b) Filtration
- (c) Evaporation

- (a) Decantation

#### Experiment 3.13

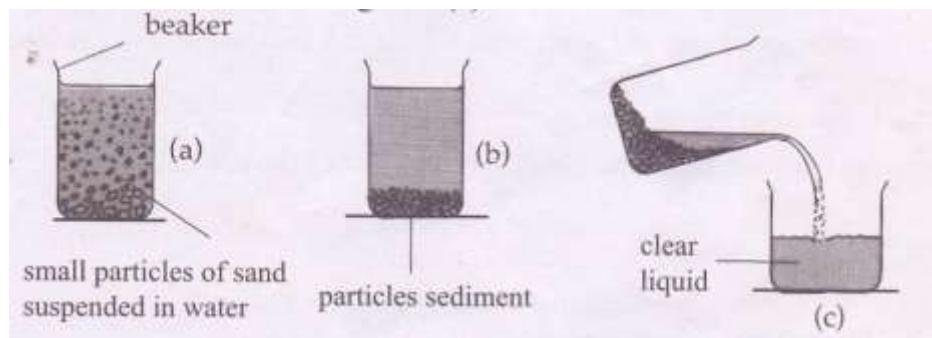
Aim: To separate sand from a mixture of sand and water

#### Apparatus and chemicals

- Beaker
- Sand
- Water

#### Procedure

1. Place some sand in a beaker
2. Add water and stir
3. Allow the sand to settle to the bottom of the beaker.
4. Pour off water as shown in Fig. 3.32 (c) #



*Fig. 3.32 Decantation*

When the insoluble sand particles settle to the bottom of the beaker (sediment), the water can simply be poured off. Sand is left in the beaker.

This process is called decantation. However, in some cases, we cannot obtain all the water from the mixture by decanting.

### (b) Filtration

#### Experiment 3.14

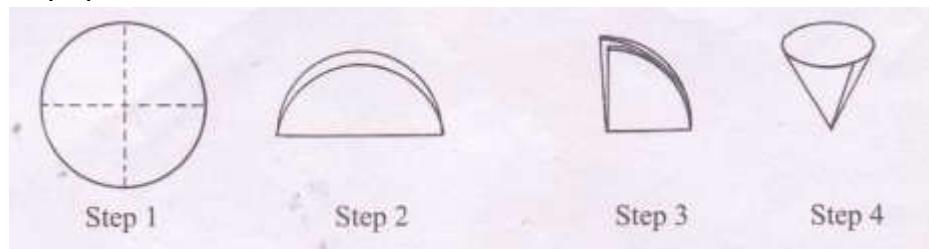
Aim: To separate soil from a mixture of soil and water

Apparatus and chemicals

- Beaker
- Filter paper
- Soil
- Conical flask
- Filter funnel
- Water mixture

Procedure

1. Fold a circular filter paper into two form a semi-circle, then again to form a quadrant.
2. Open the paper into a hollow cone

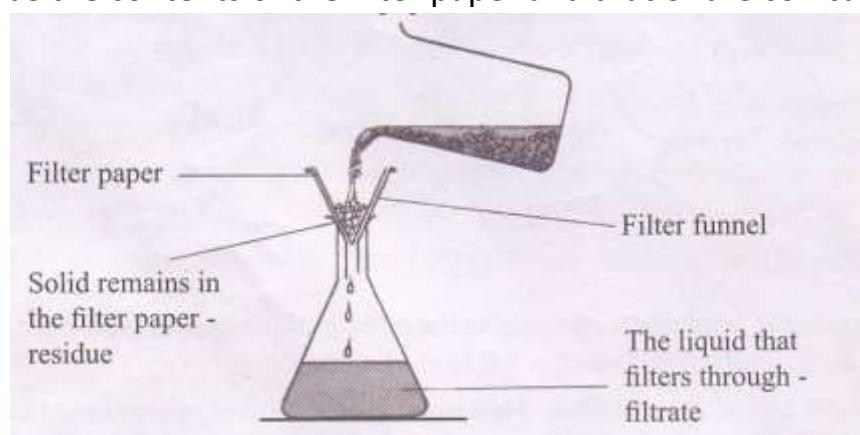


*Fig. 3.33 making a cone*

3. Wet the paper cone with a little water so that it sticks to the funnel.
4. Place it inside a filter funnel. Place the funnel on a conical flask as shown in fig 3.34
5. Stir the mixture of soil and water with a glass rod and pour it into the filter paper cone.

What happens to the mixture?

Describe the contents of the filter paper and that of the conical flask.



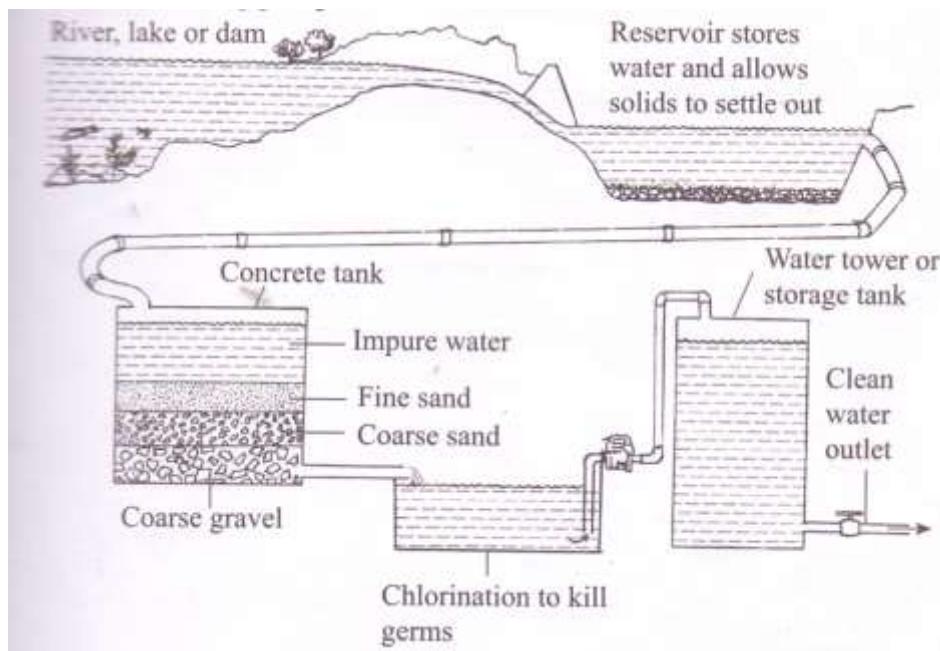
*Fig. 3.34 separation of soil from water by filtration.*

The heavy soil particles will remain in the beaker.

Filtration is generally a more useful method than decantation for separating insoluble solids from liquids. Infiltration, both the solid and liquid can be obtained in one process. We use a porous type of paper called filter paper, which acts like a fine sieve, allowing the liquid to pass through but retaining solid particles. The liquid which passes through is called the filtrate and the solid that remains or the filter paper is called the residue.

#### Application of filtration in purifying water

Filtration is important in the process of obtaining clean water for use in our homes. In this process from a source if first stored in reservoirs where most of the solid particles settle out. It is then filtered through sand and gravel which trap smaller particles of mud and other suspended solids (see Fig. 3.36). The filtered water is treated with chlorine to kill germs before being pumped for use in homes and industries.



*Fig. 3.35 main stages in the purification of water*

List other processes at home that involves filtration.

### (C) Evaporation

#### Experiment 3.15

Aim: To separate salt from sand

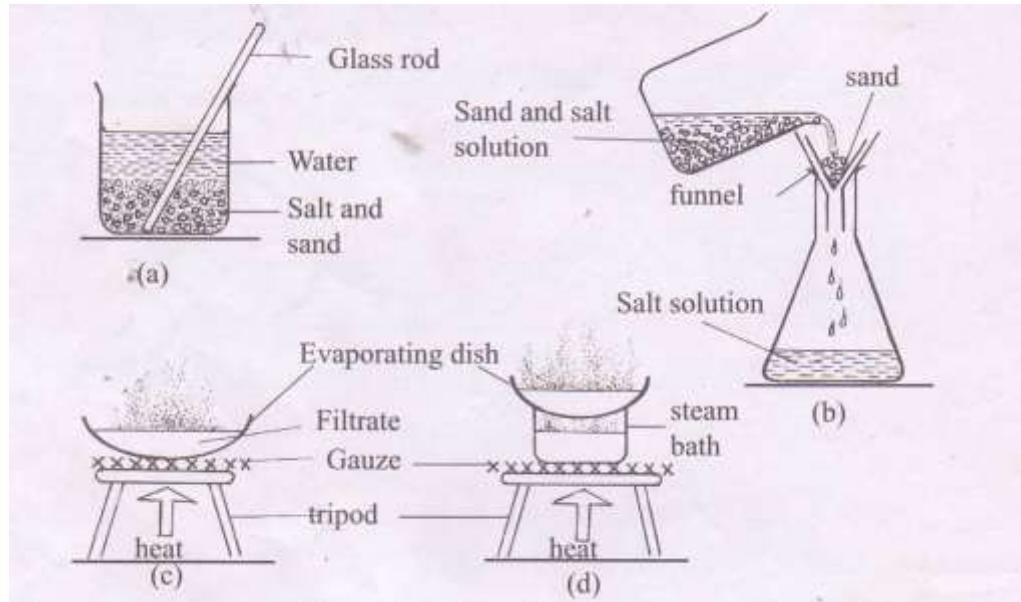
Apparatus and chemicals

- Beaker
- Filter funnel
- Glass rod
- Tripod stand
- Conical flask
- Filter paper
- Wire gauze
- Evaporating dish

#### Procedure

1. Place the mixture of salt and sand in a beaker
2. Fill the beaker half way with water
3. Stir with a glass rod to make the salt to dissolve faster (see 3.35)

4. Pour the mixture through a filter funnel fitted with a filter
5. Put the filtrate into an evaporating dish
6. Heat the filtrate gently and evaporate to dryness as shown in (c) or (d) what is the name of the white solid in the evaporation dish?



*Fig. 3.36 steps of obtaining salt from sand*

Caution: Do not taste unless instructed to do so by your teacher. Remember one of the safety rules – you do not taste anything in the laboratory unless asked to do so.

This experiment has employed three techniques:

- (a) Dissolving
- (b) Filtration
- (c) Evaporation

Sand was left on the filter as the residue. After evaporating all the water from the salt solution (filtrate), salt was left in the evaporating dish.

The gaseous form of a substance at room temperature is called vapour. The process of changing a liquid to gas or vapour is called evaporation. Evaporation can be used to separate a dissolved solid (solute) from its solution.

Evaporation is an essential process in drying wet clothes. During evaporation, only the solute is recovered while the solvent is let to escape as vapour. Evaporation takes place at any temperature.

## Separation of immiscible liquids separating funnel

This can be done by either simple or fractional distillation.

### Experiment 3.16

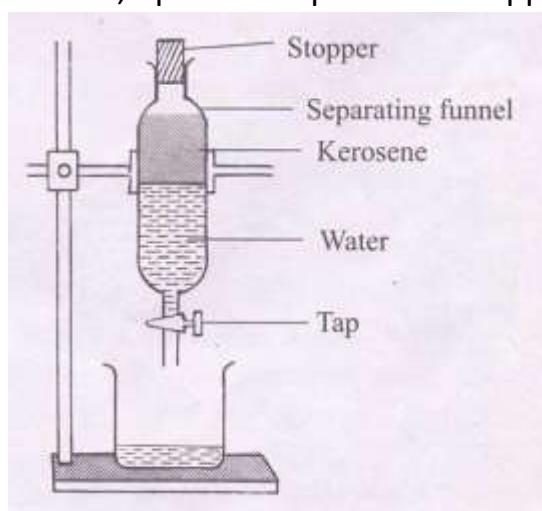
Aim: To separate kerosene and water using a separating funnel

Apparatus and chemicals

- Separating funnel
- Kerosene
- Two beakers
- Water

Procedure

1. Place the mixture of kerosene and water in a separating funnel.
2. Leave the mixture for some time until there is a clear dividing line between the two liquids.
3. Remove the stopper and open the tap to let the lower layer run out slowly into a beaker. Close the tap when the upper layer reaches the tap.
4. Run out a small quantity of the remaining liquid into a separate container and discard it.
5. Using a different beaker, open the tap to let the upper layer run into it.



*Fig. 3.37 separating two immiscible liquids using a separating funnel*

- Why should a small portion of the liquid between the first collection and the second one be discarded?

The discarded portion was a mixture of kerosene and water. Two liquids that do not mix, such as kerosene and water, are said to be immiscible. Their separation is based on differences in their densities. The less dense liquid (kerosene) floats on the denser liquid (water). The device used is called a separating funnel.

### Separation of soluble solid-liquid mixtures

This can be done using either simple distillation or fractional distillation.

#### (a) Simple distillation

#### Experiment 3.17

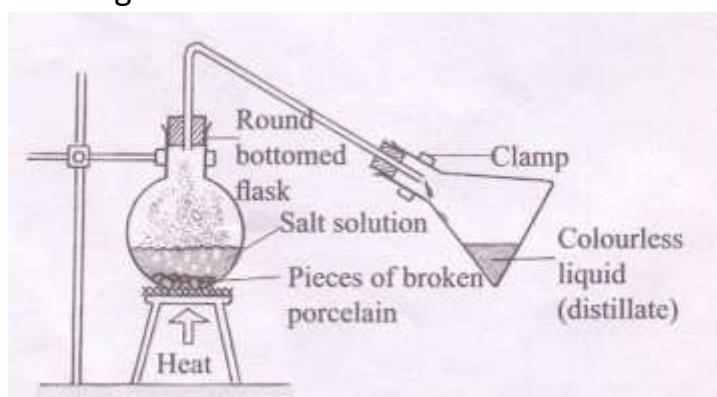
Aim: To obtain both salt and water from a salt solution using simple distillation

#### Apparatus and chemicals

- Distillation flask
- Bunsen burner
- Thermometer
- Sodium chloride (common salt)
- Round bottomed flask
- Stand and clamp
- Beaker or conical flask
- Water

#### Procedure

1. Prepare some salt solution.
2. First try to separate the components by filtration method. Does it work?
3. Pour the solution into the first and add pieces of broken porcelain. Arrange the apparatus as in Fig. 3.3.8. Boil the solution.

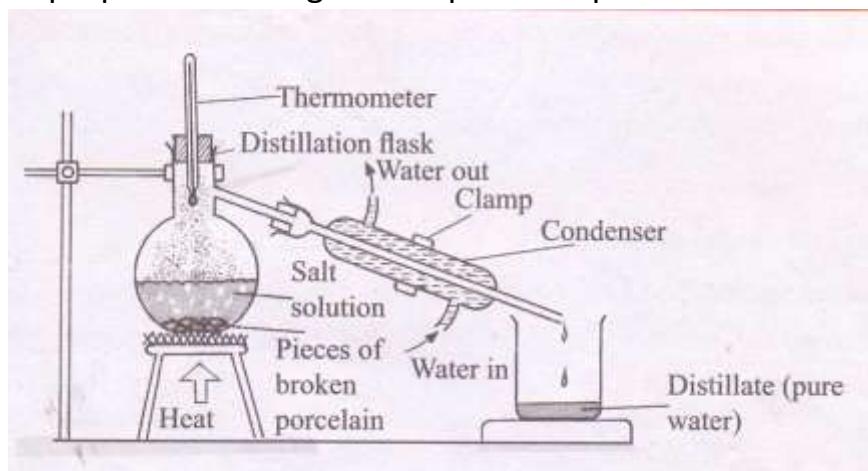


*Fig. 3.38 simple distillation of salt solution*

The residue left in the flask is salt.

The salt solution can be separated as shown in Fig. 3.39.

- What is the purpose of adding broken pieces of porcelain?



*Fig. 3.39 simple distillation of salt solution*

- What is the use of the thermometer in Fig. 3.39?
- Which of the two set-ups provides a better way of cooling the vapour? Explain
- Look at Fig. 3.39 again. What do you think would happen if the direction of cooling water is reversed? Give a reason for your answer.

The pieces of broken porcelain in the flask prevent “bumping” of the solution during boiling. Glass beads or pieces of porous pot can also be used to achieve smooth boiling.

The thermometer is used for noting the temperature at which the solution boils. The set up in Fig. 3.39 is more efficient in cooling the vapour. This is because of the cool water circulating in the condenser.

If the direction of cooling water is reversed, the water circulating in the condenser would flow down through the condenser a bit warm. Therefore it would not cool the vapour as efficiently as before.

During simple distillation, evaporation and condensation take place at the same time but at different parts of the apparatus. The liquid is evaporated in the

distillation flask. It is condensed by passing it through a condenser, which is cooled by water. The condensed liquid is transferred into a container usually referred to as a receiver. The pure liquid collected is called the distillate. The process is called simple distillation.

Note simple distillation can also be used to separate miscible liquid mixtures whose components have a difference in their boiling points of 40°C and above.

In areas where the available water is salty, like in coastal regions, pure water can be obtained by distilling the salty water.

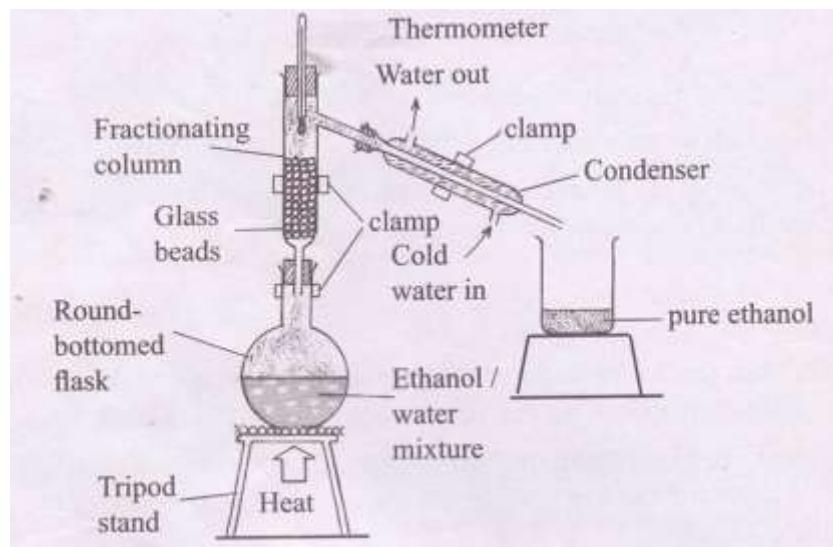
(b) Fractional distillation

Experiment 3.18

Aim: To separate a mixture of ethanol and water by fractional distillation

Apparatus and chemicals

- Round-bottomed flask
- Bunsen burner
- Fractional column
- Ethanol
- Condenser
- Thermometer
- Receive (beaker)
- Water



*Fig. 3.40 separating a mixture of ethanol and water by fractional distillation*

#### Procedure

1. Set up apparatus as shown in fig. 3.40 to separate a mixture of ethanol and water.
2. The mixture is heated. At about  $78^{\circ}\text{C}$  ethanol whose boiling point is lower, (more volatile) begins to boil. Some water evaporates too, so a mixture of ethanol vapour and water vapour rises up the fractionating column.
3. Continue heating and collect the liquid that drips between  $80^{\circ}\text{C}$  and  $100^{\circ}\text{C}$  in a different receiver and discard it.  
Why should we discard this portion?
4. When the thermometer reading eventually rises to  $100^{\circ}\text{C}$ , the water is boiling and the liquid collected now is water collect it in another receiver.
5. Perform the following tests on each of the two distillates:
  - (a) Smell each distillate.
  - (b) Place a little of each on a watch glass and test with a lighted splint with a lighted splint or matchstick. Record your observations.

The fractional column is packed with glass beads to increase the surface area for condensation. When the vapours rise, they condense on the glass beads in the column, which become hot.

When temperature of the beads reaches about  $78^{\circ}\text{C}$ , ethanol vapour no longer condenses on them. Only the water vapour condenses. The water drips back to

the lower parts of the column or into the distillation flask, while the ethanol vapour enters the condenser.

Ethanol condenses. Liquid ethanol drips into the receiver as the first distillate.

The first distillate is flammable (burns) and has a characteristic smell. This is ethanol (alcohol). The second distillate is not flammable and has no smell. It is water. We have, therefore, separated the mixture into pure ethanol and pure water. The portion discarded is a mixture of ethanol and water.

Miscible liquids can be separated by fractional distillation especially if they have different, but close boiling points.

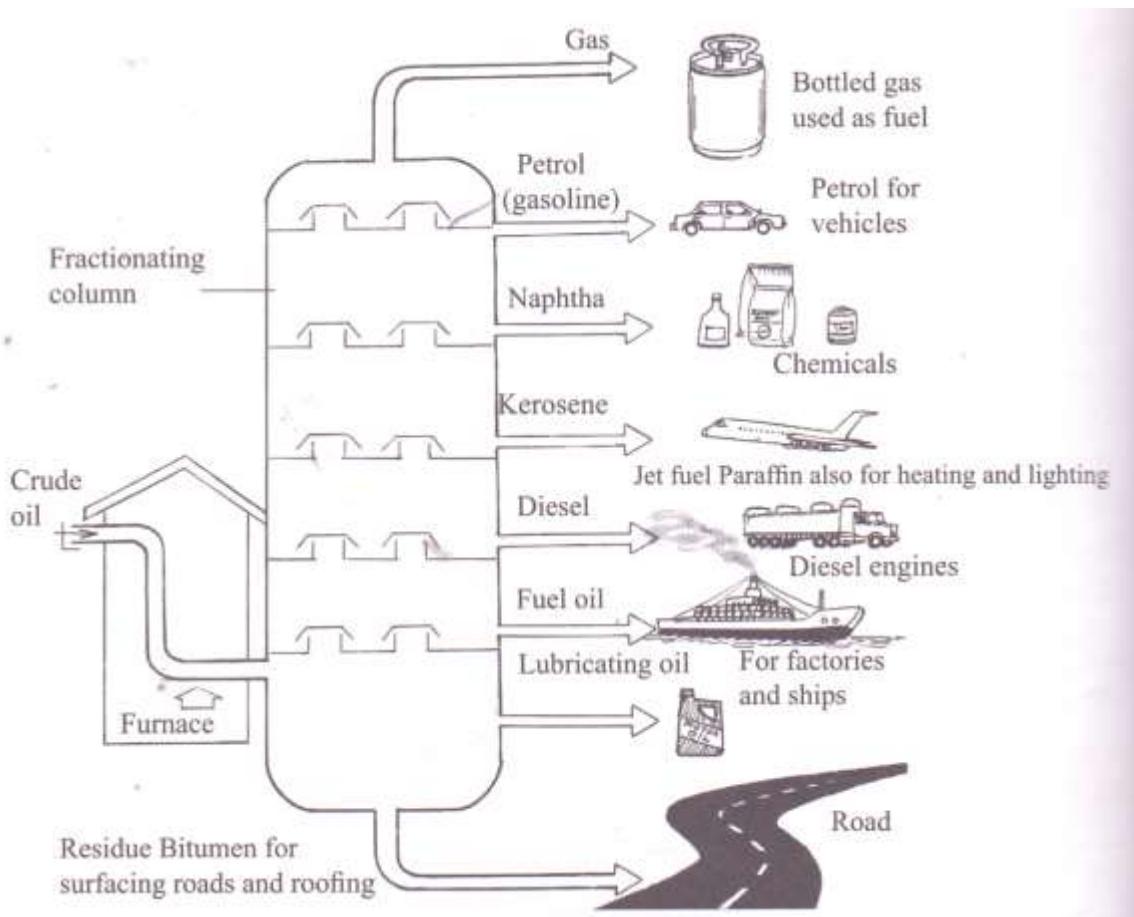
Fractional distillation separates mixture into two or more components (or fractions).

Industrial application of fractional distillation

Fractional distillation is used industrially or on a large scale to:

- Separate various fractions from crude oil.
- Manufacture oxygen and nitrogen from air, for instance at the British Oxygen Company (BOC).
- Manufacture spirits such as gin and whisky.

An example of fractional distillation is separation of crude oil in a refinery as shown in Fig. 3.14



*Fig. 3.41 fractional distillation of crude oil in a refinery*

### Separation of solid-solid mixture

Solid-solid mixture can be separated using either sublimation or by use of a magnet (magnetism)

#### (a) Sublimation

We have seen that some solids generally melt when heated. However, a few substances change directly from solid to gas on heating, and from gas to solid on cooling, without passing the intermediate liquid state. This change is called sublimation.



Substances that sublime include; iodine ammonium chloride and solid carbon dioxide (dry ice). Sublimation can be used to separate a mixture of two solids as long as one of the solids sublimes on heating.

### Experiment 3.19

Aim: To demonstrate sublimation

Apparatus and chemicals

- Boiling tube
- Bunsen burner
- Flakes of iodine
- Test tube holder or a strip of paper
- Ammonium chloride

Procedure

1. Place a little ammonium chloride in a boiling tube.
2. Hold the tube with a test tube holder or a folded strip of paper
3. Heat gently over the non-luminous Bunsen flame and observe what happens.
4. Repeat the experiment with one flake of iodine and record your observations.

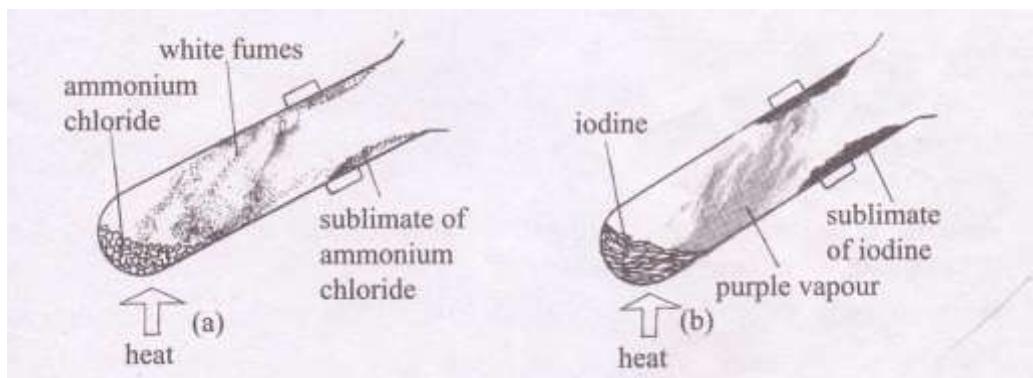


Fig. 3.42 sublimation of ammonium chloride and of iodine

Ammonium chloride does not melt. It turns into white fumes which condense back to a white solid on reaching the cooler parts of the boiling tube. The white solid is called a sublimate. Solid iodine forms purple vapour which condenses on the cooler parts of the boiling tube to form a sublimate of pure iodine (see Fig. 3.42(b)).

## Experiment 3.20

Aim: To separate a mixture of sodium chloride and ammonium chloride

Apparatus and chemicals

- Bunsen burner
- Sodium chloride (common salt)
- Funnel
- Ammonium chloride

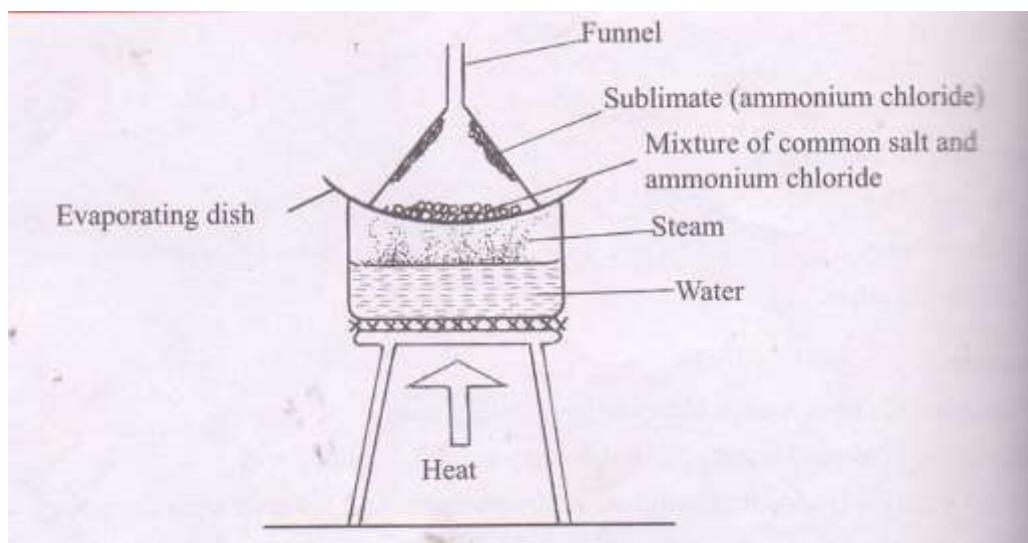


Fig. 3.43 separation of sodium chloride and ammonium chloride

Procedure

1. The apparatus is arranged as in Fig. 3.43.
2. Mix equal amounts of common salt and ammonium chloride (both salts are white and soluble in water).
3. Place the mixture in an evaporating dish and cover with an inverted glass funnel. Why should we not use a plastic funnel?
4. Heat the water to produce steam which heats the mixture. Observe what happens
5. Repeat the experiment with a mixture of iodine and sodium chloride.

Ammonium chloride sublimes to form white fumes which cool to form a white solid on the inside surface of the funnel. Sodium chloride is left in the evaporating dish. Iodine also sublimes leaving the solid white sodium chloride behind.

Note: we can also use a magnet to separate a mixture if it contains a magnetic substance. For example, in a mixture of iron and sulphur, iron can be attracted by a magnet.

### (b) Magnetism

It is a separation technique that involves the use of magnet property. In industries an electromagnet is used to separate iron from other materials in the crushed ore. The electromagnet attracts iron a magnetic substance from non-magnetic materials thereby achieving separation.

#### Experiment 3.21

Aim: To separate a mixture of iron and sulphur

Apparatus and chemicals

- Iron filings
- Beaker
- Sulphur powder
- Magnet

Procedure

1. Mix iron filings and sulphur powder in a small beaker
2. Put the mixture on a white paper
3. Pass a magnet (separated from the mixture with another white paper as shown in Fig. 3.44) over the mixture.

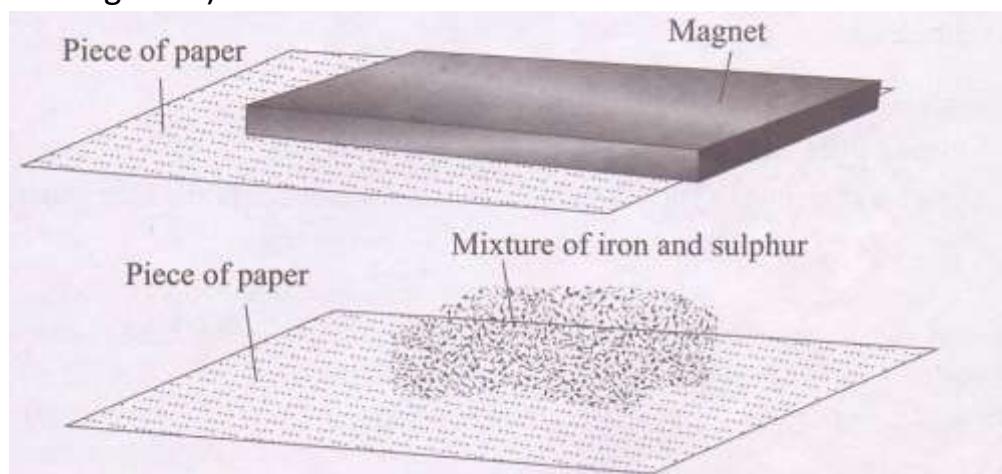


Fig. 3.44 separation of mixtures by magnetism

The iron filings are attracted by the magnet thereby separating them from sulphur powder.

### Chromatography

The name of this technique comes from the Greek word chromes, which means colour. This is because the early work was done with coloured substances.

Chromatography is a technique for the separation of components of a solution by passing it through a medium. The components move at different rates over an absorbent material such as filter paper. In paper chromatography, the absorbent material is the filter paper and different solvents, such as water, ethanol, propanone and other organic solvents can be used.

Are you writing with blank ink? Is it a one-colour substance?

Look outside! Do you see those trees, that grass, those flowers? What colour do you see on each of them?

What solvent can we use to extract these colours? Water? No. the rain would have long washed them out. Think of other solvents that we can use.

### Paper chromatography

#### Experiment 3.22

Aim: To investigate whether black ink is one single colour

#### Apparatus and chemicals

- Teat pipette
- A beaker
- Black ink
- Filter paper
- Water

#### Procedure

1. Place a filter paper on a beaker
2. Use a teat pipette to put a drop of blank ink at the centre of the filter paper.

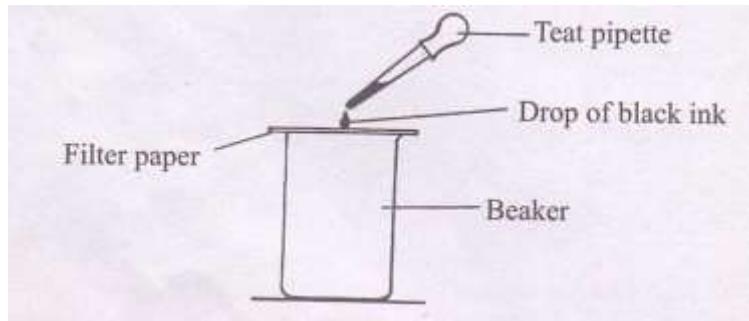


Fig. 3.45 paper chromatography

3. Wait for the drop of ink to spread and dry
4. Using a clean pipette, carefully add a drop of water to the ink spot.
5. Wait until the water has spread outwards and then add another drop.
6. Wait patiently until the water stops spreading.
7. Keep on adding drops of water until there is no further change.
8. Make a list of the different colours observed on the filter paper.
9. Finally, let your paper dry. Compare your result with those of the other groups.

You must have observed a number of different dyes. This indicates that the black ink is made of several dyes. The different dyes in the ink dissolved in water and spread out with the water to form a series of coloured rings, called a chromatogram (see Fig. 3.46).

Two factors are necessary for the chromatogram to form:

- a. The dyes must be soluble in the solvent used
- b. The dyes must be retained by the absorbent material.

The degree of solubility in water is different for the dyes found in ink. The filter paper also retains the dyes at different rates. This is the reason why they move at different rates and are finally separated. The dyes immediately after the solvent front is the most soluble. It is also the least retained.

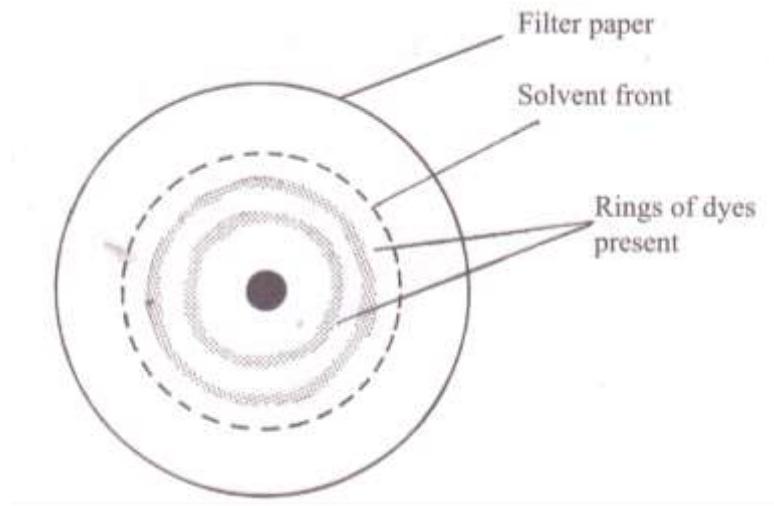


Fig. 3.46 a chromatogram of black ink

The green dye found in the grass and various coloured dyes in flowers can be separated by a similar process. You could also use ascending chromatography as described in the next experiment. This method is called ascending paper chromatography because the solvent travels upwards.

### Experiment 3.23

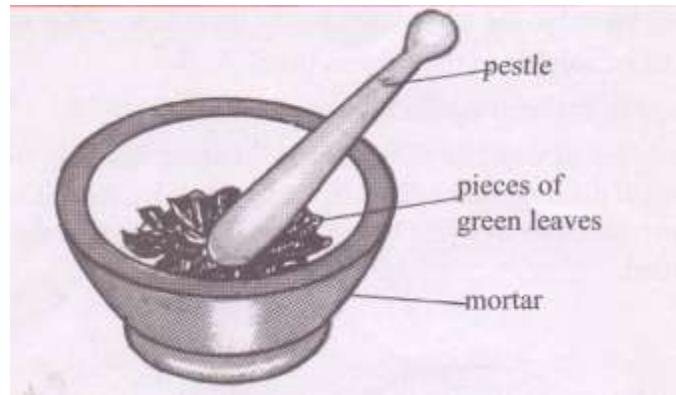
Aim: To separate the components of chlorophyll in leaves

Apparatus and chemicals

- Mortar
- Boiling tube
- Pestle
- Methylated spirit (or any other suitable solvent)

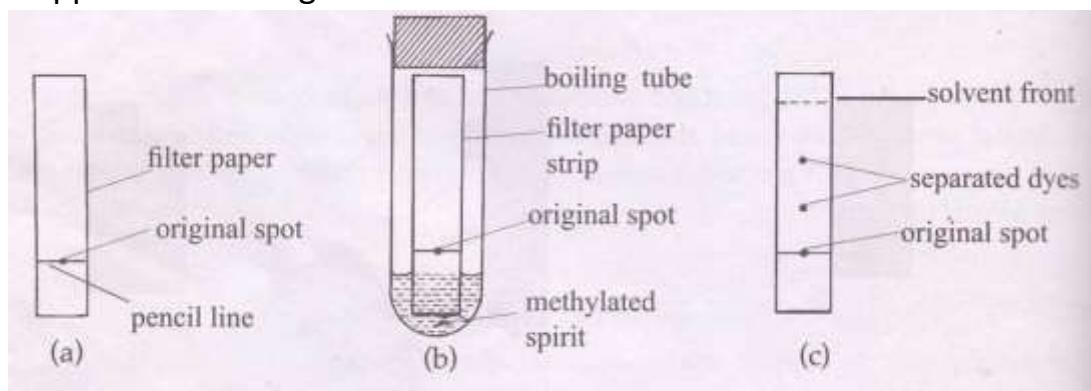
Procedure

1. Crush a handful of green leaves using a pestle and mortar. Add 3-4cm<sup>3</sup> of methylated spirit. Continue crushing until you get a green extract.



*Fig. 3.47 crushing green leaves in mortar*

2. Make a strip of filter paper  $15 \times 2$  cm. draw a pencil line on the paper about 2.5 cm from one end. Make a pencil mark in the middle of this line.
3. Using a teat pipette, place a drop of the green extract solution on the mark on the pencil line. Wait for the spot to dry. Add another drop on exactly the same place and wait for it to dry.
4. Place the paper strip in a boiling tube containing methylated spirit. Make sure the surface of spirit is about 1.5 cm below the pencil line.
5. Set apparatus as in Fig. 3.48 below and leave it for a few hours.



*Fig. 3.48 ascending paper chromatography*

When the solvent moves up, it dissolves the substances in chlorophyll and then takes with it along at different rates. The most soluble and least retained dye is just behind the solvent front. Any insoluble substance remains at the original spot.

6. Repeat the experiment with any other types of ink or coloured substances.

Compare your results with those obtained in the previous experiment  
Application of chromatography

The main use of chromatography is separating the dyes, for example, the dyes that make black ink. It is also used to analyze and identify mixtures of substances which are difficult to separate by other means. We can also use it to analyze dyes used in food colouring.

### Solvent extraction

Aim: To extract oil from nut seeds

### Apparatus and chemicals

- Pestle and mortar
- Nut seeds
- Propanone
- A white paper

### Procedure

1. Crush nut seeds in a mortar using a pestle
2. Add a suitable solvent such as propanone and continue crushing.
3. Pour the resulting solution into an evaporating dish
4. Vaporize the solvent using energy from the sun.  
What is left on the evaporating dish?
5. Put a piece of plain paper in the liquid left in the evaporating dish. Record your observations.

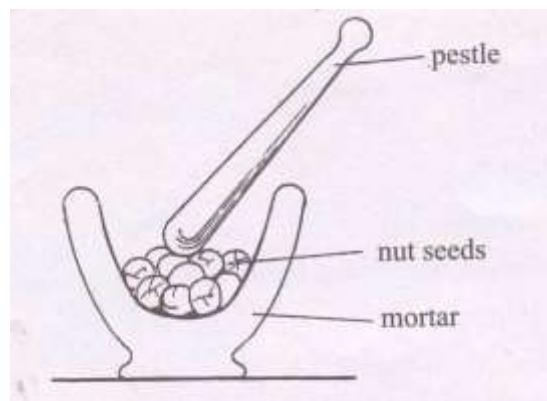


Fig. 3.49 extracting oil from nuts

Oil from the nut seeds dissolve in propanone. The solvent evaporates in the sun leaving oil. This method of extracting substances using suitable solvents is called solvent extraction.

The paper is used to test for oil. It becomes translucent after being dipped in oil. Propanone has a lower boiling point than oil and therefore vaporizes when left in the sun, leaving oil (high boiling point).

### Crystallization

Crystallization involves use of a saturated solution. A saturated solution is one that cannot dissolve any more of the solute at a given temperature and contains undissolved solute. We can make crystals from a saturated solution.

#### Experiment 3.25

Aim: To prepare a saturated solution

Apparatus and chemicals

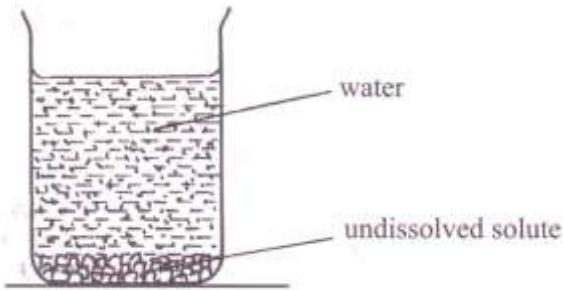
- Beaker
- Stirring rod
- Sodium chloride
- Water

#### Procedure

1. Put  $100\text{ cm}^3$  of water in a beaker
2. Add sodium chloride a little at a time while stirring well with a glass rod until no more can dissolve. One can tell it is saturated when undissolved salt settles down even after vigorous stirring.
3. Filter off the undissolved salt.

The filtrate is called a saturated solution.

When saturated solutions loss more water (solvent) a solid is left behind. The solid has a regular shape. Such a solid is called crystal. The process of obtaining crystals from a solution is called crystallization.



*Fig. 3.50 crystallization*

Note: Hot water will dissolve more sodium chloride than cold water.

### Experiment 3.26

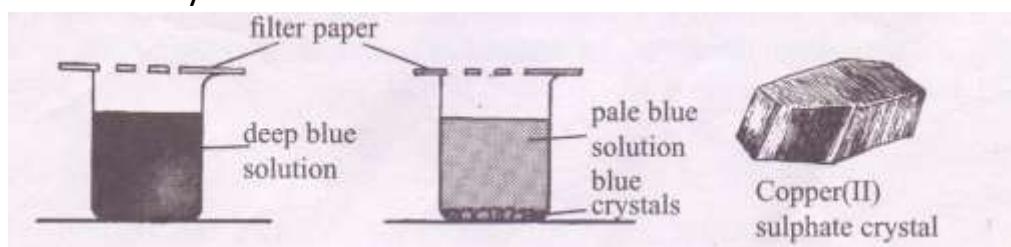
Aim: To prepare big crystals of copper (II) sulphate

Apparatus and chemicals

- Beaker
- Water
- Copper (II) sulphate
- Glass rod

Procedure

1. Prepare about  $50\text{ cm}^3$  of a saturated solution of copper (II) sulphate solution using the method in Experiment 3.25.
2. Put the filtrate in a beaker and cover it with a filter paper pierced with a few holes.
3. Leave the content of the beaker undisturbed for 2-3 weeks. After this time very beautiful crystals are obtained.



*Fig. 3.51 preparation of Copper (II) sulphate crystal*

Application of crystallization

This process of natural evaporation is used to separate trona and sodium chloride. Trona is double salt of sodium carbonate and sodium hydrogen carbonate. Sodium chloride is found in the water trapped between trona crystals.

Most of the sodium chloride for our consumption is obtained by evaporation of seawater and other salt works near the coast region.

#### Difference between mixtures and compounds

We have learnt about the different methods of separating mixtures. A mixture is formed as a result of bringing particles of different substances into close contact with each other, without chemically combining them. Usually the substances making up a mixture can be mixed in any proportions and each component retains its original physical and chemical properties.

A mixture, therefore, shows the properties of its components. For instance, a mixture of sand and salt has the properties of both salt and sand. Similarly, sugar solution (a sugar and water mixture) has the properties of both sugar and water.

A mixture may contain elements or compounds. Air is an example of a mixture. Air consists of a mixture of elements (nitrogen, oxygen and noble gases); and compounds (water vapour and carbon dioxide). We have already seen that mixture can be separated by simple means such as filtration, chromatography, crystallization, distillation (simple and fractional) and sublimation.

Which other methods of separation of mixture do you know?

We saw earlier in this unit that a compound also consists of more than one type of substance. However, compounds are different from mixtures.

From Experience 3.7, we learnt that iron filings and sulphur can be separated by physical means, using a magnet. This is possible because the two elements were not chemically combined, but were not mixed.

We can conclude from these observations that different kinds of matter can be made to combine together in two ways to form complex substances. They can merely be brought together in any proportion to form a mixture or they can be heated or allowed to react chemically to form a compound. A compound contains two or more different atoms chemically joined together.

Table 3.11 gives a summary of the main difference between a compound and a mixture. Study and understand the difference. Name two substances that can be distinguished by these differences.

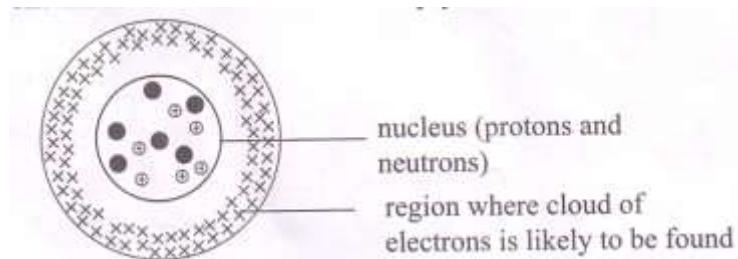
Mixture	Compound
<ol style="list-style-type: none"><li>1. The substances in a mixture can be separated by physical methods.</li><li>2. The properties of a mixture are the average of the properties of the substances in it.</li><li>3. The substances which make the mixture need not be in fixed proportions.</li><li>4. Mixtures are formed by physical methods; no new substance is formed</li><li>5. No energy (heat and light) is liberated out or absorbed during the formation of the mixture.</li></ol>	<ol style="list-style-type: none"><li>1. The substances in a compound cannot be separated by physical methods.</li><li>2. The properties are different from the properties of the substances which made it.</li><li>3. The substances in a compound are in fixed proportions.</li><li>4. A compound is formed by a chemical method; a new substance is formed.</li><li>5. Energy (heat and light) is liberated or absorbed during the formation of the compound.</li></ol>

## Characteristics of particles that make up an atom

Atoms are very small particles of an element, they contain even smaller particles called sub-atomic particles. These are three kinds of sub-atomic particles. These are the protons, electrons and neutrons. Protons and neutrons are found in the nucleus. The nucleus is the middle part of the atom. It is very dense and extremely small part of the atom. Outside the nucleus is a much larger region of the atom where electrons occur in energy levels at different distances from the nucleus of the atom. Electrons can be imagined as circulating the nucleus, at great speeds in those energy levels.

Note: An electron is represented by a dot ( $\bullet$ ) or cross (x).

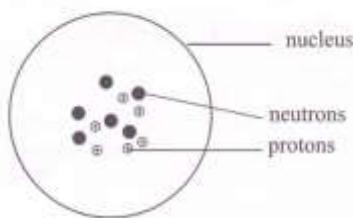
We can represent the structure of an atom simply as follows:



*Fig. 4.1 an atom showing nucleus and a cloud of electrons*

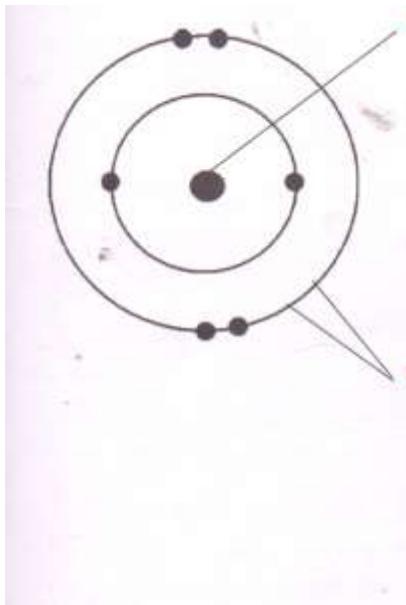
Electrons are negatively charged (-ve), protons are positively charged (+ve) and neutrons have no charge. Therefore, the nucleus has a positive charge because of the protons as illustrated in Fig. 4.1.

If we magnify the nucleus, we can show the sub-atomic particles (protons and neutrons) as in Fig. 4.2



*Fig. 4.2 a nucleus containing protons and neutrons*

Since protons have similar charge (+ve) they repel each other. However, the repulsion is minimised by the presence of the neutrons as illustrated in Fig. 4.3



### The nucleus

- It is at the **centre** of the atom.
- It contains **protons** and **neutrons**.
- It has a **positive charge** because of the **protons**.
- The **whole mass** of the atom is **concentrated in the nucleus**.
- However, it is **tiny** compared to the size of the atom.

### The electrons

- They move **around** the nucleus.
- They are **negatively charged**.
- They are **tiny**, but cover a lot of space.
- They occupy the **energy levels**.

*Fig. 4.3 atomic structure showing protons and neutrons in the nucleus and electrons in the energy level*

In a neutral atom: Number of protons = Number of electrons

This means that each proton (+ve) has an electron (-ve) as a partner.

Table 4.1 particles present in an atom

Sub-atomic particle	Mass	Charge	Where found in atom
Proton	1	+ (positive)	Inside the nucleus
Neutron	1	0 (neutral)	Inside the nucleus
Electron	$\frac{1}{1840}$	-(negative)	Outside the nucleus

### Arrangements of elements in shells of an atom

To understand the atomic structure, let us look at a suggested analogy of atomic 'city'. In a well planned town or city, houses are constructed according to the plan. The house rent also varies. Those people who live in some neighbourhoods of the city are wealthy people and therefore pay a lot more than those who live in other areas.

### Peculiarity of the 'atomic city'

These are found in a country where people live according to the following city plan.

1. There are three suggested types of houses that are allowed in each neighbourhood. Only 2 people are allowed to occupy a single house.
  - 1<sup>st</sup> neighbourhood – has one low cost stone house only
  - 2<sup>nd</sup> neighbourhood – has four middle cost stone houses only
  - 3<sup>rd</sup> neighbourhood – has one high cost stone house only
2. All houses must be built away from the city/town centre.  
How many people are in:
  - a) Low cost houses in the 1<sup>st</sup> neighbourhood?
  - b) Middle cost houses in the 2<sup>nd</sup> neighbourhood?
  - c) What is the maximum number of high cost houses expected to be in the 3<sup>rd</sup> neighbourhood?

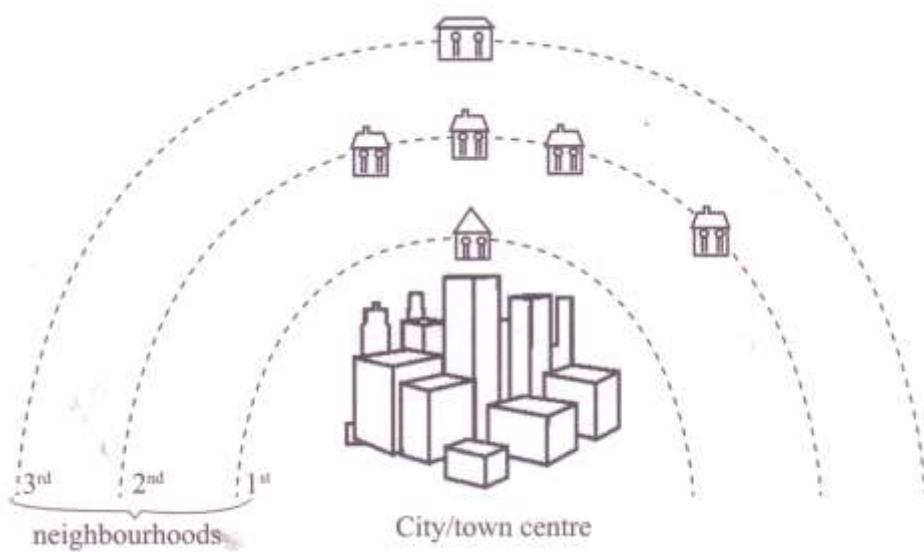
We notice that the further the houses are from the city centre, the more expensive they are. The cheapest houses are in the first neighbourhood. So the people who live in the 3<sup>rd</sup> neighbourhood are very wealthy. We mentioned that the houses are settled according to wealth and ability to pay higher rent. This implies that the 3<sup>rd</sup> neighbourhood has wealthy people. The 2<sup>nd</sup> has poor people and the 1<sup>st</sup> has the poorest people. In other words the house rent increases as we move from the city centre.

The city plan (vocabulary list) is as follows:

City	= Atom
People	= Electrons
Neighbourhoods	= Energy levels (shells)
Wealth	= Energy
City centre	= Nucleus

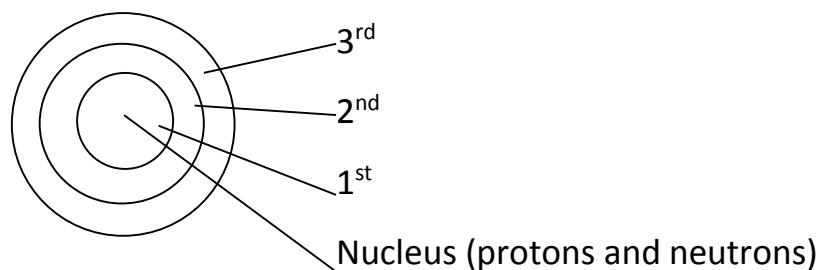
Look at Fig. 4.4. What would you get if you completed the dotted lines?

The neighbourhoods are represented by energy levels where an energy level is the path of an electron around a nucleus in an atom.



*Fig. 4.4 Atomic 'city' showing the neighbourhoods*

The various energy levels in an atom are represented by a series of circles sharing the same centre (nucleus), separated from other by roughly equal distances. The nucleus of an atom is at the centre of the circle. The electrons in the energy levels are represented by dots ( $\bullet$ ) or crosses ( $\times$ ). The energy levels are labelled  $1^{\text{st}}$ ,  $2^{\text{nd}}$ ,  $3^{\text{rd}}$ ,  $4^{\text{th}}$  and so on starting from the one nearest to the nucleus as shown in Fig. 4.6



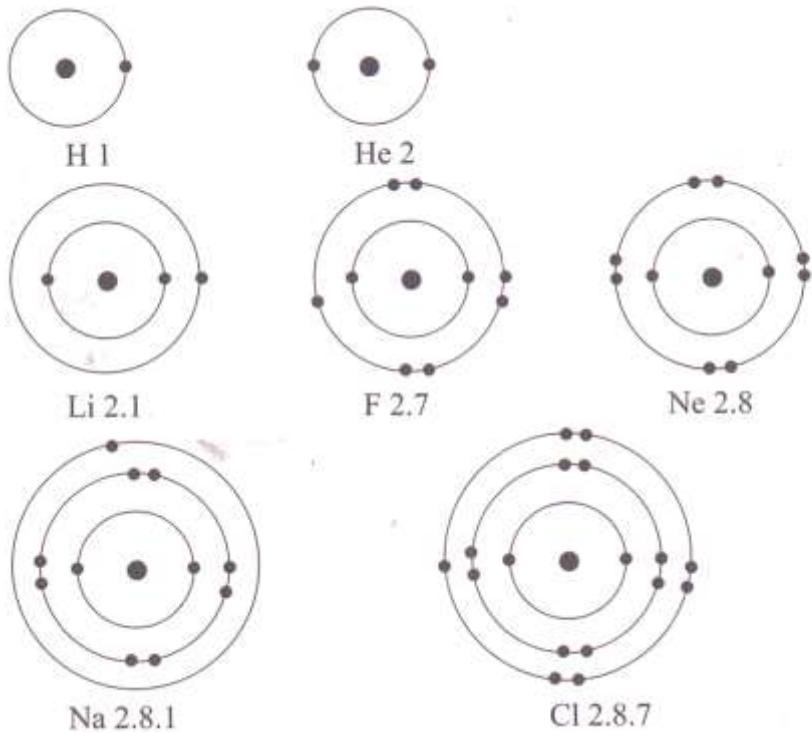
The electrons that occupy the  $1^{\text{st}}$  energy level have lower energy than those in the  $2^{\text{nd}}$  energy level. Subsequently those in the  $2^{\text{nd}}$  energy level have lower energy than those in the  $3^{\text{rd}}$  energy level and so on. The  $1^{\text{st}}$  energy level usually has a maximum of two electrons while the  $2^{\text{nd}}$  energy level has a maximum of eight (8) electrons. The maximum numbers of electrons that an energy level can hold are given by  $2n^2$ . Where  $n = \text{number of energy levels}$ ;  $1^{\text{st}}$ ,  $2^{\text{nd}}$ ,  $3^{\text{rd}}$  among others. This means that when we draw the atomic structure, we should never put more than two electrons in the  $1^{\text{st}}$  energy level or more than 8 in the  $2^{\text{nd}}$  level. The third energy level is usually not full but can accommodate one to eight electrons. (it holds a maximum of 8 electrons for the first 20 elements). The first twenty

elements of the periodic table have 1 to 20 protons and 1 to 20 electrons. We can summarise the electrons in each energy level as shown in table 4.2

Elements	Symbol	Number of protons	1 <sup>st</sup> Energy level	2 <sup>nd</sup> Energy level	3 <sup>rd</sup> Energy level	4 <sup>th</sup> Energy level	Electron arrangement
Hydrogen	H	1	•				1
Helium	He	2	..				2
Lithium	Li	3	..	•			2.1
Beryllium	Be	4	..	..			2.2
Boron	B	5	..	...			2.3
Carbon	C	6	..	....			2.4
Nitrogen	N	7	..	.....			2.5
Oxygen	O	8	..	.....			2.6
Fluorine	F	9	..	.....			2.7
Neon	Ne	10	..	.....			2.8
Sodium	Na	11	..	.....	•		2.8.1
Magnesium	Mg	12	..	.....	..		2.8.2
Aluminium	Al	13	..	.....	...		2.8.3
Silicon	Si	14	..	.....	...		2.8.4
Phosphorus	P	15	..	.....	...		2.8.5
Sulphur	S	16	..	.....	...		2.8.6
Chlorine	Cl	17	..	.....	...		2.8.7
Argon	Ar	18	..	.....	...		2.8.8
Potassium	K	19	..	.....	...	•	2.8.8.1
Calcium	Ca	20	..	.....	...	..	2.8.8.2

Arrangement of electrons in an atom helps us to explain the patterns in properties of the elements. These properties are the basis of the periodic table which we will discuss later.

Now let us illustrate the electron arrangement of the following atoms. Hydrogen, helium, lithium, fluorine, neon, sodium and chlorine respectively.



*Fig. 4.7 Atomic structure of H, He, Li, F, Ne, Na and Cl*

### Atomic characteristics

#### Atomic number and mass number

We have learnt how to draw atomic structure showing the nucleus and arrangement of electrons in various energy levels. It is important at this stage to know how to calculate the number of protons and neutrons inside the nucleus. This will enable us to show the number of these sub-atomic particles when we draw atomic structure diagrams. Atomic number and mass number are two simple numbers that tell us something about an atom.

Note the following points:

The atomic number tells us how many protons are there in a nucleus. It is denoted by letter Z. It also tells us the number of electrons in an atom

The atomic number (Z) of an element	=	The number of protons in the nucleus of its atom	=	The number of electrons in an atom
--	---	---	---	---------------------------------------

Now, why the mass number is double the atomic number?

- To get the number of neutrons, we just subtract the atomic number from the mass number, that is  $A - Z$ .

Study Table 4.3 which shows the relationship between atomic number, number of neutrons and mass number for some elements.

Table 4.3 Relationship between the number of protons, neutrons and mass number of some elements

Atom	Symbol	Number of protons Atomic number $Z$	Neutrons	Mass number $A$
Hydrogen	H	1	0	1
Carbon	C	6	6	12
Nitrogen	N	7	7	14
Sodium	Na	11	12	23
Chlorine	Cl	17	18	35

From Table 4.3, we notice that mass number,  $A$ , is a sum of protons and neutrons that is;

$$\text{Mass number } (A) = \text{number of protons} + \text{number of neutrons}$$

If we represent neutrons by  $N$ , we see that  $A = Z + N$

Using letters  $A$ ,  $Z$  and  $N$ , how can we get the number of neutrons ( $N$ ), in the nucleus of an atom?

The answer to this question is  $N = A - Z$

#### Example 4.1

Calculate the number of neutrons in a chlorine atom given that the atomic number,  $Z = 17$  and mass number,  $A = 35$

#### Solution

$$\text{The number of neutrons} = A - Z$$

$$= 35 - 17 = 18$$

Therefore, a chlorine atom has 18 neutrons.

Remember that since the number of protons and electrons are equal in an atom, once you know the atomic number which represents the number of protons, you automatically know the number of electrons in the atom.

Suppose you are required to draw the atomic structure of an atom structure of an atom showing the sub-atomic particles as in following example:

#### Example 4.2

A certain X has atomic number 6 and mass number 14.

Draw the atomic structure of element X showing all the sub-atomic particles and their numerical values.

#### Solution

The sub-atomic particles to show are protons, neutrons and electrons.

Atomic number = number of protons = number of electrons = 6

Number of neutrons = mass number – atomic number or  $A - Z$

$$= 14 - 6 = 8 \text{ neutrons}$$

This means that we will show 6 protons and 8 neutrons in the nucleus. The 6 electrons will be indicated in the energy levels as shown in Fig. 4.8.

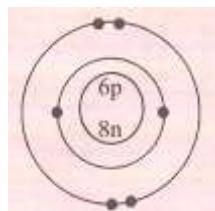


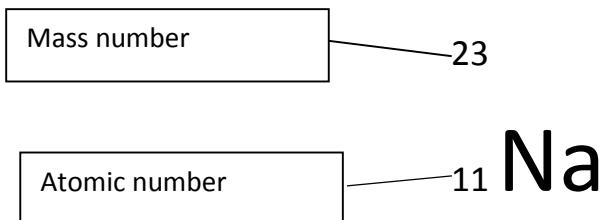
Fig. 4.8 Atomic structure of X

Usually, the atomic number,  $Z$ , and mass number,  $A$ , of an atom of an element X can be written alongside the symbol of that element, one as a superscript and the other as a subscript as shown below:

Mass number → A

Atomic number → Z  element

Note: the top number is referred to as the superscript and bottom number is the subscript. So the symbol for an atom of sodium would be written as;



We can also represent the sub-atomic particles using symbols with the mass of the particle as a superscript and the charge as a subscript as follows:

Proton - 1

P

1

Neutron n

0

Electron 0

# e

-1

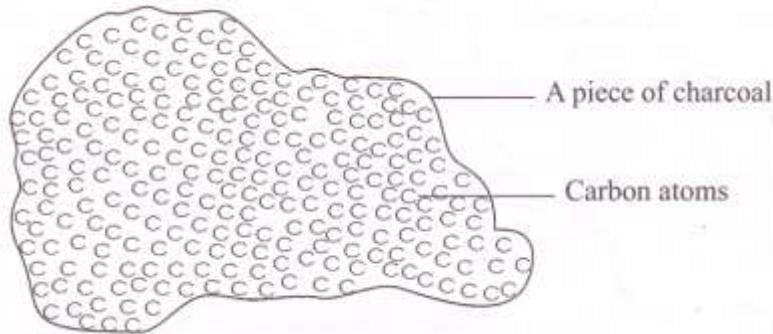
An electron has zero mass and a charge of -1. Can you tell what the letters p and n stand for?

## Isotopes and Relative Atomic Mass

We are all familiar with charcoal. May be some of us know how charcoal is formed. Burning of charcoal should be discouraged because cutting down of trees to make charcoal interfere with our water catchment areas. What is the colour of charcoal? What is charcoal in chemistry?

Charcoal is black in colour. In Chemistry charcoal is simply carbon.

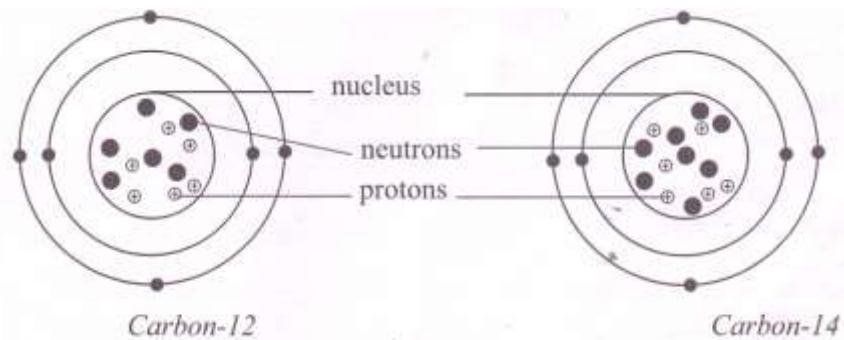
A piece of charcoal has millions and millions of carbon atoms. Let us represent a piece of charcoal as in Fig. 4.9



*Fig. 4.9 Millions of carbon atoms that make charcoal*

Carbon atoms exist in three forms represented as Carbon 12, Carbon 13 and Carbon 14.

C12 and C14 can be represented using diagrams as shown in Fig. 4.10



*Fig. 4.10 Different forms of Carbon*

What difference do you notice between the two carbon atoms?

The subscript numbers are the same but the superscript numbers are different. This means that they have the same atomic number but different mass numbers. What cause the difference in mass number?

Calculate the sub-atomic particles for each of the carbon atoms.

You will notice that the difference masses are caused by 2 extra neutrons in the nucleus of carbon -14. Such atoms are called isotopes.

Isotopes are atoms of the same element, which have the same number of protons in the nucleus (atomic number) but different mass numbers.

Some elements are made up of just one type of atom, while others exist as mixture of isotopes. Some of these elements are listed in Table 4.4.

**Table 4.4 Examples of elements with isotopes**

#### (a) Hydrogen

Hydrogen	Hydrogen (99.99%)	Deuterium (0.01%) heavy hydrogen	Tritium (Trace)
Symbols of isotopes of hydrogen	$^1_1\text{H}$	$^2_1\text{H}$	$^3_1\text{H}$
No. of protons	1	1	1
No. of neutrons	0	1	2
No. of electrons	1	1	1

Mass number (p + n)	1	2	3
---------------------	---	---	---

(b) Carbon

Carbon	Carbon-12 (98.9%)	Carbon-13 (1.1%)	Carbon-14 (Trace)
Symbols of Carbon isotopes	$^{12}_{\text{6}}\text{C}$	$^{13}_{\text{6}}\text{C}$	$^{14}_{\text{6}}\text{C}$
No. of protons	6	6	6
No. of neutrons	6	7	8
No. of electrons	6	6	6
Mass number (p + n)	12	13	14

(c) Chlorine

Chlorine	Chlorine-35 (75%)	Chlorine-37 (25%)
Symbols of chlorine isotopes	$^{35}_{\text{17}}\text{Cl}$	$^{37}_{\text{17}}\text{Cl}$
No. of protons	17	17
No. of neutrons	18	20
No. of electrons	17	17
Mass number (p + n)	35	37

(d) Oxygen

Oxygen	Oxygen-16(31.4%)	Oxygen-17 (33.3%)	Oxygen-18(35.3%)
Symbols of oxygen isotopes	$^{16}_{\text{8}}\text{O}$	$^{17}_{\text{8}}\text{O}$	$^{18}_{\text{8}}\text{O}$
No. of protons	8	8	8
No. of neutrons	8	9	10
No. of electrons	8	8	8

Mass number (p + n)	16	17	18
---------------------	----	----	----

### Relative Atomic Mass

Atoms, molecules and ions are single particles that are so tiny that their masses cannot be measured on a balance or scale directly. We cannot weigh an atom! Hydrogen is the lightest known element. A hydrogen atom is therefore the lightest. Originally, the mass of hydrogen atom was taken as the standard reference atom. Its atomic mass was arbitrarily and fixed as one unit that is, H = 1.

The mass of other element was found by comparing its mass with that of hydrogen. The idea was to find out how many times the atom of another element was as heavy as one atom of hydrogen hence relative atomic mass (R.A.M) with a symbol ( $A_r$ ). This can be expressed mathematically as follows:

$$\text{Relative Atomic Mass} = \frac{\text{mass of 1 atom of the element}}{\text{mass of 1 atom of hydrogen}}$$

For instance, an oxygen atom, (O), has a mass of 16. This means one oxygen atom is 16 times heavier than one atom of hydrogen.

Sometimes the symbol  $A_r$  is used for R.A.M with the symbol of the atom in parenthesis after the symbol. For example,  $A_r$  (O) means relative atomic mass of oxygen.

Many changes have occurred since the hydrogen scale was introduced. In the twentieth century, the hydrogen scale was replaced by oxygen as the standard reference atom since oxygen combines with most elements. But later the oxygen scale was found to be unsatisfactory because oxygen have several isotopes which would have different masses depending on which oxygen isotope was being considered (see table 4.4) (d).

In 1961 the International Union of Pure and Applied Chemistry (IUPAC) recommended the most abundant of the carbon isotopes, carbon – 12 ( $^{12}\text{C}$ ). as the standard reference atom. It has an abundance of 98.9%.

Nowadays, the atomic mass of an elements is measured by comparing it with  $\frac{1}{12}$  of the mass of one atom of a carbon- 12 ( $^{12}\text{C}$ ).

One atom of  $^{12}_{\text{6}}\text{C}$  isotope is taken to have a mass of 12.00 units.

$\frac{1}{12}$  of the mass of one atom of carbon – 12 = 1.00 units

$$A_r \text{ of an element} = \frac{\text{average mass of one atom of the element}}{\frac{1}{12} \times \text{mass of one atom of carbon} - 12}$$

The relative atomic mass ( $A_r$ ) of an element, is defined as the average mass of one atom of the element compared with  $\frac{1}{12}$  of the mass of one atom of carbon – 12

Note: relative atomic mass has no units. It is a ratio of two masses. The relative atomic masses are not whole numbers like mass numbers. This is because the abundance of isotopes of an element is different. For example, a sample of chlorine is a mixture of two isotopes,  $^{35}\text{Cl}$  and,  $^{37}\text{Cl}$  in the ratio of 3:1 respectively, 75% is,  $^{35}\text{Cl}$  and 25% is,  $^{37}\text{Cl}$ .

#### Calculation of the relative atomic mass

As mentioned earlier, many elements naturally consist of a mixture of isotopes. The abundance in which the isotopes occur in an element differs in different elements. This is why the term ‘average’ mass of one atom is used in the definition above. The proportions (abundance) in which the isotopes occur in an element may be stated as:

- a ratio
- percentage of the total
- a fraction of the total

#### Example 4.3

(You are given ratio abundance)

Chlorine consists of two isotopes, chlorine – 35 and chlorine 37 in the ratio 3: 1. Calculate the Relative Atomic Mass (R.A.M) of chlorine.

#### Solution

Suppose the sample contains 4 atoms of chlorine in the ratio 3 : 1 respectively, then 3 atoms will each have a mass of 35 and 1 atom will have a mass of 37.

The total mass of  $^{35}\text{Cl}$  =  $35 \times 3$

While the total mass of  $^{37}\text{Cl}$  =  $37 \times 1$

Therefore, the average mass of chlorine atoms will be:

$$\frac{\text{total mass of all atoms}}{\text{total number of atoms}} = \frac{(35 \times 3) + (37 \times 1)}{4}$$

$$= 26.25 + 9.25$$

$$\text{The R.A.M} = 35.5$$

Note: The value of the R.A.M is closer to the mass number of the most abundant isotope, that is,  $^{35}\text{Cl}$

#### Example 4.4

(You are given percentage abundance)

Silicon (Si = 14) consists of three isotopes: silicon – 28, 92.2%, silicon -29, 4.7% and silicon -30, 3.1%. Find the Relative Atomic Mass of silicon.

Solution

Percentage abundance simply means that if we have 100 atoms of an element called silicon, 92.2 atoms will each have a mass of 28

Therefore, the total mass of these =  $\frac{28 \times 92.2}{100}$

25.81 atoms will be silicon – 28

Therefore, the total mass of these =  $\frac{29 \times 4.7}{100}$

1.36 atoms will be silicon – 29

Therefore, the total mass of these =  $\frac{30 \times 3.1}{100}$

0.93 atoms will be silicon – 30

The average mass of a silicon atom is;

$$\begin{aligned}\text{Total mass of all atoms} &= \frac{28 \times 92.2}{100} + \frac{29 \times 4.7}{100} + \frac{30 \times 3.1}{100} \\ &= 25.81 + 1.36 + 0.93\end{aligned}$$

Therefore, R.A.M of silicon = 28.1

### Example 4.5

(Abundance is given as a fraction of the total)

A sample of an element X consists of  $\frac{9}{10}$  of  $^{18}_8\text{X}$ , show that the Relative Atomic Mass of X would be 16.2

Solution

$$\begin{aligned}\text{R.A.M} &= (16 \times \frac{9}{10}) + (18 \times \frac{1}{10}) \\ &= 14.4 + 1.8 \\ &= 16.2\end{aligned}$$

### Revision Exercise 4

1. Here is a symbol of Beryllium atom:



- (a) What is the name given to the superscript (top number)?
- (b) What is the name given to the subscript (bottom number)?
- (c) In the Beryllium atom shown above, calculate the number of protons, neutrons and electrons in the atom.

2. Select the pair which represents two atoms with the same number of neutrons.

- (a)  $^{12}_6\text{C}$  and  $^{24}_{12}\text{Mg}$
- (b)  $^{18}_8\text{O}$  and  $^{19}_9\text{F}$
- (c)  $^{23}_{11}\text{Na}$  and  $^{20}_{10}\text{Ne}$

3. Explain briefly why some elements have Relative Atomic Mass which are not whole numbers?
4. An atom is the smallest particle of an element. Name the sub-atomic particles found in an atom and state where they are found.
5. Explain the following terms:
  - (a) Atomic number
  - (b) Mass number
  - (c) Isotopes
6. The following symbols refers to isotopes of oxygen.
  - (a)  $^{16}_8\text{O}$
  - (b)  $^{18}_8\text{O}$

What is the number of protons and neutrons in the nucleus of each isotope?
7. A class was asked to select from the list below the elements whose electrons arrangements were impossible to write down  
 Which elements did the class select?  
 $^{23}_{11}\text{Na}$ ;  $^{10}\text{B}$ ;  $^6_6\text{C}$ ;  $^{12}_{6}\text{C}$ ;  $^{25}\text{Mg}$ ;  $^{26}_{12}\text{Mg}$

8. Write down and complete the following table:

Atom	Mass number	Number of protons	Number of neutrons	Number of electrons
F	19			9
Na			12	11
Ne	20	10		
C	12	6		

9. An atom of element X can be represented by the symbol  $^{16}_6\text{X}$ .
  - (a) State the:
    - (i) Number of electrons of atom X
    - (ii) Mass number of X
    - (iii) Atomic number of X
  - (b) Element X contains 90% of  $^{16}\text{X}$ , and 10% of  $^{18}\text{X}$ . Calculate the Relative Atomic Mass of X.

## **Unit 5**

## **The Periodic Table**

### **The main features of the Periodic Table**

In Unit 4 ‘Atomic structure’ you were introduced to energy levels as a series of circles in an atom, sharing the same centre (nucleus) separated from each other by equal distances. Energy levels are usually labelled 1<sup>st</sup>, 2<sup>nd</sup>, 3<sup>rd</sup> and so on depending on the number of electrons of an element. The 1<sup>st</sup> energy level can accommodate a maximum of (2) two electrons. The 2<sup>nd</sup> and 3<sup>rd</sup> energy levels can accommodate a maximum of (8) eight electrons. In this unit, we shall discuss how the number of electrons in the outermost energy level and the total number of energy levels are used to describe the groups and periods of elements in the Periodic Table.

### **Activity 5.1**

Consider the electronic configuration of the following elements:

Lithium (Li) 2.1

Carbon (C) 2.4

Sodium (Na) 2.8.1

(a) What is the similarity between electronic configuration of:

- (i) Li and C
- (ii) Li and Na

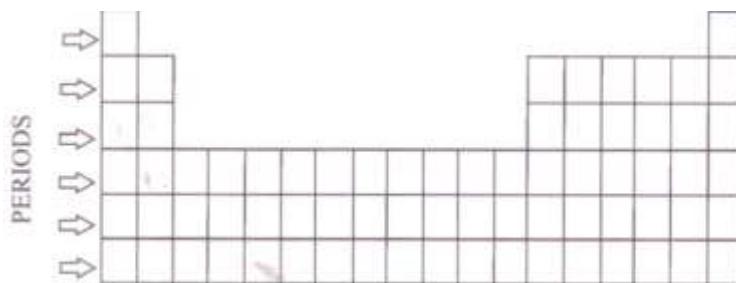
(b) What is the difference between electron configuration of Lithium and Sodium

In activity 5.1, you will note that lithium and carbon have 2 energy levels only. Lithium and sodium atoms have one electron in their outer most shell or energy level although they have different numbers of energy levels. Elements can be classified or grouped according to the number of electrons in the outer most shell. They can also be grouped according to the number of shells. Such an array of elements is called a Periodic Table.

### Structure of the Periodic Table

The Periodic Table is organized as a grid. In a grid, there are rows and columns. The Periodic Table has rows and columns. The elements are placed in specific columns because of the way they behave. They have similar chemical properties because they have the same number of electrons in their outer most energy level.

In a Periodic Table, each of the rows is called a period. See Fig. 5.1

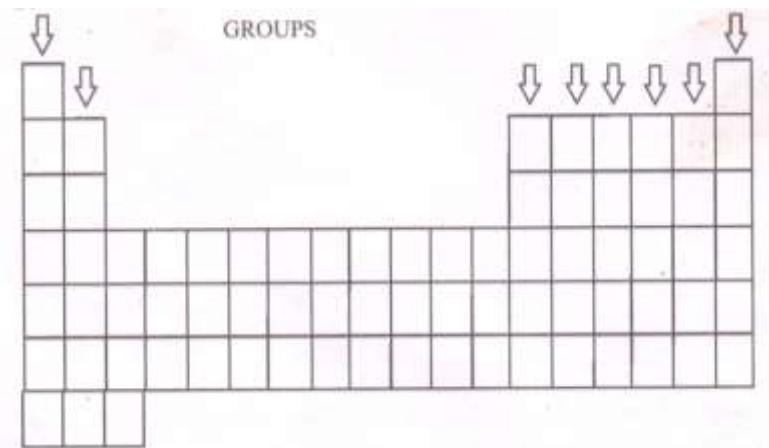


*Fig. 5.1 Periods in the Periodic Table*

All the elements in the same period have the same number of energy levels. If we want to know in which period to find an element we count the number of energy levels. If it has 4 energy levels; it is in the 4th period.

The two elements at the top row (the first period) have one energy level. All of the elements in the second row (the second period) have two energy levels. The third row elements have three energy levels and so on.

The Periodic Table has a special name for its columns too. A column in the Periodic Table is called a Group. The elements in a Group have the same number of electrons in their outermost energy levels.



*Fig. 5.2 Groups in the Periodic Table*

An element in the first column belongs to Group I. it has one electron in its outermost energy level. An element in the second column belongs to Group II and has two electrons in the outermost energy level. The element in Group III has three electrons in the outermost energy level and so on. Therefore the Group number indicates the number of electrons in the outermost energy level.

#### Distribution of Elements in the Periodic Table

The principal characteristics of the Periodic Table are:

- a. The elements are arranged in an order which depends on one of their fundamental characteristics.
- b. The elements are divided into groups in such a way that similar elements lie in the same group.
- c. The elements within the same group will not necessarily have identical properties. But the way in which properties differ from one element to another (gradation of properties) conforms to some pattern; so knowing the properties of one element in a group, it should be possible to predict the properties of another in the same group.
- d. Similarly, the gradation of properties from one group of elements to the next group of elements conform to some pattern.

The general structure of the table is:

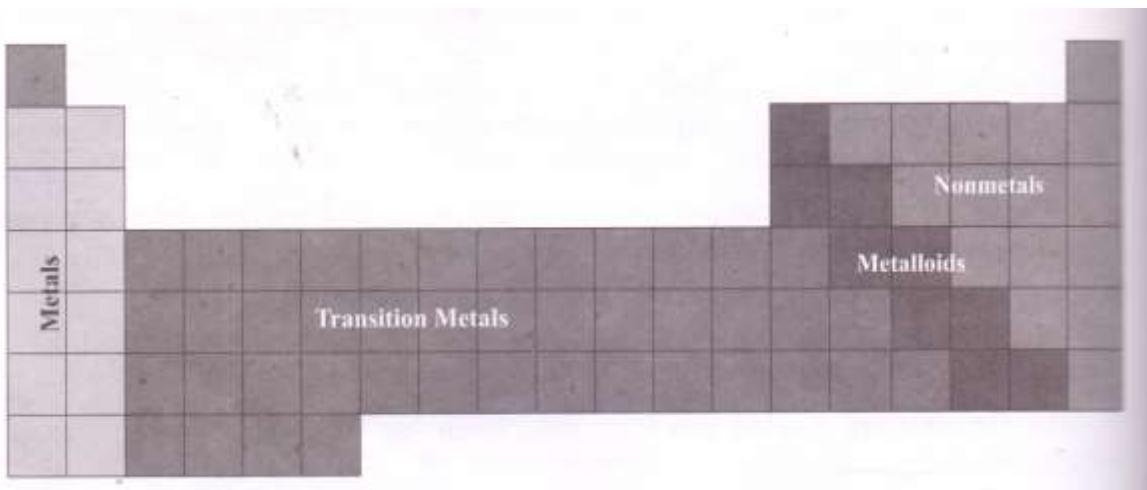
- The elements are arranged in the order of the atomic number starting with hydrogen atomic number 1.
- The elements are divided into eight groups running vertically. Some groups of elements are given names:
  - Group I – the alkali metals
  - Group II – the alkaline earth metals
  - Group VII – the halogens
  - Group VIII – the noble gases.
- The three rows of elements forming the large block in the centre of the table are called the transition metals.
- The elements are also divided into periods which run horizontally. The number of elements in each complete period is significant:
  - Period 1 has 2 elements
  - Period 2 has 8 elements
  - Period 3 has 8 elements
  - Period 4 has 18 elements
  - Period 5 has 18 elements
  - Period 6 has 32 elements
  - Period 7 has 17 elements

Note: you will notice that this arrangement corresponds to the number of electrons in successive energy levels for the first twenty elements.

- Metals are on the left-hand side of each period. The properties of the elements in the Group become more metallic as the atomic number increases (that is, down the Group).

#### Distribution of metal and non-metal elements in the Periodic Table

It is possible to predict which elements are metals and which elements are non-metals depending on how they are arranged in the Periodic Table. In general, the elements on the left of the Periodic Table are metals while those to the right are non-metals. In between the two are a special group of elements known as metalloids or semi-metals. This is because their properties are in between those of metals and those of non-metals. Fig. 5.3 shows the positions of these blocks of elements in the Periodic Table.



*Fig. 5.3 Blocks of elements in the Periodic Table.*

- Elements which are gases at room temperature and pressure are in the top right corner of the Periodic Table.
- The number of electrons in the outermost energy levels of any element is the same as the Group number of the Group containing the element, with an exception of the noble gas family.
- The most reactive elements are in Group I, the most non-reactive elements are in Group VIII
- The most reactive non-metals are in the earlier periods, for example, fluorine is more reactive than chlorine. The most reactive metals are in later periods, for example, potassium is more reactive than sodium.

#### Electron configuration and the Periodic Table

Electron configuration refers to how the electrons are arranged in the energy levels of an atom.

Given the atomic number of an element, we can write the electron arrangement in shorthand. We shall illustrate this in the following example.

#### Example 5.1

The atomic numbers of X, Y and Z are 12, 13 and 17 respectively. Write the electron arrangement of each.

#### Solution

X 2.8.2

Y 2.8.3

Z 2.8.7

All the three elements are in period 3 because each has three energy levels. X is in Group II, Y in Group III and Z in Group VII, following the number of electrons in their outermost energy levels.

	Group I	Group II	Group III	Group IV	Group V	Group VI	Group VII	Group VIII
Period 1	H 1							He 2
Period 2	Li 2.1	Be 2.2	B 2.3	C 2.4	N 2.5	O 2.6	F 2.7	Ne 2.8
Period 3	Na 2.8.1	Mg 2.8.2	Al 2.8.3	Si 2.8.4	P 2.8.5	S 2.8.6	Cl 2.8.7	Ar 2.8.8
Period 4	K 2.8.8.1	Ca 2.8.8.2						

*Fig. 5.4 Electron arrangement of the first 20 elements of the Periodic Table*

Hydrogen is a unique element. It is placed in Group I as its atoms have one electron in the outermost energy level. It also fits in Group VII because its atoms have one electron short of a fully filled outermost energy level.

Helium is different from all other elements. It has two electrons in its single energy level which can accommodate a maximum of 2 electrons. Therefore, its outermost energy level is full and cannot accommodate extra electrons.

Note: The atomic number also gives you the number of electrons. The electronic configuration gives you the number of energy levels hence the period. The valence electrons (in the outermost energy level) gives you the group of the element.

The atomic structure diagrams below show how you can determine the group number and the period of an element given the atomic number.

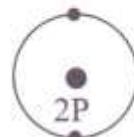
(a) Hydrogen

Atomic Number 1  
Electronic configuration 1  
Group 1  
Period 1



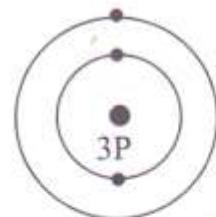
(b) Helium

Atomic Number 2  
Electronic configuration 2  
Group VIII  
Period 1



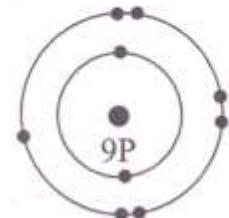
(c) Lithium

Atomic Number 3  
Electronic configuration 2.1  
Group 1  
Period 2



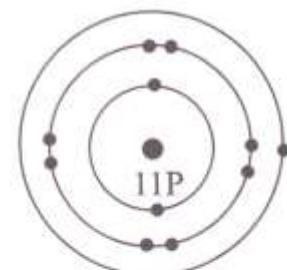
(d) Fluorine

Atomic Number 9  
Electronic configuration 2.7  
Group VII  
Period 2



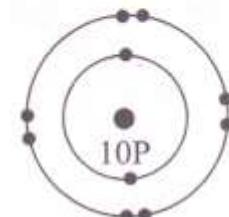
(e) Sodium

Atomic Number 11  
Electronic configuration 2.8.1  
Group I  
Period 3



(f) Neon

Atomic Number 10  
Electronic configuration 2.8.2  
Group VIII  
Period 2



### 5.3 Family names of some elements in the periodic table

Elements belonging to particular groups are given different family names as follows:

- Group I elements – Alkali metals
- Group II elements – Alkali earth metals
- Group VII Elements – halogens
- Group VIII elements – Noble gases

Examples of Alkali metals are sodium, potassium and lithium.

Examples of alkali earth metals are beryllium, magnesium and calcium

Examples of halogens are fluorine, chlorine, bromine and iodine.

Examples of noble gases are neon, argon and krypton.

I	II	III	IV	V	VI	VII	VIII
H 1	Be 2		B 3	C 4	N 3	O 2	He 0
Ei 1	Mg 2		Al 3	Si 4	P 3 or 5	S 2	Ne 0
Na 1	Ca 2					Cl 1	Ar 0
K 1							

Noble gases ↓

↑ Halogens

Alkali metals ↑

### Revision exercise 5

1. Name:
  - (a) Three alkali metals
  - (b) Three halogens
  - (c) Three noble gases
  - (d) One alkaline earth metals

2. Which of the following elements belong to the same Group of the Periodic Table? Lithium, magnesium, potassium, beryllium, helium and argon.
3. Which of the following elements belong to the same period of the Periodic Table? Sodium, lithium, fluorine, argon, hydrogen, helium.
4. Some elements (denoted by letters A and F which are not their chemical symbols) have atomic numbers as follows: A = 3, B = 9, C = 12, D = 16, E = 18, F = 20
  - (a) Which letter represents a halogen?
  - (b) Which letter represents a noble element?
  - (c) Which two elements are in the same Group of the Periodic Table?
5. An element is in period 3, Group VI. Draw its atomic structure.
6. Copy the puzzle below on a piece of paper then for each description below, write the missing letters of the word that best fits the description.
  - (a) An alkali metal.
  - (b) A gas found in the air whose atoms have an electronic configuration of 2.5.
  - (c) An element with an atomic number of 17
  - (d) An element whose atom has one electron in the outermost energy level.
  - (e) An element in the period number 3 and Group number V.
  - (f) A noble gas.
  - (g) An element with electronic configuration of 2.4
  - (h) A horizontal row in the Periodic Table.
  - (i) A halogen whose atom has 7 electrons in the outermost energy level.

C\_\_\_\_\_

  H\_\_\_\_\_

  E\_\_\_\_\_

  M

  I

  S\_\_\_\_\_

  T

  R

  Y\_\_\_\_\_

**Introduction**

Heat causes many substances to change. Do you remember what happens to ice and water when they are heated? What happened when these substances were cooled? We shall learn more about the effect of heat on substances in this section.

Heat can be dangerous, even fatal, if mishandled, not just in the laboratory, but at home as well. Always exercise caution whenever you are using heat. In chemistry, observations can be made on experiments that involves heating substances.

This section gives some guidelines to follow when heating substances in the laboratory. In addition, always follow the instructions given for specific experiments.

**Safety rules to remember when heating substances**

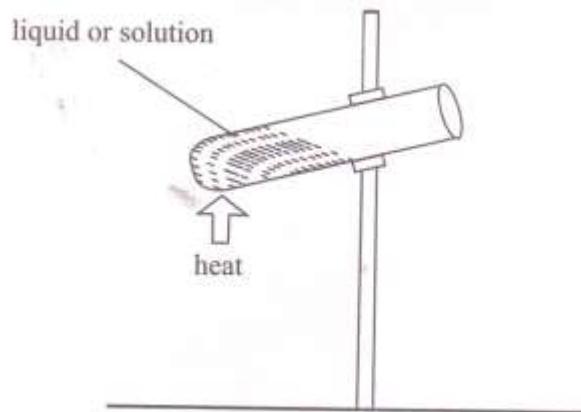
- Ensure that the test tube or boiling tube is dry before use.
- Make use of test tubes to heat solids, but, boiling tubes are preferred when boiling liquids or solutions. For large quantities of liquids or solutions, beakers are preferred. Hold the test tube or boiling tube with a test tube holder or folded paper.
- Boil liquids or solutions in boiling tubes or test tubes that are less than a fifth full. See Fig. 6.1.
- While boiling liquids or solutions, face the mouth of the test tube away from yourself, friends or books in case of spouting.
- Always heat gently at first, holding the tube at an angle to the flame. From time to time, move the tube away from the flame. If you have no smell any gases, hold the tube some distance from the nose, then using your hand, waft some gas above the tube towards your nose.

In this way you can safely smell any gas that might be given off as shown in Fig. 1.6 in Unit 1.

- When heating solids, look for changes in colour, or the colour of gas being produced, listen out for sound too.

Also look out for water droplets condensing and sublimating (sublimate) on the cooler parts of the tube.

- If necessary, heat strongly and repeat the observation for smell and any other evidence for change. Note the change when the residue cools.
- Always use a small non-luminous flame so that blackening of the glass does not occur. Blackening of the tube makes it difficult to make accurate observations.



*Fig. 6.1 The amount of liquid boiled in a test tube should be a firth or less*

## 6.1 Difference between physical and chemical changes

### Experiment 6.1 (a)

Aim: To investigate the changes that take place during the heating and cooling of a substance

### Apparatus and chemicals

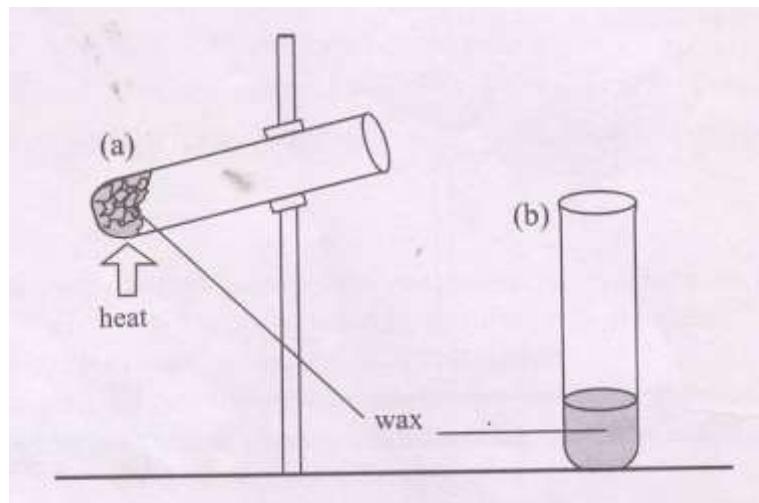
- Test tubes
- Candle wax
- Iodine crystals
- Bunsen burner
- Ice
- Test tube holder
- Zinc oxide

## Procedure

1. Place pieces of candle wax in a test tube. Hold the test tube with the test tube holder and heat. Allow it to cool.
2. Repeat the experiment with small quantities of the substances. Use a clean dry test tube in each case. Copy Table 6.1 in your notebooks and fill in the empty columns. These are your observations.

**Table 6.1 Changes occurring in some substances on heating and cooling**

Substance	Appearance	Result on heating	After heating
Iodine crystals			
Candle wax			
Zinc oxide			
Ice			



*Fig. 6.2 (a) Heating and (b) cooling wax*

## **Experiment 6.1 (b)**

Aim: To determine what happens when a substance is dissolved in water

### **Apparatus and chemicals**

- Test tubes
- Sucrose
- Water

- Bunsen burner
- Test tube holder
- Common salt (sodium chloride)

### **Procedure**

1. Place some common salt in a test tube. Add water to fill about the test tube and shake well.
2. Place the mixture in an evaporating dish
3. Heat to evaporate the water
4. Copy Table 6.2 in your notebook and record your observations.
5. Repeat the experiment with a little sugar. Record your observation in Table 6.2.
6. Pour the mixture in an evaporating dish. Heat to evaporate the water.

Table 6.2 Dissolving some substances in water and the effect of heat on them

Substance	Observation after shaking in water	Observation after evaporation
Common salt (sodium chloride)		
Sugar (sucrose)		

You may have noted the following observations when the various substances are heated:

(a) Wax

When wax is heated, it melts. On cooling, it solidifies to form the solid wax again.

(b) Iodine

Iodine forms a purple gas when heated. On cooling, the gas forms solid iodine again. We refer to such change as sublimation.

(c) Sugar

Sugar dissolves in water. When water is evaporated, it is not easy to recover the sugar. Why is this so?

(d) Zinc oxide

Zinc oxide is a white solid when cold. When heated, it does not melt but just turns yellow. On cooling, it turns to the original white substance.

(e) Common salt

When common salt is shaken with water, it dissolves. When water is evaporated the salt sample was recovered.

In the above experiments, the changes which took place are reversible. This means that the original substances were obtained upon cooling. None of the substances heated formed a new substance. Changes like melting, boiling, evaporation and sublimation are reversible. They are called non-permanent or physical changes. Some substances change colour when heated and return to the original colour when cooled; for example, zinc oxide. It is yellow when hot and white when cold. Note that zinc oxide changed in colour only. It undergoes a physical change.

## Experiment 6.2

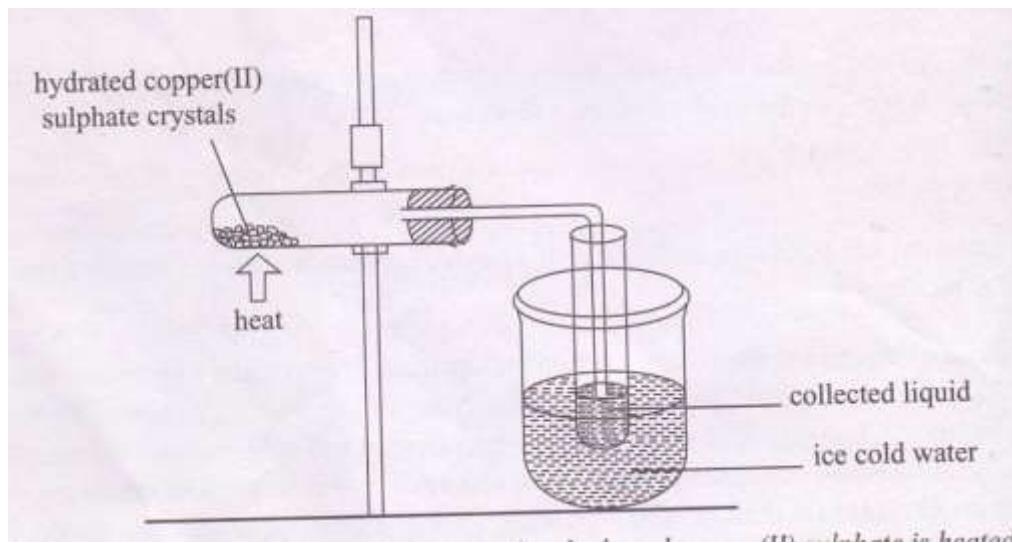
### Aim: To investigate the effect of heat on hydrated copper (II) sulphate

#### Apparatus and chemicals

- Test tubes
- Test tube holder
- Hydrated copper (II) sulphate
- Bunsen burner
- Thermometer

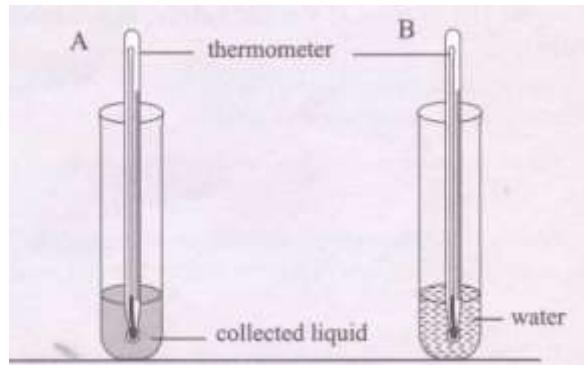
#### Procedure

1. Place some small hydrated copper (II) sulphate in a dry test tube (if crystals are wet, dry them with a blotting filter paper). Note the colour of the crystals.
2. Set up the apparatus as shown in Fig. 6.3
3. Heat hydrated copper (II) sulphate until there is no further changes.
4. Remove the delivery tube while still heating, otherwise the collected liquid will be sucked into the cooling tube.
5. Allow the test tube to cool and divide the residue solid into two portions.
6. Label the test tube containing the collected liquid as 'A'. measure and record the temperature of the liquid. Put an equal amount of water in another test tube and label it as 'B'. Measure the temperature of the water and record it.



*Fig. 6.3 Set-up to show what happens when hydrated copper (II) sulphate is heated.*

7. Add the first portion of the residue into liquid in test tube A and record the temperature change. Add the second into test tube B containing water and record the temperature as shown in (fig. 6.4).
8. Draw Tables 6.3 and 6.4 and fill them.
  - (a) Why should we remove the delivery tube from the liquid collected while still heating?



*Fig. 6.4 Temperature changes on adding heated copper (II) sulphate to different liquids*

- (b) From your results in Table 6.3, what is the liquid collected?
- (c) How can we check whether this liquid is pure?
- (d) From the change in temperature recorded in Table 6.4, is heat energy released or absorbed? Give a reason.
- (e) How can we classify the changes in this experiment? Give reasons.

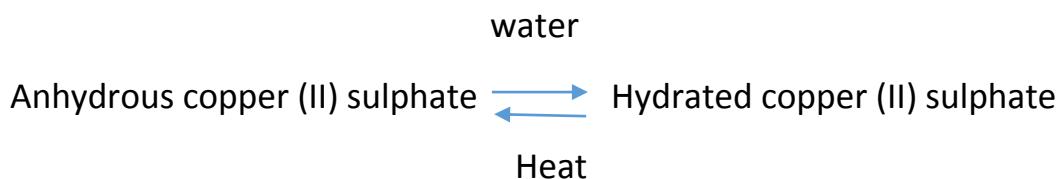
Table 6.3 Effect of heat on copper (II) sulphate

Hydrated copper (II) sulphate	Before heating	After heating
Colour		

Table 6.4 Heat absorbed or released on addition of heated copper (II) sulphate to different liquids

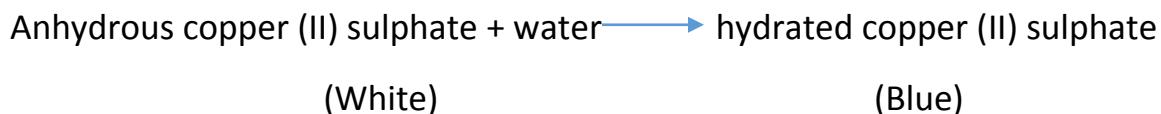
	Test tube A	Test tube B
Final temperature °C		
Initial temperature °C		
Change in temperature °C		

When hydrated copper (II) sulphate is heated, a colourless liquid is given out i.e. water. It turns white anhydrous copper (II) sulphate blue.

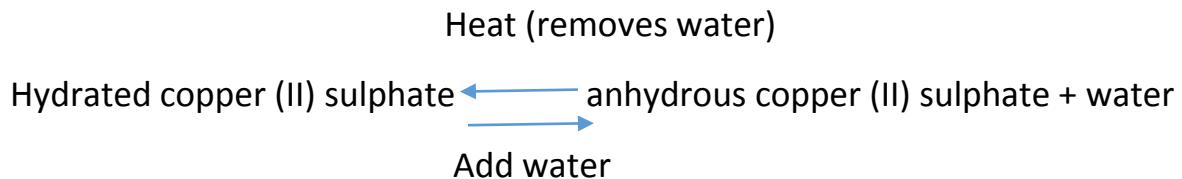


Hydrated copper (II) sulphate contains water molecules. Note that we are using heat energy to drive the water molecules out of the crystals. This water is known as water of crystallization. It is the one that gives the crystals a blue colour. Without water of crystallization, copper (II) sulphate is a white powder. It is said to be anhydrous copper (II) sulphate.

When we add water to anhydrous copper (II) sulphate, it turns blue, that is becomes hydrated. The reaction that takes place is as follows:



This is an example of a reversible chemical change. If we kept anhydrous copper (II) sulphate in a moisture-free place, it will remain white. We can write the above reaction as follows:



There are other reversible chemical changes that you will study later

### Chemical changes

All the reactions given above are examples involving physical change. Let us now find out what happens during chemical change.

### Experiment 6.3

Aim: To investigate the effects of burning a piece of paper, wood and magnesium ribbon.

### Apparatus and chemicals

- A piece of wood
- Paper
- Magnesium ribbon

### Procedure

1. Cut a piece of paper
2. Hold with a pair of tongs or a split piece of wood.
3. Light a burner and the piece of paper.
4. Repeat the experiment using a piece of wood.
5. Then cut a piece of magnesium ribbon about 4-5cm.
6. Hold the magnesium ribbon securely with a pair of tongs.
7. Then hold the end of tongs in the flame until the magnesium ribbon catches fire.

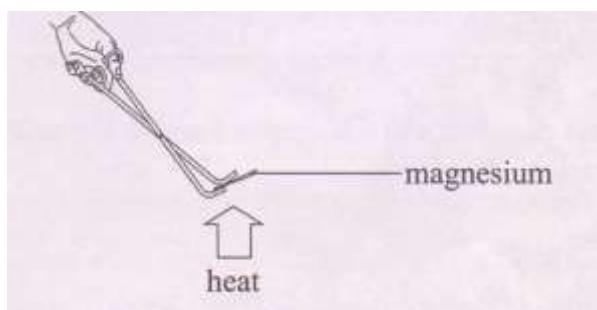


Fig. 6.5 Burning magnesium in air

**Caution:** Do not look at the flame of burning magnesium directly.

8. Draw Table 6.5 in your notebook and record your observations.

Table 6.5 Effect of heat on some substances

Substance	Describe the type of flame produced during burning	Colour of substance formed
Paper		
Wood		
Magnesium		

What is the difference between the original substance and the substance formed?

Can one re-form the original substance from the product formed?

The substance, in experiment 6.3, burnt with different types of flames. The paper and wood burnt with a luminous yellow flame. Magnesium burnt with a very bright dazzling flame. It forms white ash which is a new substance called magnesium oxide. You cannot get magnesium from the oxide. It is an irreversible reaction. The reaction that takes place is as follows:



The substances formed are totally different from the original substances. This indicates that new substances are formed after burning. It is impossible to reverse the products formed to the original substances. The change that has taken place is permanent. Such changes are also called chemical changes or chemical reactions and are difficult to reverse.

There are many examples of chemical changes. They include oxidation of food in our bodies, rusting of iron, burning of kerosene, petrol and charcoal.

Table 6.6 Difference between physical and chemical changes

Physical change	Chemical change
1. No new substance is formed	A new substance is formed
1. No energy is either given out or absorbed.	Energy is usually given out or absorbed
2. The mass of the substance does not change.	The mass of the substance changes

3. The change is usually reversible	The change is usually irreversible
-------------------------------------	------------------------------------

However, there are some exceptions to difference 2 and 4. For example, there are some chemical reactions which are reversible.

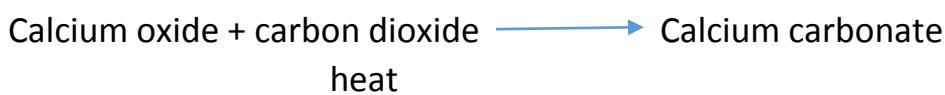
### Example 6.1

heat



This reaction can be reversed as follows:

Heat



This can be reversed as follows:

heat



We will learn more about reversible chemical changes later.

### Chemical reactions

In chemistry the substances we burn or heat are generally called reactants and the substances formed are referred to as products. When burnt substances usually combine chemically with oxygen in the air.



### Simple word equations

When a chemical reaction takes place, we can represent it in form of word equation. When magnesium burnt in air, it combined with the oxygen in the air to form magnesium oxide. We can summarize this in form of a word equation.



- The plus sign (+) in chemistry when used in an equation means 'reacts with' (when used on the left)
- The arrow ( $\rightarrow$ ) means to form the products on the right.

When there is only one substance on the left as in the following equation:



This means when calcium carbonate is heated it decomposes (splits up) to form calcium oxide and carbon dioxide.

- Two full-headed ( $\longleftrightarrow$ ) arrows mean it is reversible chemical reaction.

Remember, you cannot write chemical equations if you do not know the common symbols and valences.

The number of electrons an atom requires to attain the stable noble gas electron arrangement is known as valency or combining power of the atom or group of atoms. A group of atoms which occur in compounds but cannot exist on their own are called a radical, for example sulphate,  $\text{SO}_4^{2-}$  or hydroxide,  $\text{OH}^-$ . In metals, the valency is just the number of electrons in the outermost energy level. For non-metals it is the difference between Group VIII and Group number of the element, for example the valency of oxygen is  $8 - 6 = 2$ , valency of phosphorus is  $8 - 5 = 3$  etc. Valencies of various elements are given in the table 6.7

Table 6.7 Valencies of the first twenty elements in the Periodic Table

Atomic number	Element	Symbol	Electron arrangement	To gain stability of a noble gas atoms gain or lose electron	Valency
1	Hydrogen	H	1	Not stable, can lose $1e^-$	1
2	Helium	He	2	Stable, cannot lose/gain	0
3	Lithium	Li	2.1	Not stable, loses $1e^-$	1
4	Beryllium	Be	2.2	Not stable, loses $2e^-$	2
5	Boron	B	2.3	Not stable, loses $3e^-$	3
6	Carbon	C	2.4	Not stable, loses $4e^-$	4
7	Nitrogen	N	2.5	Not stable, easy to gain $3e^-$	3
8	Oxygen	O	2.6	Not stable, easy to gain $2e^-$	2
9	Fluorine	F	2.7	Not stable, easy to gain $1e^-$	1
10	Neon	Ne	2.8	Stable, cannot gain or lose electron	0

11	Sodium	Na	2.8.1	Not stable, loses $1e^-$	1
12	Magnesium	Mg	2.8.2	Not stable, loses, $2e^-$	2
13	Aluminium	Al	2.8.3	Not stable, loses $3e^-$	3
14	Silicon	Si	2.8.4	Not stable, loses 4 or gains $4e^-$	4
15	Phosphorus	P	2.8.5	Not stable, gains $3e^-$	3 or 5
16	Sulphur	S	2.8.6	Not stable, gains $2e^-$	2
17	Chlorine	Cl	2.8.7	Not stable, loses $1e^-$	1
18	Argon	Ar	2.8.8	Stable (octet) cannot gain or lose electron	0
19	Potassium	K	2.8.8.1	Not stable, loses $1e^-$	1
20	calcium	Ca	2.8.8.2	Not stable, loses $2e^-$	2

Table 6.8 Valencies of other common elements

Name of metal	Symbols	Valency
Zinc	Zn	2
Iron	Fe	2 or 3
Tin	Sn	2 or 4
Lead	Pb	2 or 4
Copper	Cu	1 or 2
Silver	Ag	1
Barium	Ba	2

Table 6.9 Valencies of common radicals

Valency 1		Valency 2		Valency 3	
Radical	Formula	Radical	Formula	Radical	Formula
Ammonium	$\text{NH}_4^+$	Carbonate	$\text{CO}_3^{2-}$	Phosphate	$\text{PO}_4^{3-}$
Hydroxide	$\text{OH}^-$	Sulphate	$\text{SO}_4^{2-}$		
Nitrate	$\text{NO}_3^-$	sulphite	$\text{SO}_3^{2-}$		
Chloride	$\text{Cl}^-$				
	$\text{HCO}_3^-$				

Hydrogen carbonate Hydrogen sulphate	$\text{HSO}_4^-$			
---	------------------	--	--	--

We already mentioned that valency is the combining power of an element or radical. However some elements have varying valencies, for example, copper can have valency of 1 and 2. The valencies of these elements are indicated in Roman numbers in brackets after the name of the element when naming compounds. This is illustrated in Table 6.10.

Table 6.10 Valencies of elements in compounds

Compound	Elements	Valency
Copper (I) oxide	Copper	1
Copper (II) oxide	Copper	2
Iron (II) sulphate	Iron	2
Iron (III) chloride	Iron	3
Sulphur (IV) oxide	Sulphur	4
Sulphur trioxide	Sulphur	6
Carbon (IV) oxide	Carbon	4
Carbon (II) oxide	Carbon	2

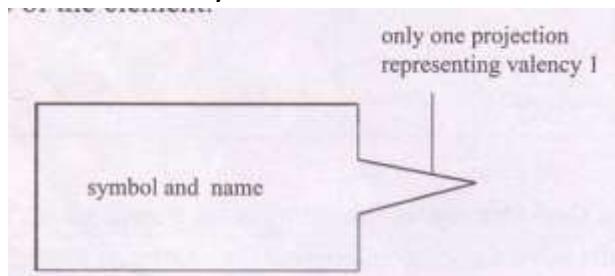
### Writing chemical formulae

Every subject has its own special language as you might have realized now. For example, in Chemistry we have substances with names like sodium chloride. However, we have a short notation of writing names called formula. The formula of a compound shows the atoms present in the compound in their simplest ratio. The formula of sodium chloride is  $\text{NaCl}$ . The formula for water is  $\text{H}_2\text{O}$ . Why then do we write the number 2 in the formula for water whereas we do not in sodium chloride? These are some of the questions we shall answer in this section as we discuss how to write chemical formulae.

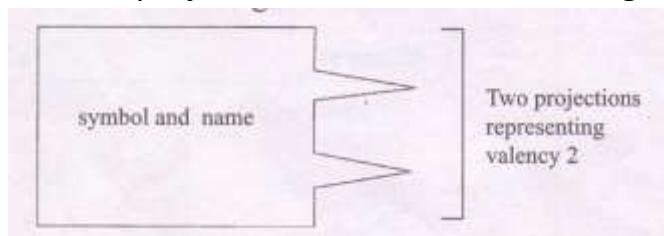
### Activity 6.2

Aim: To write a chemical formulae using a game of cards

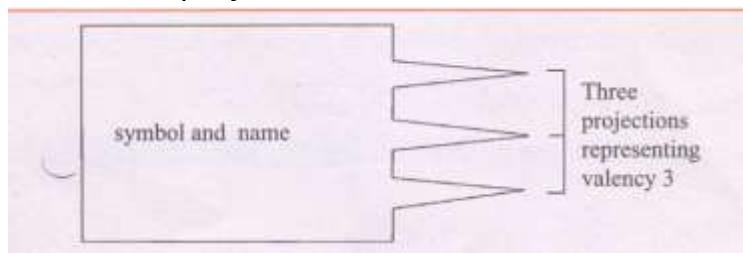
1. Get a piece of plain paper.
2. For group 1 elements or cations, with valency 1, cut these piece of paper into the following shape but make sure that they are of the same size. Indicate the symbol and the name of the element.



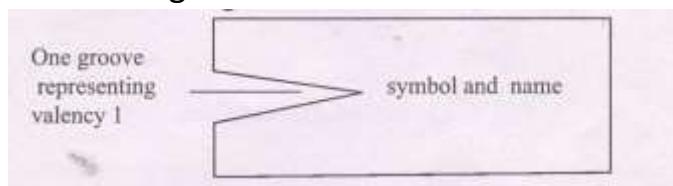
3. For group 2 elements or cations with valency 2, cut the paper such that it has two projections as shown as in the figure below.



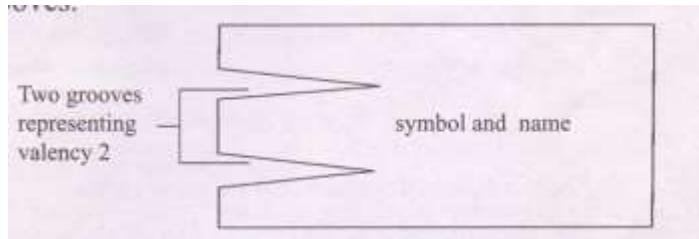
4. For group 3 elements or cations with valency 3, your piece of paper will have three projections as shown below.



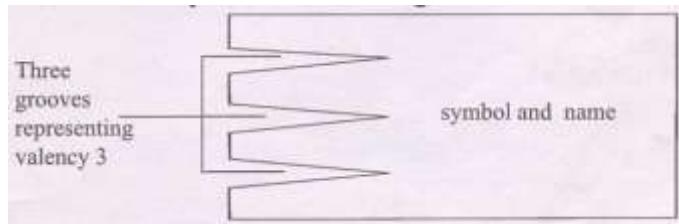
5. For group 7 elements or anions with valency 1, cut the piece of paper so that it has groove as shown below.



6. Group 6 elements or anions with valency 2, will be represented by a piece of paper with two grooves.



7. For elements with valency 3, cut out three grooves.



Note: It is important to take care of the following point before you play this game.

- When writing the formula, metals are always written first. For example, it is not correct to write ClNa or Br<sub>2</sub>Ca even though it gives us the same information. The correct way is NaCl and CaBr<sub>2</sub> respectively.

Now let us use another simple method of writing chemical formulas. Always remember that to write a correct formula we must write down the;

- Correct symbol of the element or radical
- Correct valency of the symbol or radical

Let us write the formula for sodium sulphate again following the steps below:

Step 1

Write the symbols of the elements and radical.

Na              SO<sub>4</sub>

Step 2

Write the valencies of the element and radical above and to the right side of each.



Step 3

Exchange the valencies by writing them below the symbols as shown by the arrows.



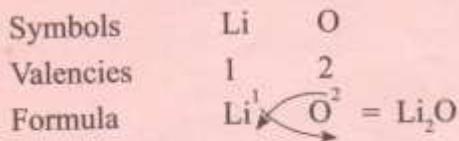
Step 4

Write the symbols close together

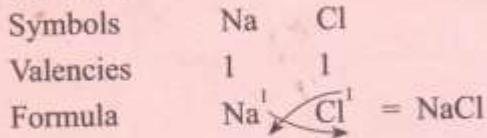


### **Other examples**

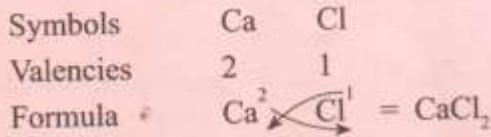
**(a)** Lithium oxide



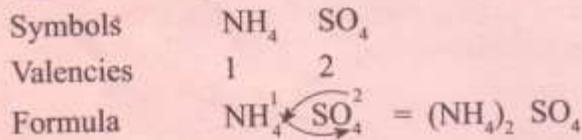
**(b)** Sodium chloride



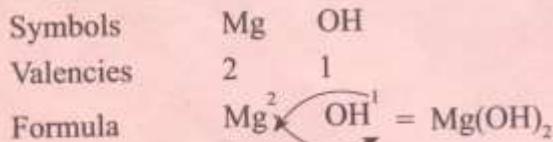
**(c)** Calcium chloride



**(d)** Ammonium sulphate

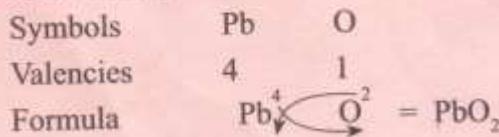


**(e)** Magnesium hydroxide



**Note:** We write brackets for (d) and (e) because the radical consists of two different elements for example OH contains oxygen atom and hydrogen atom;  $\text{NH}_4$  consists of nitrogen and hydrogen atoms. But it is wrong to put bracket for (c)  $\text{Ca(Cl)}_2$  as "C" and "I do not represent different elements.

**(f)** Lead dioxide.



**Note:** Before we bring the valencies down, we must divide the valencies by the common factor, 2.

(g) Iron trioxide

Symbols	Fe	O
Valencies	3	2
Formula	$\text{Fe}^3 \curvearrowright \text{O}^2 = \text{Fe}_2\text{O}_3$	

Here the valencies have no common number which can divide them and give us a whole number. Therefore, we just exchange the valencies as shown  $\text{Fe}_2\text{O}_3$

Use the card game for the above worked examples and to see whether you get the same formulae.

### Activity 6.3

Write the formulae for the following compounds:

- (a) Calcium carbonate
- (b) Lead oxide
- (c) Iron oxide
- (d) Iron hydroxide
- (e) Lead chloride
- (f) Sodium dioxide
- (g) Copper bromide
- (h) Iron chloride

### Writing simple balanced chemical equations

Chemical equations are short, clear and accurate descriptions of chemical reactions. A reaction process can be explained using an equation. For example, when oxygen reacts with magnesium ribbon, a white solid of magnesium oxide is formed. We have been using a word equation to describe such a reaction. Thus:



The (+) sign here is not used to mean addition, but in chemistry it is used to mean 'reacts with'. The  $\longrightarrow$  sign is used to indicate formation of a product. Using equal sign instead of an arrow, is wrong.

Names of starting substances like magnesium and oxygen in the above example are written on the left side of the arrow; these substances are called reactants. The new substances produced by the chemical reaction are called products and are written on the right side of the arrow.



Chemical equations have various notations that indicate the physical states of the reactants and products. These notations are very important. In fact, whenever one writes an equation and misses to write them, the equation is not complete. These notations are as follows:

Table 6.11 State symbols

Physical state	Representation of state	Description
Solid	(s)	A solid can be a precipitate or suspension
Liquid	(l)	A pure liquid like water and paraffin
Aqueous	(aq)	A solute or liquid dissolved in water
Gas	(g)	A gas or vapour

### Balancing chemical equations

In order for an equation to describe a reaction accurately, the equation must be balanced. Chemical reactions should always follow the law of conservation of mass so that the total mass of reactants must be equal to the total mass of products.

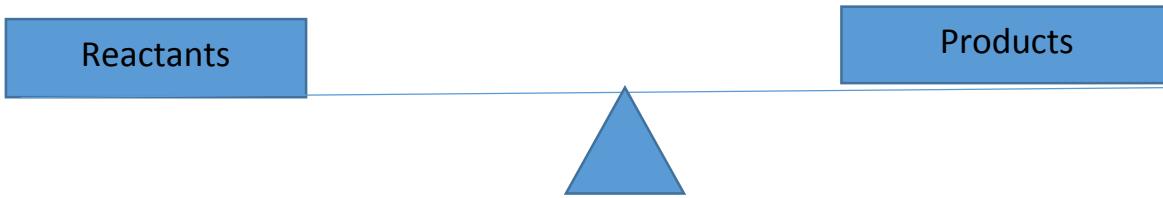


Fig 6.17 Conservation of mass in a balanced chemical reaction

Let us consider our previous example of magnesium ribbon reacting with oxygen. We can write an initial equation containing the formulae of reactants and products.



This equation is not balanced. Count the number of atoms of the reactants and the products.

Reactants	Products
1 magnesium atom	1 magnesium atom
2 oxygen atoms	1 oxygen atom

The left side (reactants) has two oxygen atoms but the right side (product) has only one oxygen atom.

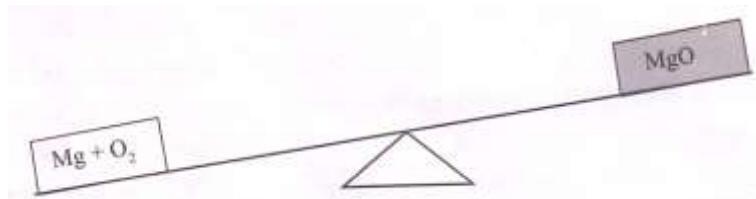


Fig. 6.18 Unbalanced equation, more oxygen atoms on the left

An equation can be balanced using a number of steps. Let us follow these steps to balance the above equation:

#### Step 1

Write the equation using correct formulae for the reactant and products.



#### Step 2

Count the number of atoms of each element in the reactants and in the products. Check whether they are equal as in Table 6.13.

Table 6.12 Balancing number of atoms for each element in the reactants and product

Atoms	Reactants	Product
Mg	1	1
O	2	1

We notice that the oxygen atoms are not equal. We have 2 atoms on the left and only one on the right.

### Step 3

To make oxygen atoms equal, balance the equation by writing numbers in front of the formula. Remember that 1 is assumed to be there already. Therefore start by inserting 2. If this does not balance, go to 3, 4 until the equation is balanced. Usually we do not go to very big numbers.



The number 2 now balances the equation. When you have  $2\text{O}_2$  it means two oxygen molecules. The number in front of a formula means, everything following is multiplied by that number.

### Example 6.3

$2\text{Mg}$  means  $2 \times \text{Mg}$  that is why we have 2 Mg atoms on the left and 2 Mg on the right. This equation  $2\text{Mg} + \text{O}_2 \longrightarrow 2\text{MgO}$  means 2 atoms of Mg react with 1 molecule of oxygen (containing 2 atoms) to form 2 molecules of magnesium oxide.

### Step 4

Count again the number of atoms of each element on the reactants and product side. Note that all atoms are balanced as illustrated in Table 6.14

Table 6.13 Balancing number of atoms for each element in the reactants and products

Atoms	Reactants	Product
Mg	2	2
O	2	2

### Step 5

Insert the correct state symbols for each substance.



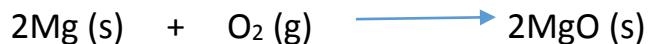
In a chemical reaction you are supposed to learn the reactants and products in words. Then change the words into correct formulae and finally balance the equation. Do not memorize chemical equation.

Examples

(a) Magnesium

When magnesium metal burnt in oxygen if formed a white ash called magnesium oxide. So we summarize the above statement in a word equation followed by a balanced chemical equation:

Magnesium + oxygen  $\longrightarrow$  magnesium oxide



Other metals that react with oxygen to from metal oxides include:

(a) Zinc + oxygen  $\longrightarrow$  zinc oxide



(b) Calcium + oxygen  $\longrightarrow$  calcium oxide



(c) Iron + oxygen  $\longrightarrow$  iron dioxide



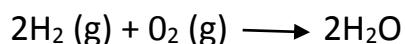
Non-metals combine with oxygen to from non-metal oxides for example charcoal (carbon), hydrogen, sulphur, phosphorus among others. Their chemical reactions are shown below:

(b) Carbon

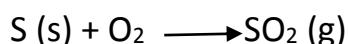




(c) Hydrogen



(d) Sulphur



### 6.3 Calculation of masses from equations

As we have suggested earlier, it is important to write correct balanced equations. When correctly written they show quantities that react and quantities produced. Given the relative atomic masses of the various elements in a compound, we add them up to get the mass of the formula. We can express the masses in grams.

Let us find the formula masses of the following compounds, using the relative atomic masses given in Table 6.15.

Table 6.14 Relative Atomic Masses of some compounds

Name of compound	Formula	Formula mass expressed in grams
Water	H <sub>2</sub> O	(2+16) = 18g
Magnesium oxide	MgO	24+16 = 40g
Sodium hydroxide	NaOH	23+16+1 = 40g
Calcium carbonate	CaCO <sub>3</sub>	40+12+ (3×16) = 100g

R.A.Ms H = 1, O = 16, Mg = 24, Na = 23, Ca = 40, C = 12

Calculation of mass of a reactant and mass of product

Balanced chemical reactions give the ratios of reacting masses of the products formed. From such information we are able to calculate the mass of a reactant used and also the mass of product formed.

Let us see how you can determine masses of products from a given chemical reaction.

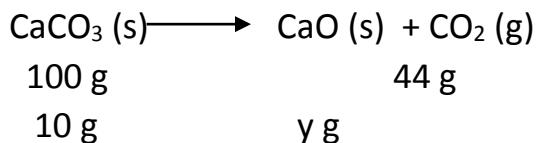
### Example 6.4

What mass of carbon dioxide will be given off when 10 g of calcium carbonate are heated? (Ca = 40, C = 12, O = 16)

Solution

- Write a balanced equation
- Write the reacting mass in grams below the formula

heat



- Cross multiply and solve for y

$$100 \times y = 10 \times 44$$

$$y = \frac{10 \times 44}{100}$$
$$= 4.4 \text{ g}$$

Therefore 10 g of calcium carbonate produce 4.4 g of carbon dioxide.

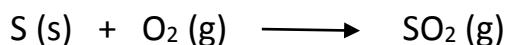
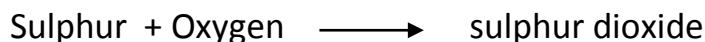
Note: We did not write the mass of Carbon dioxide ( $\text{CO}_2$ ) below the equation because it is not mentioned in the problem.

### Example 6.5

What mass of sulphur would be required to convert 15 g of oxygen to sulphur dioxide?

(S = 32.0, O = 16.0)

Solution



32 g    32 g

X g    15 g

$$X \times 32 = 32 \times 15$$

$$32x = \frac{32 \times 15}{32}$$

$$= 15 \text{ g}$$

### Example 6.6

How many grams of copper would be deposited if Ann added 6.5 g of zinc into copper (II) sulphate solution?

(Zn = 16, Cu = 63.5)

Solution



$$65 \text{ g} \qquad \qquad \qquad 63.5 \text{ g}$$

$$6.5 \text{ g} \qquad \qquad \qquad y \text{ g}$$

$$Y = \frac{6.5 \times 63.5}{65}$$

$$= 6.35 \text{ g}$$

### Experiment 6.4

Aim: To show that in a chemical reaction the mass of the reactants is equal to the mass of products

Apparatus and chemicals

- 250 cm<sup>3</sup> conical flask with rubber stopper
- Small test tube
- A piece of cotton thread
- Barium chloride (common salt) solution
- Dilute sulphuric acid
- Access to a balance

Procedure

1. Pour some barium chloride solution in a clean conical flask

2. Tie a piece of cotton thread tightly around the neck of a test tube and leave so that some thread hanging loose so that it hold the test tube as shown in Fig. 6.18
3. Add dilute sulphuric acid into the test tube until it is nearly full. Wipe the outside of the tube.
4. Lower the test tube carefully into the flask using the thread.
5. Push the rubber stopper into the flask so that it holds the thread firmly. The two solutions must not mix.
6. Weigh the flask and its contents and record its mass in a table like table mix.
7. Open the stopper so that the thread is free and test tube drops and the two solutions mix. Replace the stopper. What do you observe?
8. Again weigh the flask and its contents. Draw Table 6.15 in your notebook. Record the mass in the table.

Table 6.15 masses of substance before and after reaction

Mass of flask before reaction	x g
Mass of flask after reaction	y g

Are the value x and y equal or different?

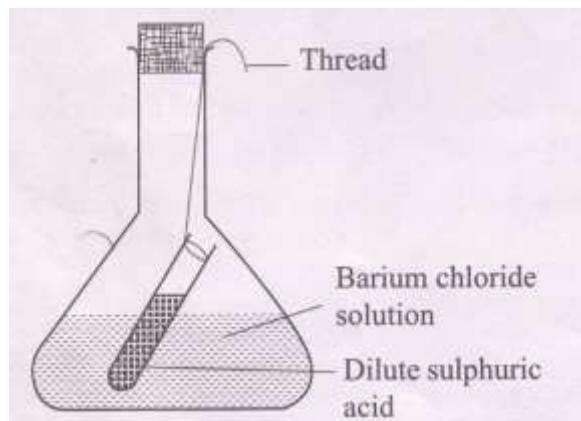


Fig. 6.19 To illustrate the law of conservation of mass

The solutions are colourless at the beginning of the experiment. But when mixed they form a white substance. This indicates that a reaction has taken place.

The values of x and y are equal. This shows the mass of reactants and products are equal. In other words nothing is lost or gained during reaction we say that matter is conserved. This shows a law known as h law of conservation of matter.

The law states that matter is neither created nor destroyed in a chemical reaction.

This explains why we must write balanced chemical equations for chemical reactions because the reactants are equal to the products.

Percentage composition by mass of elements in a compound

We have seen that to write a correct chemical formula we have to know the symbols of elements as well as their valencies. A correct formula gives us the ratio of the elements in that compound. For example, the correct formula of magnesium oxide is MgO, water is H<sub>2</sub>O and calcium carbonate CaCO<sub>3</sub>.

The last formula carries the following; 1 molecule of calcium carbonate contains; 1 atom of Calcium, 1 atom of Carbon and 3 atoms of oxygen. Study example 6.8 to learn how you calculate the parentage composition by mass of each atom in the formula.

### Example 6.7

Calculate the percentage composition of each element in calcium carbonate. (Ca = 40, C = 12, O = 16)

Solution

First write the correct formula of the compound and find the formula mass.

The formula mass of CaCO<sub>3</sub> = 40 + 12 + (16 × 3) = 100

Therefore,

$$\text{The \% of Ca} = \frac{\text{Ca} \times 100}{\text{CaCO}_3} = \frac{40 \times 100}{100} = 40\%$$

$$\text{The \% of C} = \frac{\text{C} \times 100}{\text{CaCO}_3} = \frac{12 \times 100}{100} = 12\%$$

$$\text{The \% of O} = \frac{O \times 100}{\text{CaCO}_3} = \frac{16 \times 3}{100} \times 100 = 48\%$$

## REVISION EXERCISE 6

1. Calculate the percentage by mass of each element in each of the following compounds:
  - (a) Water
  - (b) Calcium oxide
  - (c) Carbon dioxide
2. List difference between physical and chemical changes.
3. When Zinc oxide and lead dioxide are heated, they change colour when hot and when they cool. Is this a chemical or physical change? Explain
4. When a heap of an orange powder of substance X is touched with a hot wire, a large quantity of a green powder is formed, the mass glows red-hot and steam is given off. Is this a chemical or physical change? Explain your answer.
5. Balance the following equations:
  - (a)  $\text{Mg (s)} + \text{O}_2 \text{ (g)} \longrightarrow \text{MgO (s)}$
  - (b)  $\text{S (s)} + \text{O}_2 \text{ (g)} \longrightarrow \text{SO}_2$
6. How many grams of copper would be deposited if you added 13.0 g of zinc into copper (II) sulphate solution? ( $\text{Zn} = 65$ ,  $\text{Cu} = 64$ )  
$$\text{Zn (s)} + \text{CuSO}_4 \text{ (aq)} \longrightarrow \text{Cu (s)} + \text{ZnSO}_4 \text{ (aq)}$$

## **Unit 7**

## **Organic Compounds**

### **Meaning of organic compounds**

Organic compounds are studied under a branch of chemistry called organic chemistry. They are mainly known as hydrocarbons meaning they are compounds of carbon and hydrogen. In addition, they are known to link to form long chain carbon compounds. Other compounds such as carbon monoxide, carbon dioxide and those of carbonates and hydrogen carbonates are studied under the branch of chemistry known as inorganic chemistry. Although such compounds contain carbon atoms these atoms as we will see later in this course, do not link to form chains.

Therefore organic compounds are mainly hydrocarbons such as petrol, diesel, kerosene and liquid petroleum gas (LPG) etc. and their derivatives.

### **Natural sources of organic compounds**

#### **(a) Plants and animals**

These organisms synthesize many organic compounds. These include sugar, starches, fats, dyes, drugs among others.

#### **(b) Petroleum**

This is fossil fuel which originates from dead organisms such as algae buried under rocks and has undergone chemical changes over millions of years.

#### **(c) Natural gas**

This is a hydrocarbon found underground on top of crude oil deposit. It is also found under the sea. Natural gas is 95% methane.

#### **(d) Coal**

This is a fossil fuel formed from decayed plant materials that has been subjected to heat and pressure over millions of years.

### **Organic compounds as fuels**

A substance that undergoes combustion readily and gives out heat energy is known as a fuel. A fuel can be in form of a solid, liquid or gas.

## Substances used as fuels in homes

### (a) Solid fuel

- Charcoal
- Wood
- Coal

### (b) Liquid fuel

- Petrol
- Kerosene
- Diesel
- Fuel oil
- Alcohol (ethanol)
- Methylated spirit (ethanol mixed with methanol in spirit lamps and stoves)

### (c) Gas fuel

- Natural gas for example methane
- Liquid petroleum gas (LPG) for example propane and butane

### (d) Biogas

Methane is also a by-product of the metabolism of some living organisms. In developing countries, such as India and China, farmers build digesters. A digester is an apparatus that makes methane from waste organic matter, such as animal cow dung. In the absence of air, bacteria feed and give off methane. The local people use the gas for heating and lighting. Such biogas systems are also being investigated in developed countries.

They may provide a productive means of disposing of large amounts of waste from the intensive raising of livestock animals.

### (e) Liquefied alkanes

This is a convenient form of portable fuel supply. Butane, for example, is used in cigarette lighters camping stoves and lamps. Auto-mobiles and trucks can also be fuelled by liquid alkanes. Such alkanes are called liquid petroleum gas (LPG) in automotive applications.

Note: when we burn a fuel, the process is called combustion. Combustion of a substance requires oxygen. In actual fact, it is oxidation. During the process of combustion, energy is released. For example, when biogas (methane) is burned it produces carbon dioxide, water and heat energy.

### Composition and uses of fractions of petroleum

In unit 3, we saw that petroleum is a mixture of compounds which can be separated into fractions of different hydrocarbons by fractional distillation.

### Uses of fractions of petroleum

Some of the fractions and their uses are indicated in the following table:

Table 7.1: Fractions of petroleum and their uses

fraction	Use
• Liquefied Petroleum Gas (LPG)	• Fuel
• Petrol (gasoline)	• Powering vehicles
• Kerosene (paraffin)	• Jet fuel, heating and lighting
• Diesel	• Powering heavy vehicles for example lorries
• Lubricants (for example lubricating oil or grease)	• Reducing friction in moving parts of machines and vehicles
• Bitumen	• Tarmacking roads, repairing leaking roofs

### Revision Exercise 7

1. Define the following terms;
  - (a) Organic chemistry
  - (b) Organic compound
  - (c) Hydrocarbons
2. What is formed when a hydrocarbon is burnt in sufficient supply of air or oxygen?
3. What is the test for water and carbon dioxide?
4. Poisonous gas and soot are produced during incomplete combustion of fossil fuels. Give the chemical names of the two products.

