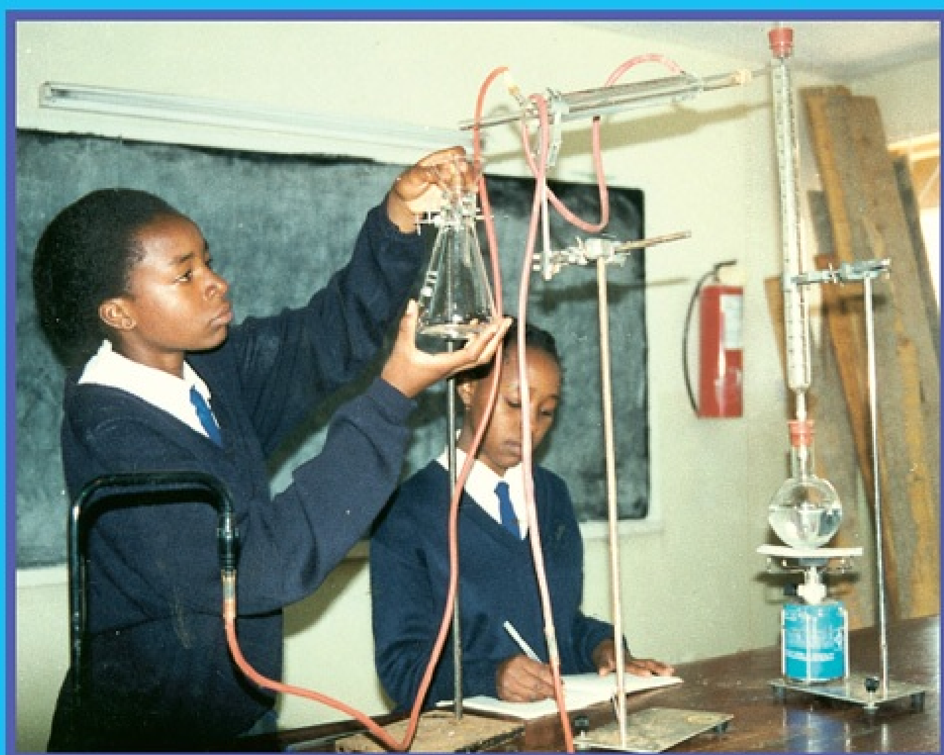
Fifth Edition

Secondary

CHEMISTRY

Form One Students' Book



KENYA LITERATURE BUREAU

Approved by the Ministry of Education

Secondary Chemistry

Form One

Students' Book

(Fifth Edition)

KENYA LITERATURE BUREAU
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Prologue

This student's book has been written to assist Secondary Chemistry Form One students. The book meets all the requirements of the current syllabus. It also conforms to current international trends in the teaching of the subject.

Chemistry is a practical subject which equips students with concepts and skills that come in handy in solving the day today problems in life. The subject aims at providing the learner with the necessary knowledge for individual benefit in daily life and for further education. This book recognises these aspects and provides adequate practical exercises to sharpen the student's practical skills. In addition, it also provides end of topic questions for self evaluation.

In this fifth edition, the book has been reworked and given a new easy to read lay out, the revision exercises fully address the syllabus requirements and there is a whole new section of sample examination style questions with answers.

I am grateful to the panel of writers and every body who took part in the writing, editing and production of this fifth edition of the book.

THE MANAGING DIRECTOR
Kenya Literature Bureau

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Acknowledgement

The Managing Director of Kenya Literature Bureau would like to thank the following writers who participated in the revision of this Fifth Edition of the Secondary Chemistry Form One Student's Book.

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CHAPTER ONE
INTRODUCTION TO CHEMISTRY

In Primary Science you studied science as a general course. Science enables us to understand our environment. At Secondary school level, science is divided into a number of subjects. The major science subjects are **Biology, Chemistry** and **Physics.**

Topics covered in primary science, which are studied under Chemistry include: Properties of matter, mixtures and methods of separation, drugs and pollution. In the secondary Chemistry course, we shall begin by reviewing some of those aspects of science already learned.

By the end of this chapter, you should be able to:

- Define Chemistry.
- Identify the topics studied in Primary School Science which are part of Chemistry.
- State the role of Chemistry in society and identify some applications of Chemistry in everyday life.
- Identify and state the use of common apparatus in the laboratory.
- Operate a Bunsen Burner.
- Observe safety in the laboratory.

Review of Chemistry topics learned in the primary science course

Matter

In our daily lives we encounter different substances. The air we breathe, the water we drink and soil are all forms of matter. *Matter is anything that has mass and occupies space*. Matter exists in three states namely solid, liquid and gas (vapour).

The three states of matter are interconvertible.

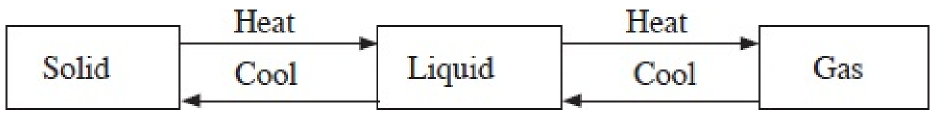


Fig 1.1 Relationship between the three states of matter.

Properties of matter

Experiment 1.1: What are some of the physical properties of solids?

Examine the shape of the solid provided. Place the solid on a weighing balance and record its mass. Half fill a glass tumbler and mark the water level. Place the solid in the tumbler, record your observation.

Answer the following questions

1. Describe the shape of the solid.
2. Is it possible to change the shape easily? Explain.
3. How does the mass of the solid you have used compare with the masses of those used by the other groups?
4. Explain the difference in the water levels before and after immersing the solid.

NOTE

Mass is the amount of matter in a substance and is measured in Kilograms. Mass remains constant everywhere. Weight is a product of mass and gravitational force. Weight is measured in newtons and varies from place to place.

Discussion

Solids have definite shapes that are not easily changed. Solids also have a definite mass and volume. Mass is a measure of the amount of matter in a substance. Volume is the space occupied by a substance. Different sizes of the same substance contain different amounts of matter and therefore have different masses. When the solid is put in water, the water level rises. The rise in water level represents the volume of the solid.

Experiment 1:2 What are some of the physical properties of liquids?

Weigh a 250ml beaker. Record the mass. Place 250ml of water in the beaker. Weigh and record the mass. Transfer the water into a 250ml measuring cylinder, 250ml conical flask, 250ml volumetric flask and then into a trough successively, each time, note the shape of the water in the container.

Answer the following questions

1. Determine the mass of water in the beaker.
2. Describe the shape of the water in each container.
3. State the volume of water in the trough.
4. Why was it possible to transfer the water from one container to another?

Discussion

When 250ml of water is added to the beaker and weighed, there is an increase in mass. The increase in mass represents the mass of water. This indicates that liquids have mass. When the 250ml of water is transferred into containers of different shapes, the volume remains the same while the water takes the shape of the container. It is possible to transfer liquids from one container to another by pouring because they flow. When the water is poured into containers of different shapes it flows and takes the shape of the new containers. Thus liquids have a definite mass and volume but take the shape of the containers in which they are placed due to their ability to flow.

Experiment 1.3: What are some of the physical properties of gases?

Weigh a deflated ball and record its mass. Inflate the ball using a hand pump. Observe and record what happens as the ball is being inflated. Detach the pump and weigh the inflated ball. Draw air into the pump. Block the nozzle of the pump with the thumb. Gently push the plunger in as far as it can go without letting any air to escape. While still blocking the nozzle, release the plunger. Record your observation.

Answer the following the questions.

1. Comment on the shape and size of the ball before and after it is inflated.
2. Determine the mass of air in the ball.
3. What is observed.
   (i) when the plunger is pushed in,
   (ii) when the plunger is released? Explain.

Discussion

As the ball is inflated, its size increases and it becomes spherical. The increase in size is due to the air occupying the space within the ball.

When air is pumped into the ball, the mass of the ball increases. This is due to the mass of the air being pumped into the ball. This shows that gases have mass.

When air is drawn in the pump and the nozzle blocked, the pump contains a fixed amount of air. When the plunger is pushed in, the volume of the fixed amount of air decreases. When the plunger is released, the air in the pump pushes out the plunger to occupy a large space. The experiment shows that a fixed volume of air can be compressed to occupy a smaller space. If allowed it can also spread to occupy a large space.

The above discussion shows that gases have a definite mass but do not have definite volume or shape.

Mixtures

Matter is found either as pure substances or mixtures. A pure substance is one that consists of only one type of matter.

A mixture consists of two or more substances mixed together and in which the individual components forming the mixture retain their physical and chemical properties.

A mixture can be separated by physical means such as winnowing, sieving, filtering, evaporating, decanting and use of magnets.

The choice of method to separate a given mixture depends on the nature and properties of the individual components forming the mixture.

Conductors and non conductors

In the primary science course, electricity was studied as one of the different forms of energy. Changes which take place in matter involve energy. Electricity is one form of energy which is important in the study of Chemistry. Substances which allow electrical energy to flow through them are **conductors.** Substances which do not allow electricity to flow through them are **non conductors.**

Drugs and drug abuse

In Chemistry, the word chemical is used to refer to the many substances that comprise matter. Chemists have been able to isolate some chemicals which are useful to the human body. Such chemicals are known as drugs. *A drug is any substance, natural or manufactured which when used alters the way the body functions*.

Tour officials wary of more dope scandals

NARBONNE, France

Tour de France organisers, adamant that they are winning the war against doping, will be praying for a scandal-free day during yesterday's 13th stage. On Thursday, Italian Riccardo Ricco fell by the roadside after failing a test for the blood-boosting drug EPO. The entire Saunier Duval team withdrew from the race.

Ricco, 24, who is the third rider to test positive for erythropoietin on this year's Tour, had won two stages on the three-week race.

Friday's 182 kms 13th stage from Narbonne to Nimes is flat from start to finish and expected to again favour sprinters.

Briton Mark Cavendish, winner of the three sprints on the Tour this season, will be closely watched by rivals still without a win, such as Spain's Oscar Freire, German Erik Zabel and Belgian Gert Steegmans. The last

Many sporting careers have been ruined by drug abuse.

Drugs used to treat diseases in human beings and other animals are known as medicines. Medicines are administered by qualified medical officers in specific amounts called **doses.** The written instructions by a qualified medical officer, giving details on the type of drugs and how the drugs should be used is called a **prescription.**

The use of a drug for a purpose other than what it is meant for, or use of overdose or underdose of prescribed drugs constitutes **drug abuse.** Drug abuse has harmful effects on the state of health of the user. The harmful effects include stress, depression, hallucination, addiction and dependency or may be fatal. The commonly abused drugs are tobacco, alcohol, bhang and khat (miraa). Harmful effects of smoking tobacco include lung cancer and heart failure. Alcohol abuse leads to liver problems (liver cirrhosis). Misuse of bhang leads to mental disorders. Prolonged use of khat leads to addiction, dependency and vascular disorders.

What is Chemistry?

States of matter and its properties, mixtures and their methods of separation and drugs are studied under **Chemistry.** *Chemistry is the study of the structure, properties and composition of matter and the changes that matter undergoes*.

The study of Chemistry involves carrying out experiments, making observations, analysis, interpretation and making conclusions.

Role of chemistry in society

People normally use water, soap, body lotions, perfumes, hair oils and plastic combs among other things. Their meals may include sugar, porridge, milk, tea, bread, vegetables and fruits.

Have you ever stopped to think about these things? How are they obtained? Who makes them? Have you ever wondered how life would be without the items mentioned above?

Many products avaiable thanks to Chemistry.

Some substances like water, milk and herbal medicines occur naturally. Others like soap, salt, panadol, chloroquine, body oils and cooking oils are prepared from naturally occurring materials. In Chemistry, substances are referred to as chemicals. The people who work with chemicals are Chemists.

Chemistry offers various career opportunities in various fields such as medicine, pharmacy, food technology, education and engineering. Chemistry has helped to improve standards of living in areas such as:

1. Manufacture of drugs to fight diseases.

2. Food production to fight hunger.
3. Manufacture of cheaper alternative fabrics such as nylon, polyester and tetron.
4. Manufacture of plastics for roofing, packaging and domestic use.
5. Manufacture of detergents.
6. Production of fuels for transport and domestic use. This includes alternative fuels to reduce global pollution as well as to supplement the fossil fuels.

NOTE

Fuel is any substance which burns to produce heat energy which is then used for different purposes.

Apparatus used in Chemistry

The study of Chemistry involves practical activities in form of experiments. The experiments are performed using chemicals and pieces of apparatus. Generally, these experiments are carried out in a **laboratory.**

A laboratory is a building or special room where chemicals and apparatus are kept and in which practical subjects such as Chemistry are studied. Some of the apparatus in the laboratory are used as sources of heat while others are used for the measurement of volume, temperature, mass and time.

Most laboratory apparatus which are used as containers and reaction vessels are made of transparent glass or plastic. This is to allow one to see through while observing the reactions taking place or to determine the level of the liquids held there in. Glass and plastic also do not react with most of the reagents used in the laboratory.

Apparatus for measuring volume

All apparatus used for measuring volumes of liquids are usually of transparent glass or plastic. The apparatus used for measuring volumes of liquids include a graduated beaker and flask, a measuring cylinder, volumetric flask, syringe, pipette and burette.

Each piece of apparatus is designed for a specific use and may come in various sizes. Figure 1.2 shows some of these apparatus.

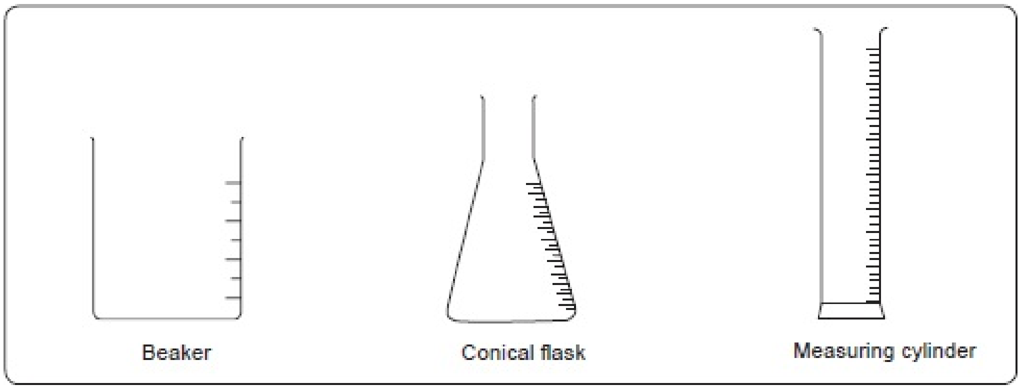


Fig 1.2a

Graduated beakers, flasks and measuring cylinders are used to measure approximate volumes of liquids.

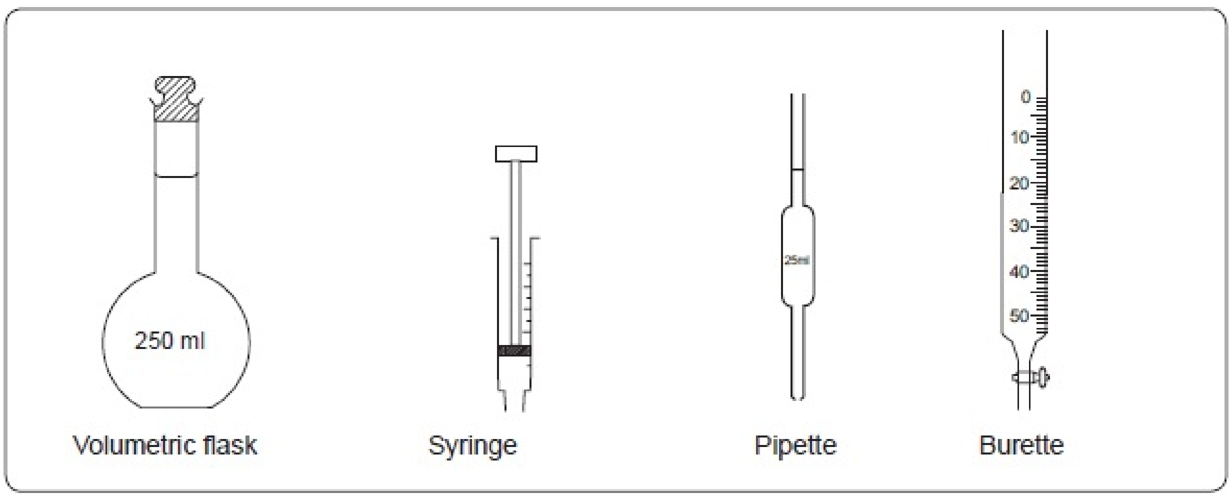


Fig 1.2b

When fairly accurate volumes are required, volumetric flasks, syringes, pipettes and burettes are used.

Apparatus for measuring temperature

Temperature is measured using thermometers. There are different types of thermometers such as maximum and minimum, clinical and general purpose thermometers. Figure 1.2 shows the general purpose thermometer used in Chemistry laboratories.

Fig 1.3 General purpose thermometer

Apparatus for measuring mass

Mass is measured using weighing balances. There are different types of weighing balances such as beam balances, electronic balances and top pan balances. Figure 1.4 shows electronic and beam balances.

An analytical balance with an accuracy of 10^{-3} g

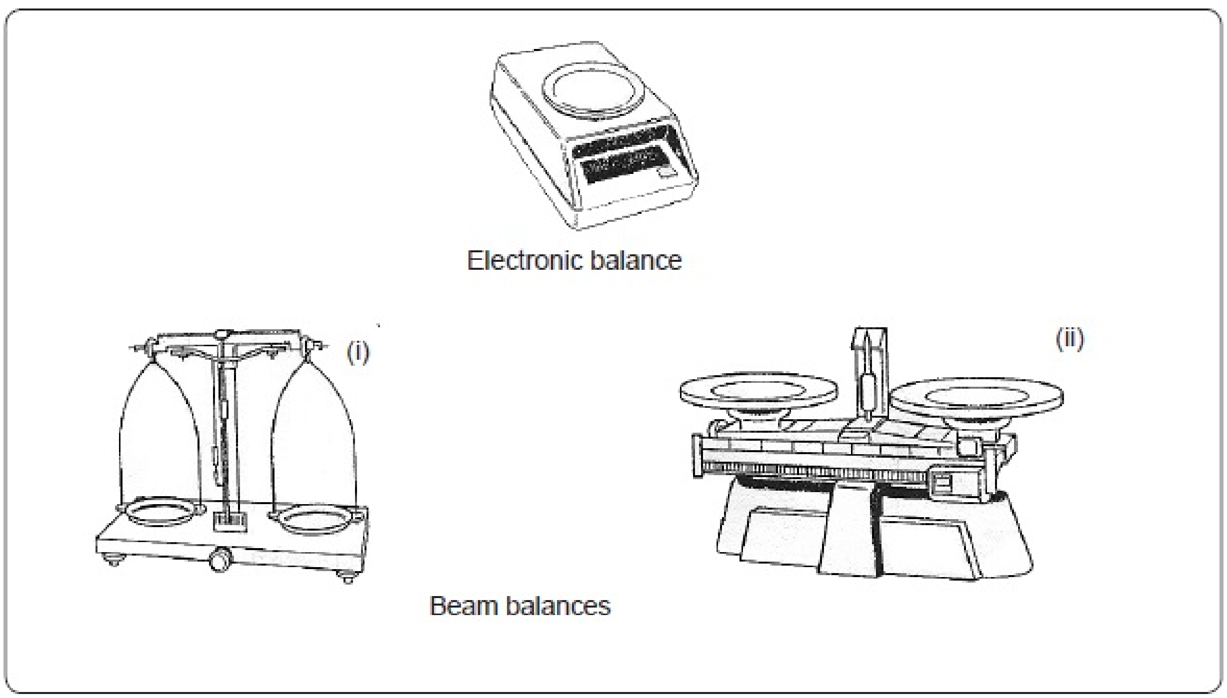
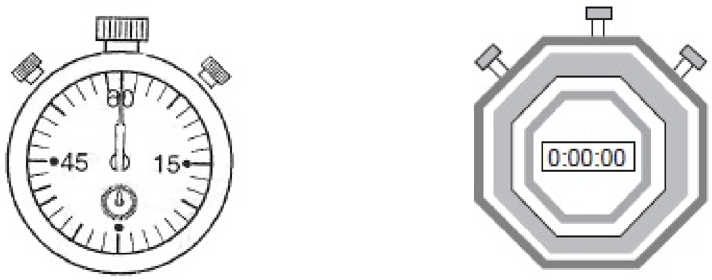


Fig 1.4 Common laboratory weighing machines

Apparatus for measuring time

The apparatus for measuring time are watches and clocks. For accuracy during experiments in the laboratory, stop watches and stop clocks are used. Some of the common types are shown in figure 1.5.

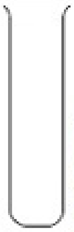


Analogue stop watch   Digital stop watch

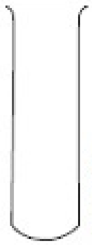
Fig 1.5 Apparatus for measuring time

Other commonly used apparatus are represented diagrammatically in figure 1.6.

Plastic, glass and ceramic are the most commonly used materials for Chemistry apparatus

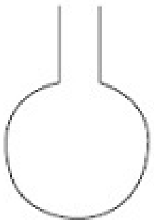


1. Test-tube
Used for general laboratory experiments

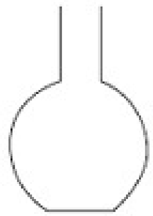


2. Boiling tube
Used when heating substances

3. Trough
Used for holding water for experiments especially when collecting gas over

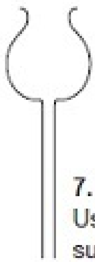


4. Round–bottomed flasks
Used when heating liquid susbtances because heat is supplied uniformly so that the flask does not crack as it expands

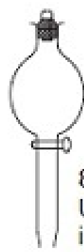


5. Flat-bottomed flask
Used for general laboratory experiments

6. Filter funnel
Used for delivering liquids carefully into vessels

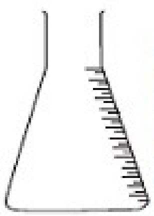


7. Thistle funnel
Used for delivering liquid substances



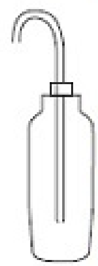
8. Separating funnel
Used for separating immiscible liquids

9. Dropping funnel
Used to add controlled amounts of liquids into reaction vessels

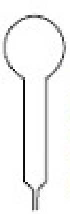


10. Conical flask
Used for general laboratory experiments and for measuring approximate volumes of liquids

11. Reagent bottle
Used for storing bench reagents

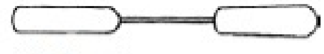


12. Wash bottle
Used to hold water for rinsing vessels and for adding water to vessels with narrow necks

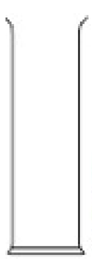


13. Teat pipette (dropper)
Used for delivering liquids drop-wise

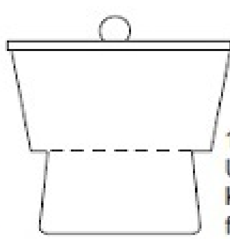
14 Evaporating dish
Used when evaporating liquids



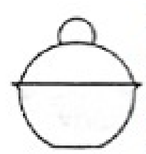
15. Spatula
Used for scooping solid substances from containers



16. Gas jar
Used for gas collection



17. Desiccator
Used for drying or keeping substances free from moisture



18. Crucible
Used when heating solid substances that require strong heating

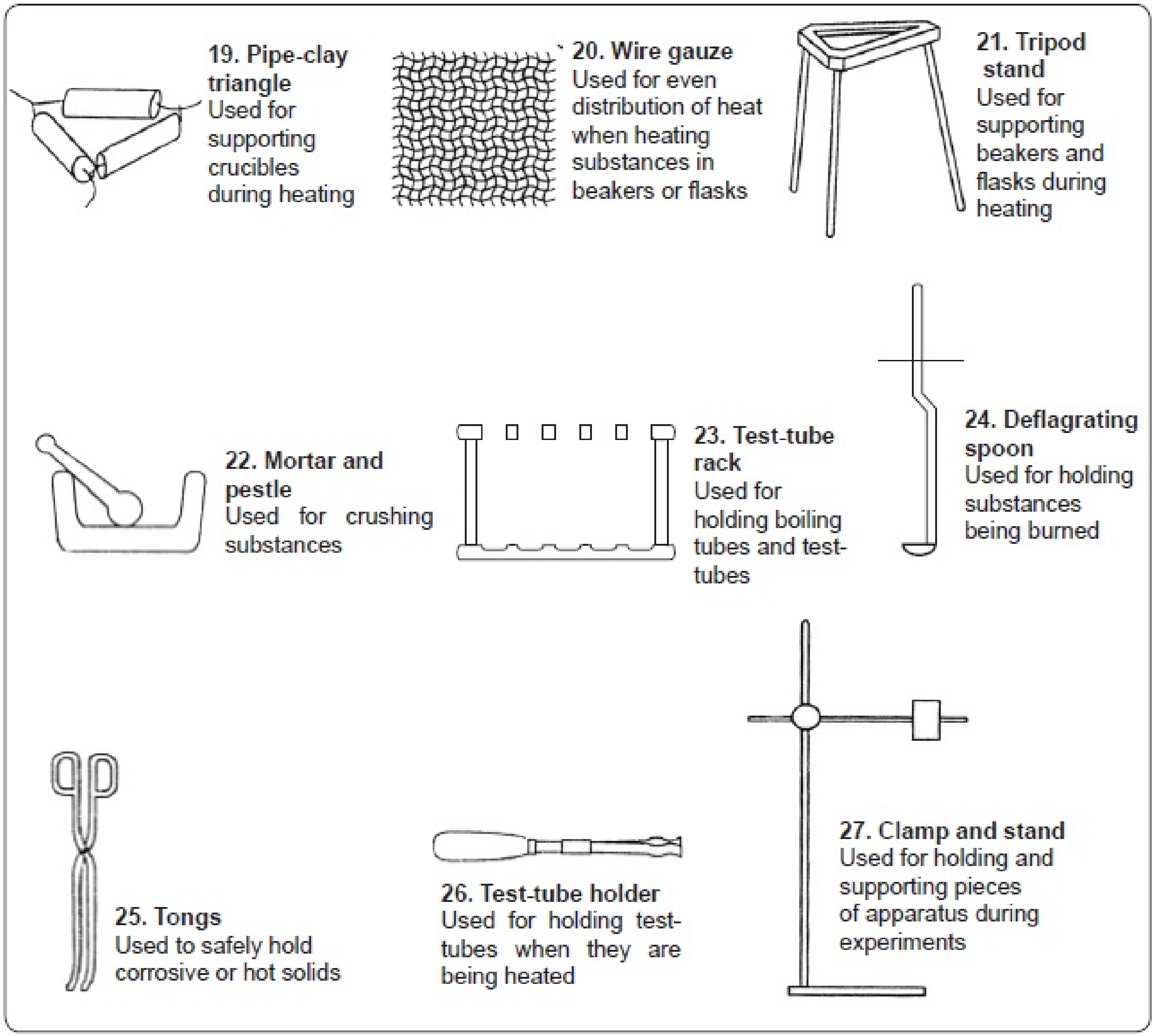


Fig 1.6

Sources of Heat

The pieces of apparatus used as sources of heat include the Bunsen burner, spirit lamp, candle, gas stove (portable burner), kerosene stove and electric heater. The Bunsen burner is the most suitable source of heat in laboratories.

The Bunsen Burner

A Bunsen burner consists of three major parts. These are the chimney, the collar and the base as shown in figure 1.7 (a). The chimney is a hollow metallic cylinder with an air hole near its lower end. The collar is a metal ring which may have an air hole whose diameter is the same size as that of the hole in the chimney. The diameter of the collar is slightly bigger than that of the chimney so that the chimney can just fit into it. The base is made of thick metallic material into which a small hollow metal with a jet is fitted. The Bunsen burner is normally connected to an external source of laboratory gas by rubber tubing. See figure 1.7 (b).

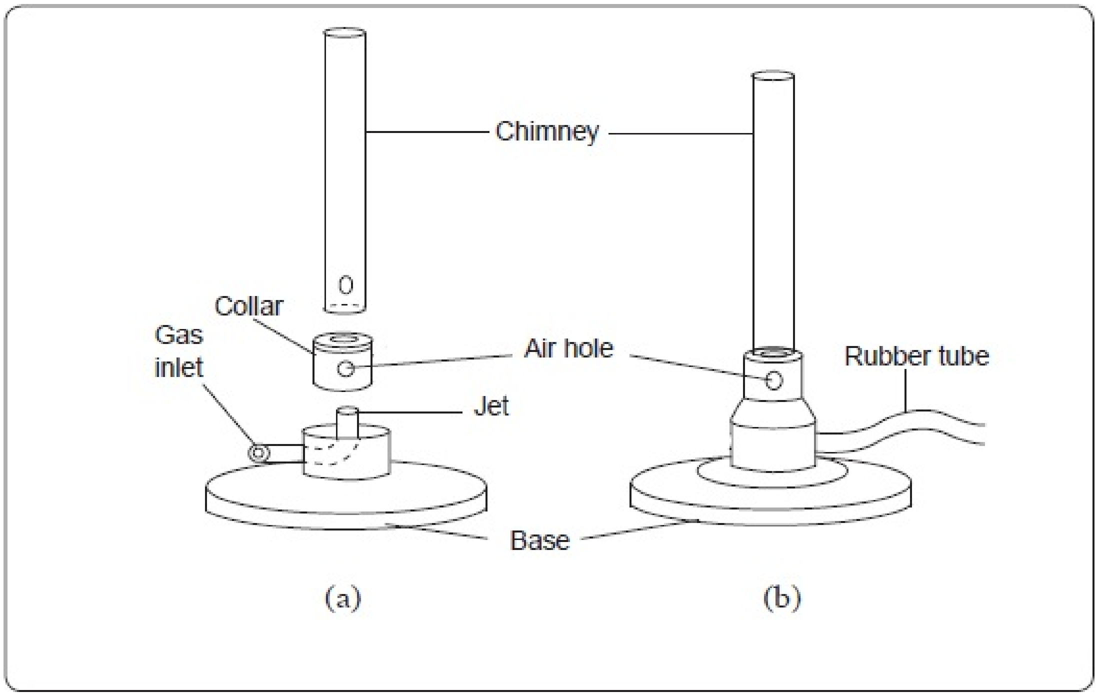


Fig 1.7: Parts of a Bunsen burner.

Functions of the different parts of a Bunsen burner

The gas inlet is connected to an external source of laboratory gas by a rubber tubing. The jet allows the laboratory gas into the chimney. The collar is used to regulate the amount of air entering the chimney. The air hole in the chimney allows air to enter and mix with the laboratory gas from the jet. This mixture of gases (laboratory gas and air), when ignited burn at the top of the chimney to produce a flame.

Experiment 1.4: What are the types of flames produced by a Bunsen burner?

Connect the burner to a gas tap and close the air hole. Turn on the gas fully and light the burner. Note the shape and colour of the different parts of the flame.

Slowly turn the collar until the air hole is half-way open and observe what happens to the flame. Continue turning the collar gradually until the hole is fully open. Observe the colours of the different parts of the resulting flame.

Draw and label the parts of the flame when the air hole is:

(a) closed.

(b) fully open.

Answer the following questions

1. What is a flame?
2. Suggest a reason why the two types of flames differ.

Discussion

A flame is a mass of burning gases. When in use, a Bunsen burner produces two types of flames depending on the amount of air allowed into the chimney as shown in figure 1.8.

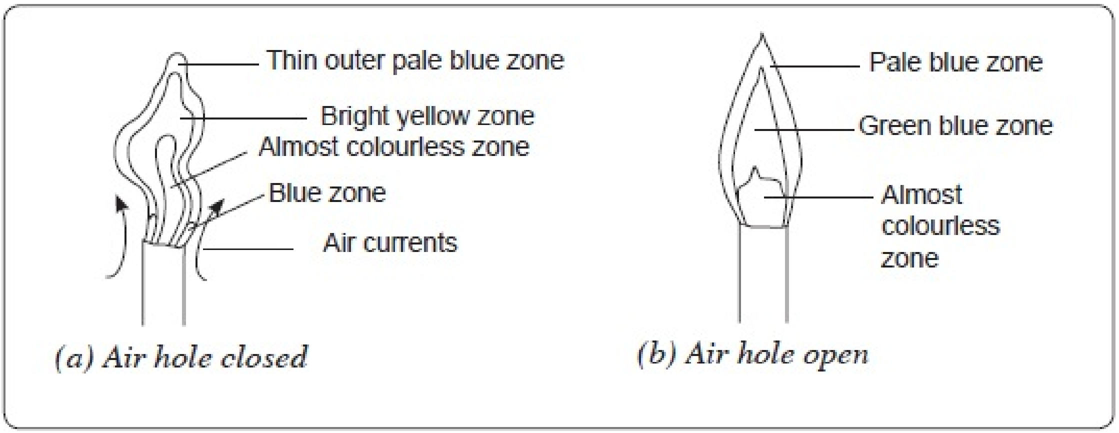


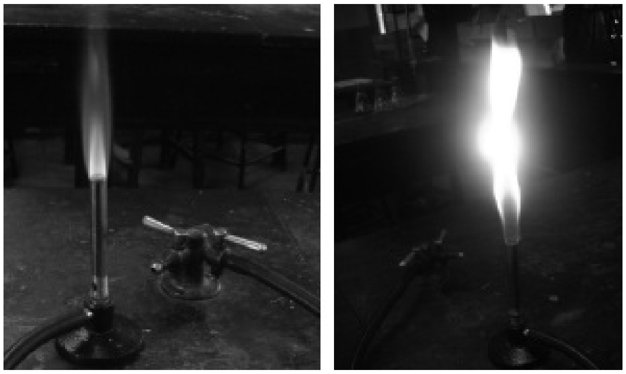
Fig 1.8 The flames of a Bunsen burner.

NOTE
The colour of a flame is related to its temperature. **A** white flame is the hottest followed by blue, red and yellow is the least hot.

When the air hole is closed, no air enters the chimney. The flame produced is bright yellow, large and wavy. It gives out much light and is described as a **luminous flame.** A luminous flame has four zones.

The blue zone occurs at the bottom of the flame. Air near the flame rises rapidly and mixes with the burning gas. This makes burning almost complete.

The almost colourless zone of the flame consists mainly of unburnt gases. The luminous bright yellow zone consists mainly of unburnt tiny particles of hot glowing solid carbon which give out light. The unburnt carbon particles form the black soot which makes apparatus dirty during heating. Air supply in the bright yellow zone is limited and there is incomplete combustion of the gas. In the thin outer pale blue zone, the gas burns completely because it mixes with plenty of air. However, this region is normally difficult to see.



Bunsen burner flames

NOTE
Candle and spirit burner flames are not efficient for heating because they are not hot enough.

When the air hole is slowly opened, more air enters the chimney. The bright yellow colour of the luminous flame gradually changes. When the air hole is fully opened, more air enters the chimney and mixes with the laboratory gas. The mixture of gases burn more quickly and completely. The flame obtained is pale blue in colour and is described as **non-luminous flame** because it does not give out much light.

The non-luminous flame has three zones. The almost colourless zone consists of unburnt gases. The green blue zone contains partially burnt gases due to insufficient supply of air. In the outer pale blue zone, the gases burn completely because there is plenty of air. The non-luminous flame gives out only a little light because it contains fewer hot carbon particles.

Experiment 1.5: What are the heating effects of the luminous and non-luminous flames?

Light a Bunsen burner and adjust the collar to produce a luminous flame. Pour 30cm^3 of water into 100ml glass beaker. Heat the water in the beaker and note the time it takes to boil.

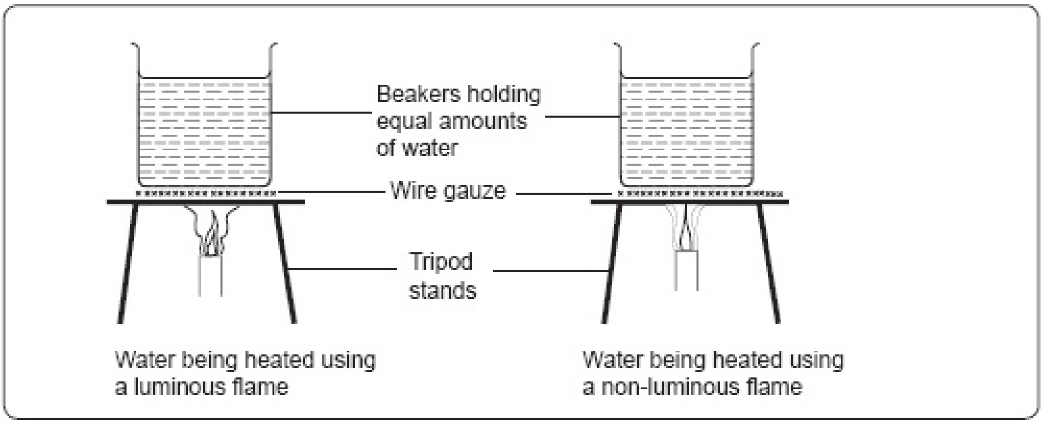


Fig 1.9 Using a Bunsen burner.

Repeat the experiment using 30cm^3 of water in an identical beaker and heat it with a non-luminous flame of the same Bunsen burner. In each case, observe the part of the beaker that was in contact with the flame.

Answer the following questions

1. Which water sample took a shorter time to boil?
2. What was observed at the bottom of each beaker?
3. Explain the observations you have made.

Discussion

Water heated by a non-luminous flame took a shorter time to boil than the same amount of water heated by a luminous flame. This shows that the non-luminous flame is hotter than the luminous flame. The bottom of the beaker heated using the luminous flame is covered with soot while the beaker heated using a non-luminous flame remains clean.

The non-luminous flame is normally preferred for heating substances. Luminous flames such as the candle flame and flame of lantern lamps are normally used for lighting.

Experiment 1.6: Which is the hottest part of a non-luminous flame?

Light a Bunsen burner and adjust the collar to obtain a non-luminous flame. Slip a piece of white manilla paper into the flame in region (i) as shown in figure 1.10 (a).

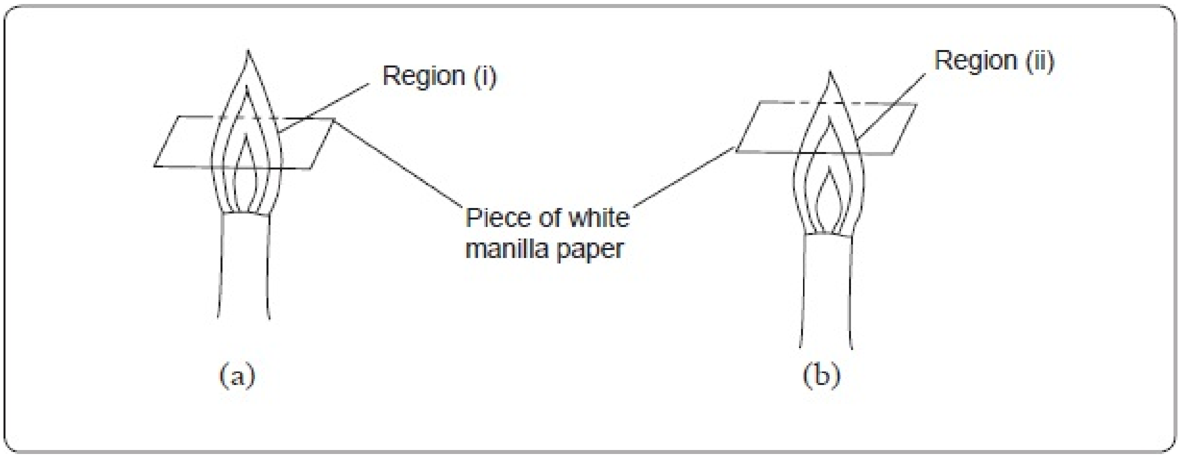


Fig 1.10: The heating power of different parts of the bunsen flame.

Quickly remove the paper before it catches fire. Slip a fresh piece of white manilla paper into region (ii) of the flame as shown in figure 1.0 (b). Quickly remove it before it catches fire.

Now repeat the experiments using wooden splints instead of white manilla paper. Let each splint stay in the flame until some of its parts get charred. Record your observations.

Draw diagrams to show how the pieces of paper and the splints were affected when placed in:

(a) region (i).

(b) region (ii) of the flame.

Answer the following question

Explain the observations made.

Discussion

When a piece of paper is quickly slipped in and out of region (i) of the non-luminous flame, the paper is partly burnt in the regions that are in contact with the pale blue zone. The middle part of the paper remains unburnt (no charring). When a piece of paper is slipped in and out of region (ii) of the non-luminous flame, it burns uniformly. The results of the experiments are shown in figures 1.11 (a) and (b).

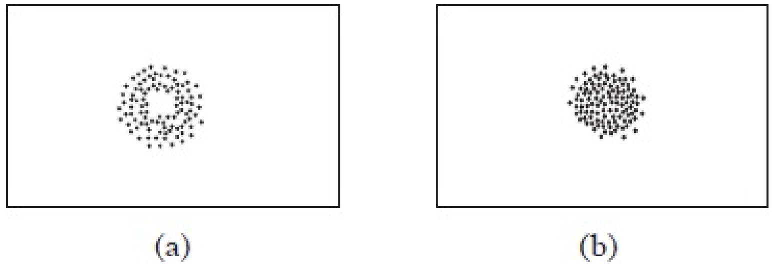


Fig 1.11 (a) and (b): Effect of different parts of the Bunsen flame on paper.

When wooden splints are used instead of paper, similar results are obtained as shown in figures 1.12 (a) and (b). These experiments show that the outer most zone in a non-luminous flame is hotter than the inner zones. An object being heated should therefore never be placed very close to the base of the flame where we have the cooler regions of the flame. Instead, the object should be placed at the outermost region of the flame. This is the hottest region of the flame.

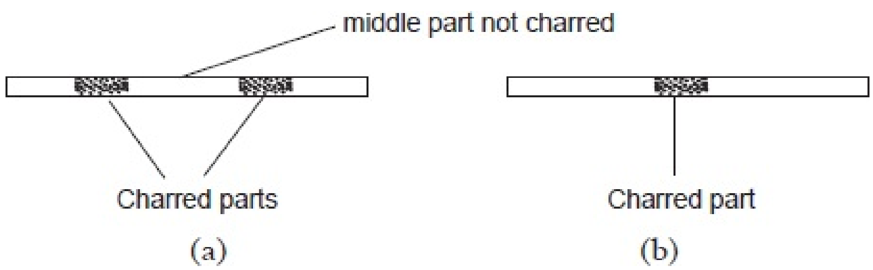


Fig 1.12 (a) and (b): Effects of different parts of the non-luminous flame on wooden splints.

Safety in the laboratory

When carrying out experiments, it should never be assumed to be completely harmless. Even the experiments that appear simple should be regarded as a potential source of danger. All chemicals and apparatus should be handled with care. For example, if a piece of glass apparatus is handled carelessly, it may break and injure the user.

Since learning chemistry emphasises on practical work, it is necessary that certain rules are followed to ensure safety in the Chemistry laboratory. The two common causes of accidents in the laboratory are ignorance and carelessness. Accidents are minimised when safety rules are followed.

All chemicals with environmental and health impacts must be stored in well labelled containers with appropriate safety warning symbols clearly visible.

A well labelled chemical container.

Laboratory safety rules

1. NEVER run while in the laboratory because you may trip, fall and injure yourself or other users of the laboratory.
2. NEVER taste or eat anything in the laboratory to avoid poisoning.
3. Always consult your teacher before trying any experiment to avoid accidents.
4. Label all the chemicals you are using to avoid confusion.
5. Always use a clean spatula for scooping a substance from a container to prevent contamination.
6. Always hold test-tubes or boiling tubes using a test-tube holder when heating to avoid being burned.
7. When heating a substance in a test tube or boiling tube, NEVER let the open end face you or anybody else because the liquid may spurt out and cause injury.
8. NEVER look directly into flasks and test-tubes where reactions are taking place, because the chemicals may spurt into your eyes and cause injury.
9. NEVER smell gases directly, instead waft the gas towards your nose with your hand.
10. Experiments in which poisonous gases are produced must be carried out in a fume cupboard or outdoors.
11. Always keep flammable substances away from flames because they easily catch fire.
12. Report any accidents to the teacher or the laboratory technician immediately

for necessary action.

13. In case of a serious accident such as fire, calmly walk out, DONT SCRAMBLE for the exit. Doing so would hinder easy escape.
14. Always extinguish flames that are not in use to avoid accidents and minimise fuel wastage.
15. If a chemical gets on your skin or in your mouth rinse it immediately with a lot of clean water.
16. Chemicals already used must always be disposed off safely to avoid contamination.
17. Always work on a clean bench. Clean all the pieces of apparatus used and store them at the end of each experiment.
18. Other safety regulations such as use of gloves , safety goggles, gas masks and **laboratory** coats should be adhered to as required.

Summary

1. Chemistry is the study of the structure, properties and composition of matter and the changes that matter undergoes.
2. Matter is anything that has mass and occupies space.
3. A drug is any substance, natural or manufactured which when used alters the way the body functions.

4 Drug abuse is the use of a drug for a purpose other than that for which it is meant.

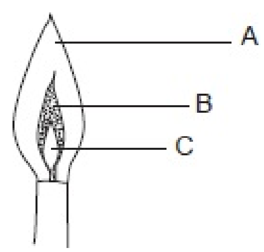
5. A laboratory is a special room or building where experiments are carried out.

6 Glass and plastic are prefered for most laboratory apparatus because they are unreactive.

7. Chemistry laboratory apparatus can be grouped according to use e.g. apparatus for measuring volume, mass, time and temperature.
8. A bunsen burner is the most appropriate source of heat for routine laboratory experiments.
9. Most accidents in the laboratory are due to ignorance and carelessness. Laboratory safety rules must be observed at all times.

Revision exercise

1. (a) Name three frequently abused drugs.
   (b) State two long-term effects of drug abuse.
2. Name four career opportunities open to a Chemist.
3. Explain why most laboratory apparatus are made of glass.
4. The following diagram represents a non-luminous flame of the Bunsen burner.



   (a) Name the parts of the flame labelled A, B and C.
   (b) Which of the parts in (a) above is the hottest?
   (c) A non-luminous flame is preferred for heating. Explain.
   (d) (i) Name the other type of flame produced by a Bunsen burner.
   (ii) Under what conditions does the Bunsen burner produce the flame you have named in d (i)?
5. (a) After use, a non-luminous flame should be put off or adjusted to a luminous flame. Explain.
   (b) Putting off flames after use is one of the safety rules in the laboratory. State five other rules.
6. State what students should do in case of a major accident such as fire outbreak in the Chemistry laboratory.

CHAPTER TWO
SIMPLE CLASSIFICATION OF SUBSTANCES

In our daily lives, substances such as sugar, salt, clean water and soap are used. These are final products obtained through processes such as separation from mixtures and purification. Indeed most other products undergo similar processes before they can be used.

In order to achieve this, substances have to be grouped together according to shared characteristics and properties. The process of grouping together substances with similar properties is called **classification.** Knowledge about properties can be applied to obtain pure substances from mixtures.

By the end of this chapter, you should be able to:
- Define mixture, element, atom, molecule, compound, melting point, boiling point and sublimation.
- Write chemical symbols of common elements and identify the constituents of compounds.
- Determine purity of a substance and use seperation techniques to obtain pure substances.
- Differentiate (i) permanent and non permanent changes (ii) physical and cchemical changes.
- Explain the arrangement of particles of matter in the three physical states in term of the kinetic theory of matter.

Mixtures

As discussed earlier matter can be classified into solids, liquids and gases. In nature, matter exists often as mixtures in various combinations. Fig. 2.1 is a flow chart representing the various categories of mixtures.

Separation of solid-solid mixtures

In primary school science, methods of separating solid-solid mixtures such as picking and winnowing were studied. The choice of each method depended on the size of the particles in the mixture. Size is a physical property. The method chosen to separate a given mixture depends on the physical properties of the components of the mixture. In this section, other methods of separating solid-solid mixtures will be studied.

Experiment 2.1: How can a mixture of iron filings and sulphur be separated?

Place the mixture provided on a piece of white paper and spread it out. Examine the mixture carefully and note its appearance. Hold a magnet above the mixture and observe what happens.

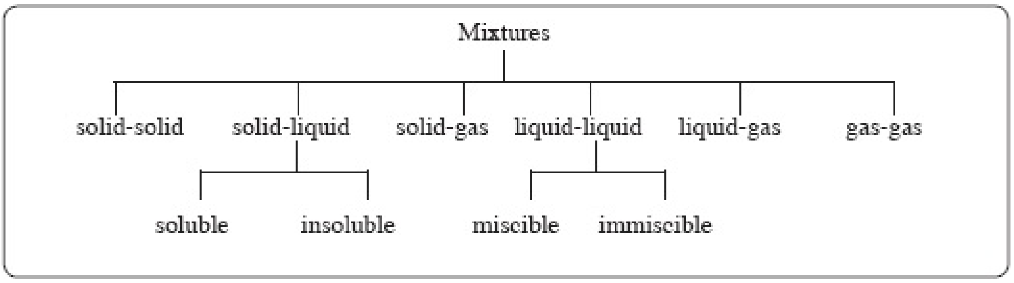


Fig 2.1: Table summarising types of mixtures.

Answer the following questions

1. Describe the colour of the mixture.
2. What is observed when a magnet is held above the mixture?
3. What is the colour of the solid;
   (a) attracted by the magnet?
   (b) left on the paper?
4. Which substance is:
   (a) sulphur?
   (b) iron? Explain

Discussion

The mixture provided is grey-yellow in colour. When a magnet is held above the mixture, the grey particles are attracted leaving a yellow powder on the paper. The grey substance attracted by the magnet is iron and the yellow substance left on the paper is sulphur.

A mixture of iron and sulphur can be separated by **use of a magnet** because iron is magnetic whereas sulphur is not. This method is applied in industries such as iron recycling, glass recycling and flour mills to remove iron particles.

NOTE
Recycling involves collecting and reprocessing used items into new products. Recycling saves on energy, money and keeps environment clean.

Experiment 2.2: How can a mixture of sodium chloride and iodine

be separated?

Examine the mixture provided and record your observation. Place a spatula full of the mixture in a 100ml glass beaker. Place the beaker on a tripod stand. Cover the beaker with a watch glass containing cold water. Heat the beaker gently until there is no further change. Carefully remove the watch glass from the beaker. Pour out the water from the watch glass and observe the lower surface. Examine the solid left in the beaker and record your observation.

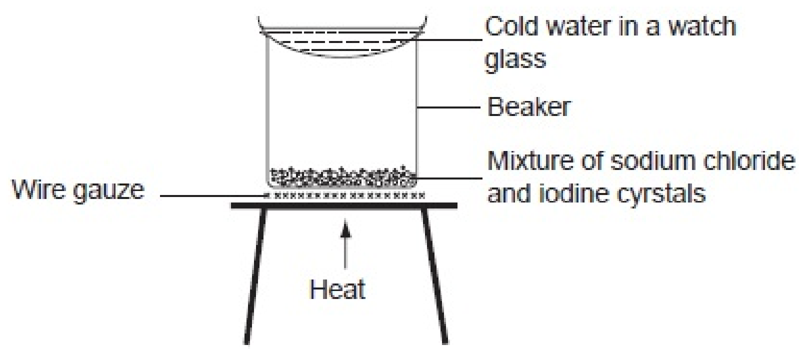


Fig 2.2: Separation of a mixture of iodine and sodium chloride.

Answer the following questions

1. Describe the appearance of the mixture.
2. What is observed inside the beaker as heating continues?
3. What is the colour of the:
   (a) solid left in the beaker?
   (b) solid formed on the surface of the watch glass?
4. Explain how the solid on the surface of the watch glass is formed.

CAUTION
Iodine vapour is harmful. Experiment should be conducted out–doors or in a well ventilated room.

NOTE
Sodium Chloride is soluble while iodine is slightly soluble in water. Therefore dissolving and evaporation would result into a sample of sodium chloride that has traces of iodine.

Discussion

The mixture contains white and shiny-black crystals. As the beaker is heated, a purple vapour is observed. On cooling, the purple vapour forms shiny black crystals on the surface of the watch glass. White crystals remain in the beaker. The shiny black crystals are iodine crystals whereas the white crystals are sodium chloride.

When solid iodine is heated, it changes directly into vapour (the purple vapour seen). The vapour cools to form solid iodine on the cold watch glass. The process where a substance changes from solid to vapour directly or vapour to solid without forming the liquid is known as **sublimation.** The solid formed when the vapour cools is known as a **sublimate.** It is possible to separate a mixture of iodine and sodium chloride because iodine sublimes while sodium chloride does not.

Other substances that sublime are anhydrous iron (III) chloride, aluminium chloride, benzoic acid and carbon (IV) oxide (dry ice).

"Dry ice" Pedal cart.

Application

Dry ice is used in cold boxes by ice cream vendors. Dry ice is preferred over ordinary ice because it sublimes leaving no wetness. It is also a better coolant compared to ordinary ice.

Solid-liquid mixtures

There are two categories of solid-liquid mixtures. In one category, the solid dissolves in the liquid and is said to be **soluble.** In the other category the solid does not dissolve **(insoluble)** resulting to an insoluble solid-liquid mixture.

Experiment 2. 3: What happens when solids are mixed with liquids?

Pour about $10cm^3$ of water into a test-tube. Add to it half a spatulaful of sodium chloride and shake well. Repeat the experiment with sugar, sand, sulphur, oxalic acid crystals, potassium nitrate, iodine and naphthalene respectively. Repeat the experiment using $10cm^3$ of propanone (acetone) as the liquid in place of water. Record your results in a table 2.1. indicating whether soluble or insoluble.

CAUTION
Propanone should not be handled in the presence of any flames because it is flammable (catches fire easily).

Table 2.1: Solubility of different substances in water and propanone

Substance	*Solubility in*	
	Water	*Propanone*
Sugar		
Sodium chloride		

Answer the following questions

1. From the experiment,
   (a) Which substance dissolved in;
      (i) water?
      (ii) propanone?
   (b) Which substances did not dissolve in;
      (i) water?
      (ii) propanone?
2. Which substances dissolved in both water and propanone?

Discussion

When each substance is mixed with water and propanone separately, it is found that sodium chloride, sugar, potassium nitrate, and oxalic acid crystals dissolve in water. Naphthalene, sugar and oxalic acid dissolve in propanone. Oxalic acid and sugar dissolve in both water and propanone. Naphthalene, sand and sulphur did not dissolve in water. Sodium chloride, potassium nitrate, sand and sulphur did not dissolve in propanone. Sand and sulphur did not dissolve in both water and propanone. Substances that dissolve in a liquid are said to be **soluble,** while substances that do not dissolve are **insoluble.** When a substance dissolves in a liquid, the substance is called a **solute** and the liquid is called a **solvent.** The resulting mixture is called a **solution.** When the solution is stirred, it forms a **homogeneous** mixture.

NOTE
Homogenous mixture is a mixture in which the solute and solvent particles are evenly distributed.

Separation of insoluble solid-liquid mixtures

Experiment 2.4 (a): How can sand be separated from water-sand mixture?

Transfer $10cm^3$ of the mixture provided into 100ml beaker. Allow the mixture to settle and carefully pour the liquid into another beaker. See figure 2.3. Preserve the contents of beaker 2 for experiment 2.4 (b).

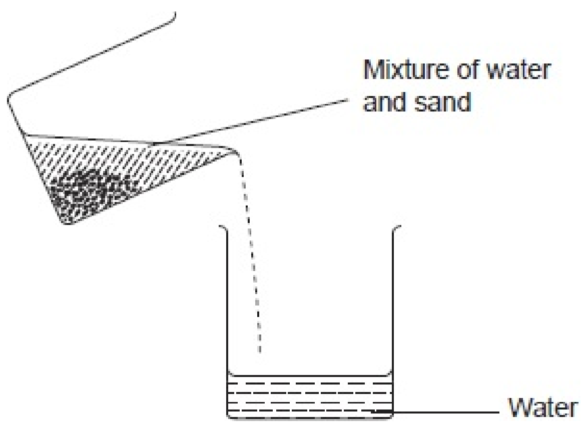


Fig 2.3: Decantation

Answer the following questions

1. How clear is the water collected in the second beaker?
2. Why is the method not good in separating such a mixture?

Discussion

Sand is insoluble in water. When a mixture of sand and water is allowed to stand, the sand settles at the bottom. The water can then be poured off carefully. This method of separation is called **decantation.** It is used in separating insoluble solids from liquids. However, it is not efficient because some solids still pass into the liquid in the process of decanting. That is why the water collected in the second beaker is not clear. It still contains small suspended particles.

Experiment 2.4 (b): Which other method can be used to separate sand from water?

Fold a filter paper into a cone shape as shown in figure 2.4 (a). Place the filter paper in the funnel. Pour the mixture from experiment 2.4 (a) carefully on the filter paper and collect the liquid in a conical flask as shown in figure 2.4 (b).

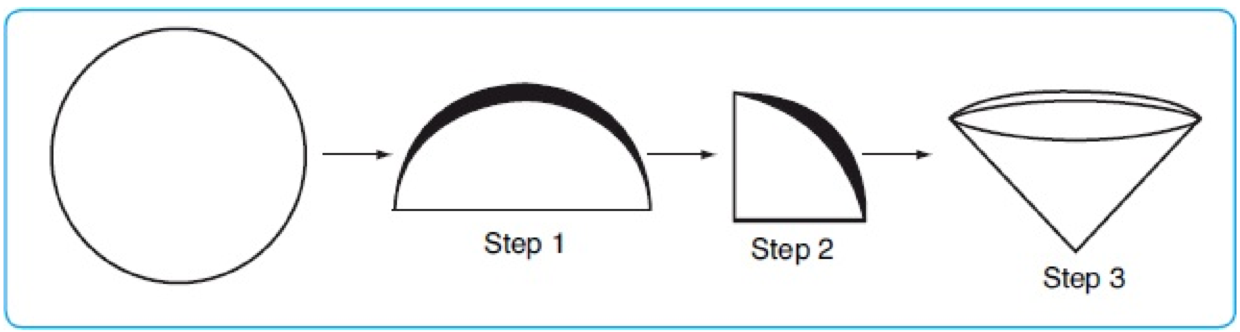


Fig 2.4 (a): How to fold a filter paper

Answer the following questions

1. What do you observe at the end of the experiment?
2. State an advantage of this method over decantation.

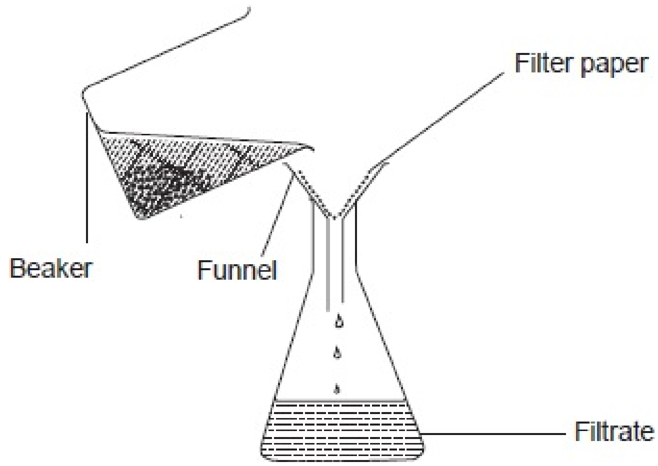
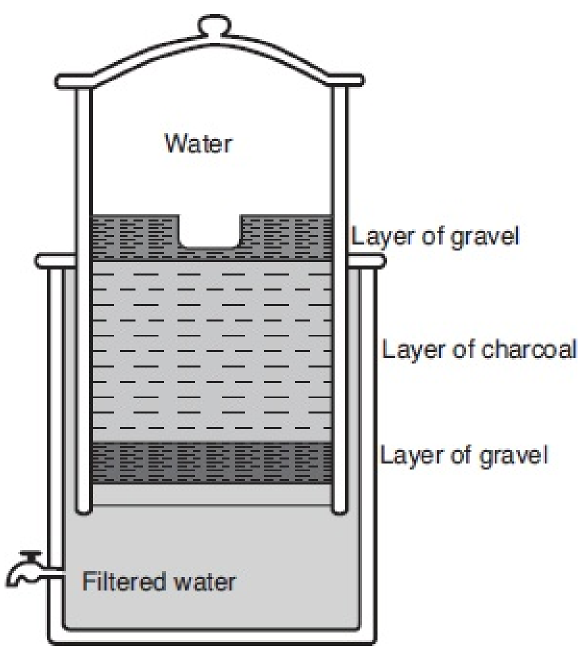


Fig 2.4 (b): Filtration

Discussion

Fine sand particles are trapped and collect on the filter paper. The solid particles that are left on the filter paper are called a **residue.**

Water collected in the conical flask is clear because the filter paper does not allow solid particles to pass through. The liquid collected in the conical flask is called a **filtrate.** This method of separation is known as **filtration.**



A Domestic water filter.

Filtration is used on a large scale in water purification plants. Dirty water is allowed to pass through a filter bed made of layers of gravel and sand.

Suspended particles are trapped by the gravel and sand while the water passes through. Domestic water filters use the same principle.

Separation of soluble solid-liquid mixtures

Experiment 2.5: How can sodium chloride be obtained from a mixture of sodium chloride and water?

Put $10cm^3$ of sodium chloride solution provided into an evaporating basin. Heat the evaporating basin until crystals start forming.

Transfer the evaporating basin on to a water bath as shown in figure 2.5 and continue heating to dry the crystals. Once dry, allow the crystals to cool.

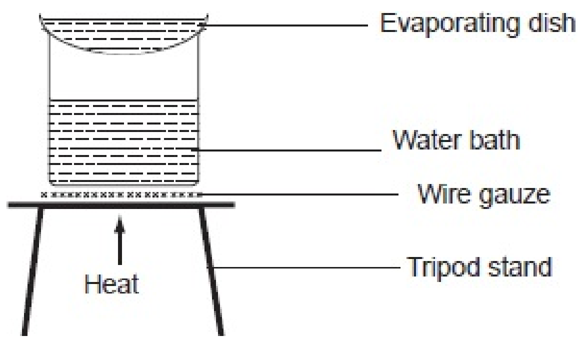


Fig 2.5 Evaporation using a water bath.

NOTE
A sand bath may be used in place of a water bath but it takes long to be heated up. However it retains heat much longer than the water bath.

Answer the following questions

1. Why was the evaporating basin containing the mixture heated?
2. Why was the evaporating basin transferred to the water bath?

Discussion

Sodium chloride (common salt) is soluble in water therefore cannot be separated by decantation or filtration. The evaporating basin is heated to drive away water into the atmosphere. This process is known as **evaporation.** When crystals start forming the evaporating basin is transferred to a water bath so that the salt does not spit out of the basin as heating continues. This process is used to obtain salt from sea water.

Experiment 2.6: How can a mixture of sand and sodium chloride be separated?

Put a mixture of sand and sodium chloride in a beaker. Add water to it and warm while stirring. Allow the mixture to cool.

Filter the mixture using a filter paper and collect the filtrate on an evaporating basin. Evaporate the water until crystals start forming. Allow the solution to cool for more crystals to form.

NOTE
Filtration is used to separate an insoluble solid from a solution. Evaporation is used to separate a solute from a solvent.

Answer the following questions

1. What was the residue left on the filter paper?
2. Name the substances contained in the filtrate.
3. Explain how sodium chloride would be recovered from the filtrate?
4. Which methods were used to separate the mixture?

Discussion

Sand is insoluble in water and is separated by filtration. It is collected as the residue on the filter paper. Sodium chloride is soluble in water and is contained in the filtrate. Therefore, to recover the sodium chloride, evaporation is carried out. The hot concentrated filtrate is then cooled to allow more crystals of sodium chloride to form.

Both **filtration** and **evaporation** methods were used in this experiment to separate the mixture.

Experiment 2.7: How can copper (II) sulphate crystals be obtained from copper (II) sulphate solution?

Put $10cm^3$ of concentrated copper (II) sulphate solution into an evaporating basin. Arrange the apparatus as shown in figure 2.6.

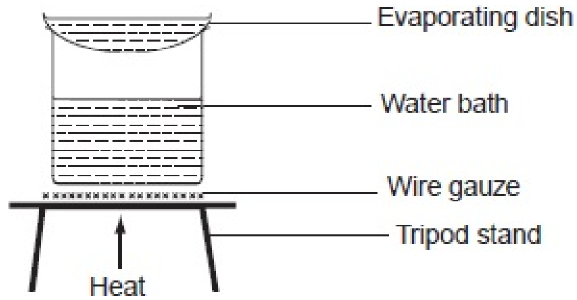


Fig 2.6: Obtaining crystals from copper (II) sulphate solution.

NOTE
For large crystals of copper (II) sulphate to form, the evaporation process must be slow.

Using a water bath. Heat the solution to evaporate excess water. As heating continues, dip a glass rod into the solution regularly and allow it to cool in the air. When crystals start forming on the glass rod, remove the evaporating basin from the water bath and allow it to cool. Record the observations.

Answer the following questions

1. Why is the glass rod dipped in the solution while heating continues?
2. What observations are made when the solution cools?

Discussion

The glass rod is dipped into the solution to find out whether the solution can form crystals on cooling. When crystals form on the glass rod, this is an indication that the solution is ready to form crystals. At this point the solution is said to be **saturated.** *A saturated solution is one in which no more solute can dissolve at a given temperature*. The process of obtaining crystals from a saturated solution is known as **crystallisation.** This method can be used to separate most soluble substances from their solutions.

A worker at a Salt extraction pan.

Application

1. Extraction of salt from salty water e.g Lake Magadi and Ngomeni in Malindi.
2. Extraction of sugar from sugar cane.
3. Extraction of medicinal substances from plants.

Experiment 2.8: How can a solvent be obtained from a solution?

Place about 20cm^3 of sodium chloride solution in a boiling tube or a round-bottomed flask. Arrange the apparatus as shown in figure 2.7. Heat the solution until all the solvent has evaporated. Record your observation.

Place about 20cm^3 of the sodium chloride solution on a watch glass. Place the watch glass on a water bath. Heat the sodium chloride solution to dryness. Remove the watch glass from the water bath and allow it to cool. Record your observations.

Place the liquid collected in tube B on another watch glass. Place the watch glass on the water bath. Heat the liquid to dryness. Allow the watch glass to cool. Record your observation.

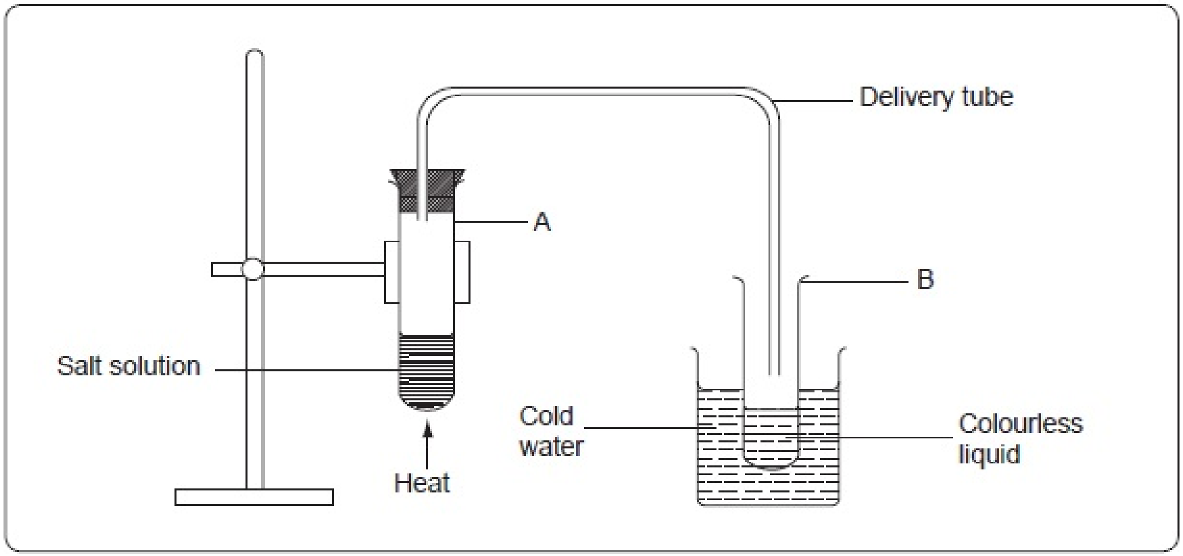


Fig. 2.7: Set- up for simple distillation.

Answer the following questions

1. What is observed in the delivery tube as the solution boils?
2. What is the purpose of cold water in the beaker?
3. What is observed in test state tube A after all the solvent has evaporated.
4. What is observed on the watch glass containing:
   (i) sodium chloride solution?
   (ii) liquid from test state tube B when heated to dryness? Explain.

Discussion

The steam passes through the delivery tube where some of it condenses on the cooler parts of the delivery tube to a colourless

liquid. The steam is cooled and condenses to liquid water in test state tube B. The cold water in the beaker therefore acts as a coolant.

A white solid remains in tube A. The white solid is sodium chloride. This method of separating a solute from a solvent is called **simple distillation.** The liquid collected in tube B is called a **distillate.**

NOTE
Distilation is the process of evaporating a liquid from a solution and condensing the vapour produced back into liquid.

When sodium chloride solution is placed on a watch glass and evaporated to dryness, a white residue is left behind. When the distillate is placed on a watch glass and evaporated, no residue remains. This shows that the distillate no longer contains dissolved sodium chloride.

To improve the condensation process, a **Liebig condenser** is used as shown in figure 2.8.

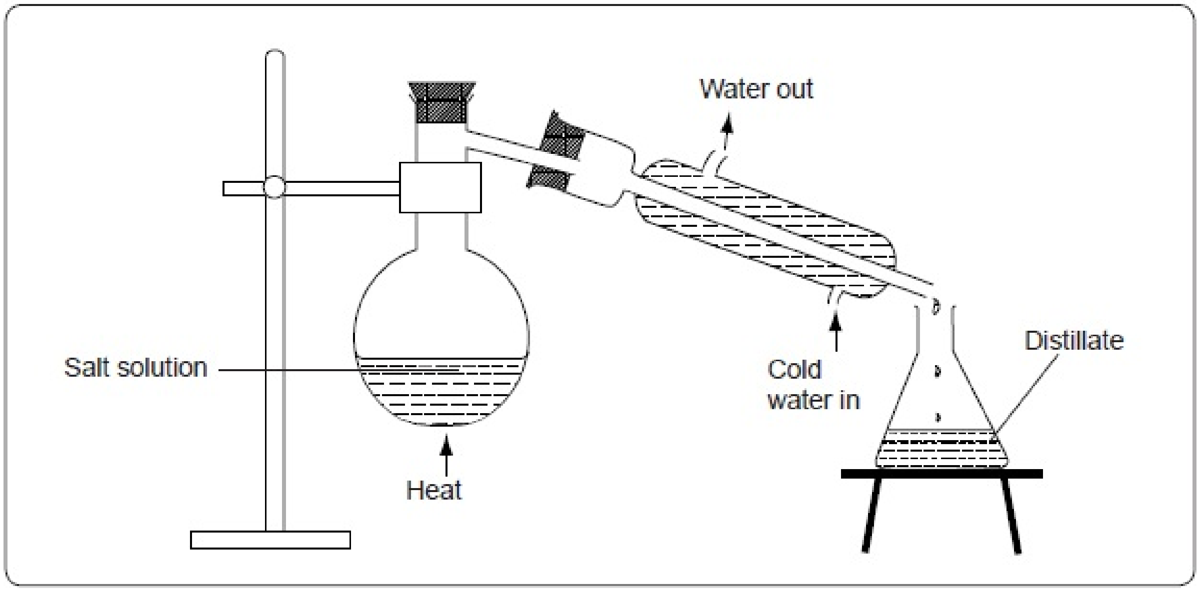


Fig 2.8: Simple distillation with Liebig condenser.

Liquid-liquid mixtures

There are two types of liquid-liquid mixtures namely immiscible and miscible liquids.

Immiscible liquids do not mix but form distinct layers. Examples of such mixtures are vegetable oil and water, kerosene and water. Miscible liquids mix to form a homogenous mixture. Examples are water and ethanol, water and milk.

Separating Immiscible Liquids

Experiment 2.9: How can immiscible liquids be separated?

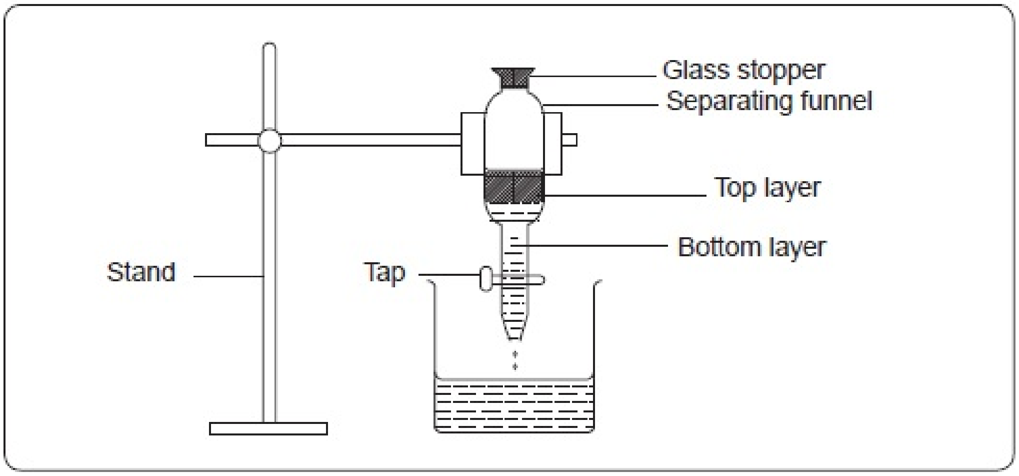


Fig 2.9 (a): The separating funnel in use.

Put 10cm^3 of coloured water in a conical flask. Add 10cm^3 of kerosene and shake well. Transfer the mixture into a separating funnel as shown in figure 2.9 (a). Allow the mixture to stand for a while and observe. Remove the stopper. Open the tap of the separating funnel and allow the bottom layer to flow into a beaker. Discard the interphase leaving the top layer in the separating funnel.

Answer the following questions

1. Which layer is at the top in the separating funnel? Explain
2. Explain why the interphase was discarded.
3. Which other method could be used to separate the two layers?

NOTE
Interphase is the boundary between water and kerosene.

Discussion

A mixture of water and kerosene is immiscible. The water is coloured to distinguish it from kerosene.

Kerosene floats on water because it is lighter. The interphase contains both water and Kerosene. It is discarded because it is not easy to separate the two liquids at the interphase. Immiscible liquids are separated by use of **a separating funnel.** Decantation can also be used to separate immiscible liquids but it is not efficient. See figure 2.9 (b)

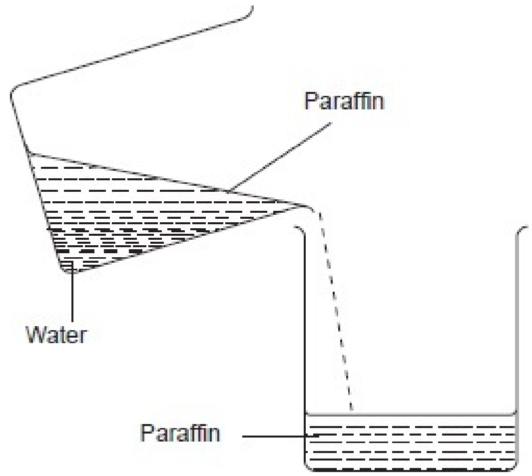


Fig 2.9 (b): Decanting in liquids.

NOTE
A burette may be used to separate the immiscible liquids.

A dropper can also be used to separate immiscible liquids by sucking the upper layer. The dropper is used to suck one layer transferring it to another beaker repeatedly. This method too is not accurate. See figure 2.9 (c).

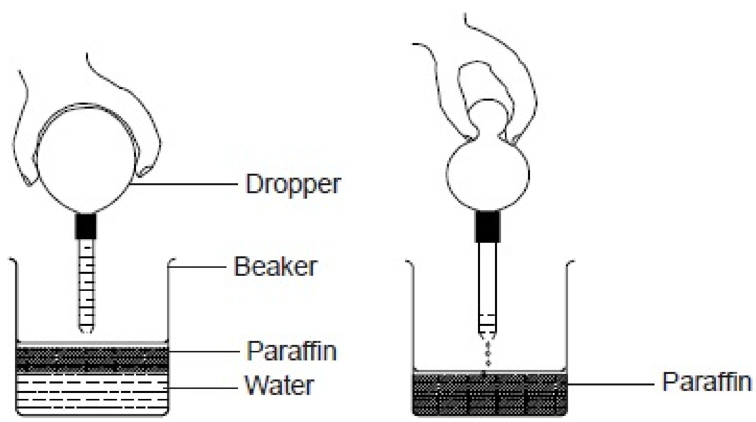


Fig 2.9 (c): Use of the teat pipette in separation of immiscible liquids

Separating Miscible Liquids

Experiment 2.10: How can a mixture of water and ethanol be separated?

Place 20cm^3 of ethanol on a watch glass and ignite. Record your observations. Place 20cm^3 of water in a round-bottomed flask. Add 20cm^3 of ethanol and shake the mixture. Set up the apparatus as shown in figure 2.10.

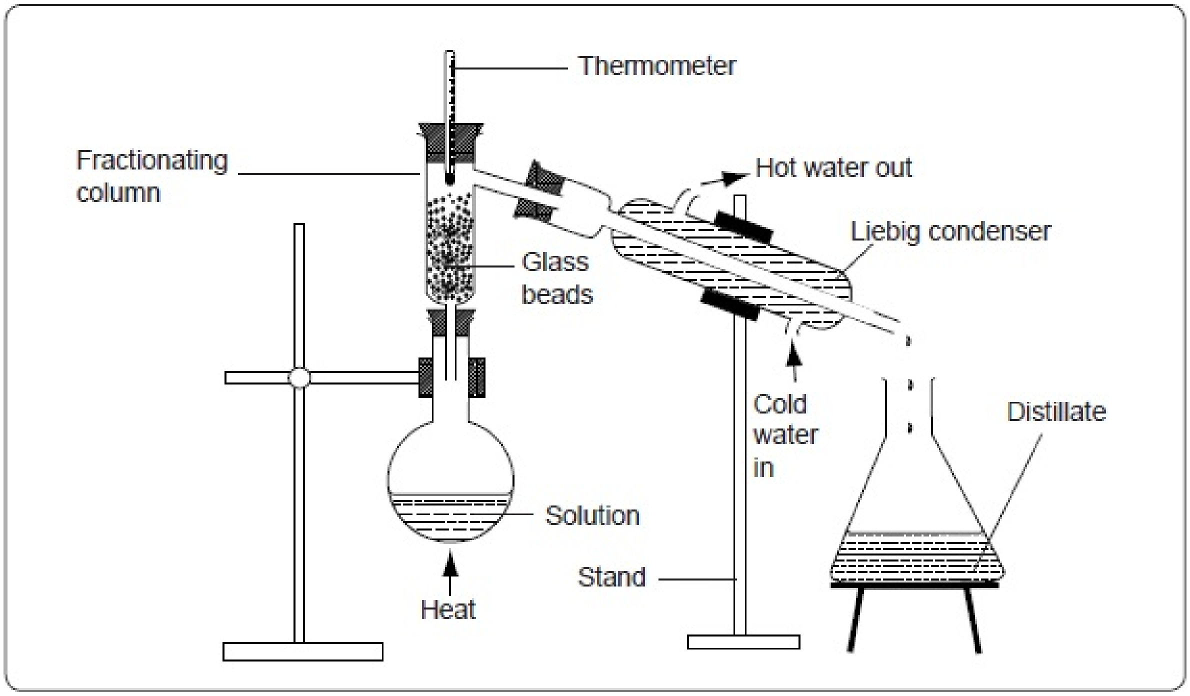


Fig 2.10: Separating miscible liquids.

Heat the mixture gently and note the temperature changes. Collect the distillate in the conical flask. Transfer a little distillate on a watch glass and light it. Record your observations.

NOTE
Ethanol should not be close to the flame since it is flammable.

Answer the following questions

1. Why is it possible to separate ethanol from water?
2. Explain the role played by the fractionating column.
3. State the role of the glass beads in the fractionating column.
4. At what temperature does ethanol distil off?
5. What would happen if the inlet and outlet were exchanged in the Liebig's condenser?

Discussion

Water and ethanol are miscible liquids. They are separated because they have different boiling points. Pure ethanol boils at 78.2°C while pure water boils at 100°C at sea level. When the mixture is heated, ethanol boils off first at about 78°C, and is collected as the first fraction of the mixture.

The temperature remains fairly constant until the ethanol distils off. At this stage the temperature starts rising and the distillate collected thereafter is mainly water as a second fraction. This process of separation is called **fractional distillation.** The purpose of the fractionating column is to allow water vapour to condense into liquid and flow back into the flask before the boiling point of water is reached. The glass beads increase the surface area for condensation. Fractional distillation is used to separate miscible liquids that have different but close boiling points.

A Liebig condenser uses the counter flow principle to cool the vapour efficiently. If the inlet and outlet were exchanged in the Liebig condenser, condensation would still occur but less efficiently.

Applications

1. Distillation of:

(a) crude oil to obtain fractions such as diesel, petroleum, cooking gas. (Kenya Petroleum Refinery in Changamwe, Mombasa),

(b) recycling of used oil (at Athi River and Kikuyu town).

2. Liquid air in the manufacture of nitrogen and oxygen. (British Oxygen Company, BOC)

Oil storage tanks at an oil refinery.

Experiment 2.11: How can oil be extracted from nuts and seeds?

Crush about 20 groundnut seeds in a mortar using a pestle. Continue crushing the nuts while adding propanone a little at a time. Decant the resulting solution into an evaporating basin. Leave the solution in the sun for sometime. Put a drop of the residual liquid on a piece of paper and hold the paper against light. Record your observations. Cashew nuts, coconuts, castor oil seeds, sunflower seeds and cotton seeds may also be used.

Answer the following questions

1. What is the role of propanone in the experiment?
2. Explain why water is not used as a solvent in this experiment.
3. Why is the solution left in the sun?

NOTE
Solvent extraction, also known as liquid-liquid extraction is a process that allows the separation of two or more components due to their unequal solubilities. The solvent, also known as an extractant is chosen to selectively extract a certain component from a mixture.

Discussion

Groundnuts, cashew nuts and coconuts contain oil which is useful. The oil can be extracted using a suitable solvent such as propanone. The nuts are first crushed to increase the surface area in contact with the solvent. Water which is a common solvent cannot be used in this extraction because it will not dissolve oil.

Once the oil has dissolved in the propanone the solution is left in the sun for the solvent to evaporate. The oil having a higher boiling point than the solvent is left in the evaporating dish. This method of extraction is known as **solvent extraction.** Oils leave a translucent patch on paper. This can be used as a simple test for their presence. Oil obtained this way can be made more pure by washing the product in water and separated from the water using a separating funnel.

Applications

1. Extraction of:
   (i) oil from nuts and seeds.
   (ii) natural dyes from plants.

(iii) some herbal medicines from plants

(iv) caffeine from tea and coffee

2. In dry-cleaning to remove dirt.

Separation of Coloured substances

Naturally occurring substances may contain several pigments (colours). For example green grass has several pigments.

Experiment 2.12: How can a mixture of pigments be separated?

Crush some green leaves or grass in a mortar using a pestle. Add propanone as you continue crushing the leaves.

Decant the extract into a beaker. Place a filter paper on top of an empty beaker. Using a dropper place one drop of the extract at the centre of the filter paper and allow it to spread as far as possible. Add a second and third drop at the same spot. Each time allow the extract to spread as far as possible.

Once the spot of the extract has stopped spreading, add the solvent drop-wise each time allowing the solvent to spread. Continue adding the solvent until it spreads out close to the edge of the filter paper. Draw the diagram of the filter paper showing the results obtained.

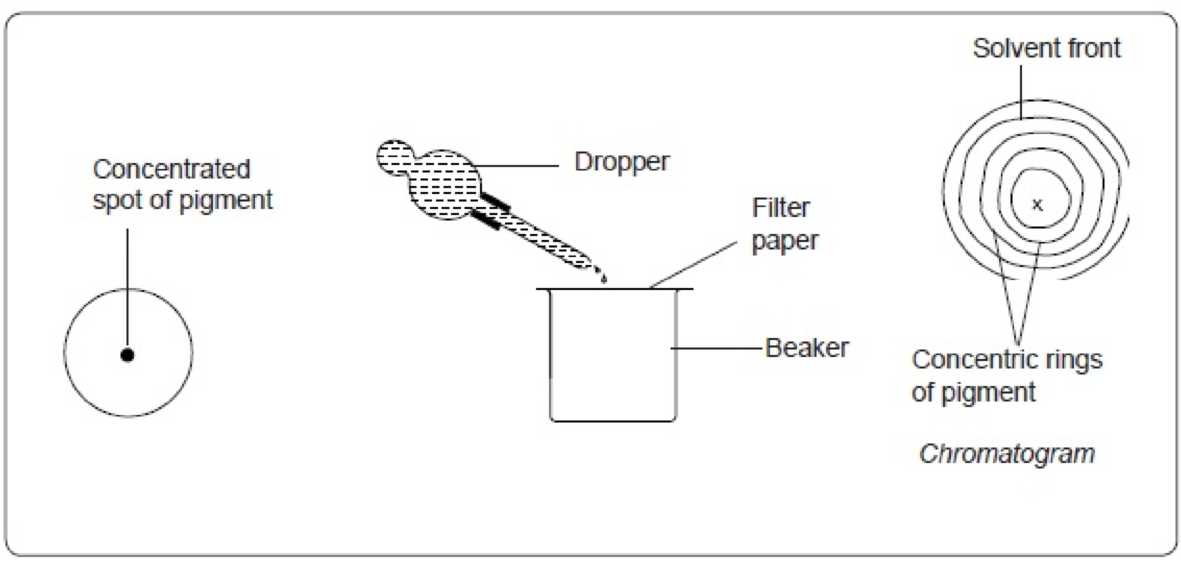


Fig 2.11: How to conduct a simple paper chromatography.

Answer the following questions

1. What is observed at the end of the experiment?
2. How many pigments are there in green leaves extract?
3. Explain how the pigments separate?

NOTE

Adsorption involves the binding of molecules or particles of a gas or liquid to the surface of a solid. It is a temporary and reversible process. It differs from absorption, the filling of pores in a solid.

Discussion

The colouring matter in green leaves is composed of different substances. Propanone is used to dissolve the colouring matter. Each coloured substance has a different solubility in propanone and a different extent of **adsorption** on the filter paper **(adsorbent material).** As the propanone spreads the pigments which are more soluble and less adsorbed are carried furthest while the less soluble and more adsorbed are left behind; as a result separation takes place. The furthest point where the solvent reaches on the adsorbent material is called the **solvent front.**

This method of separation is called **chromatography.** *Chroma* means colour and *chromatology* means study of colours. The coloured matter in leaves separates into two distinct pigments; a green colour due to a substance called **chlorophyll** and a yellow colour due to **xanthophylls.** The dry filter paper showing the separated pigments is called **chromatogram.** Xanthophyll is more soluble and less adsorbed.

Chromatography can be used to determine the presence of a substance in a mixture by comparing it with a pure substance. The suspect mixture is placed on an adsorbent medium alongside the pure substance on the same baseline as shown in figure 2.12 (a).

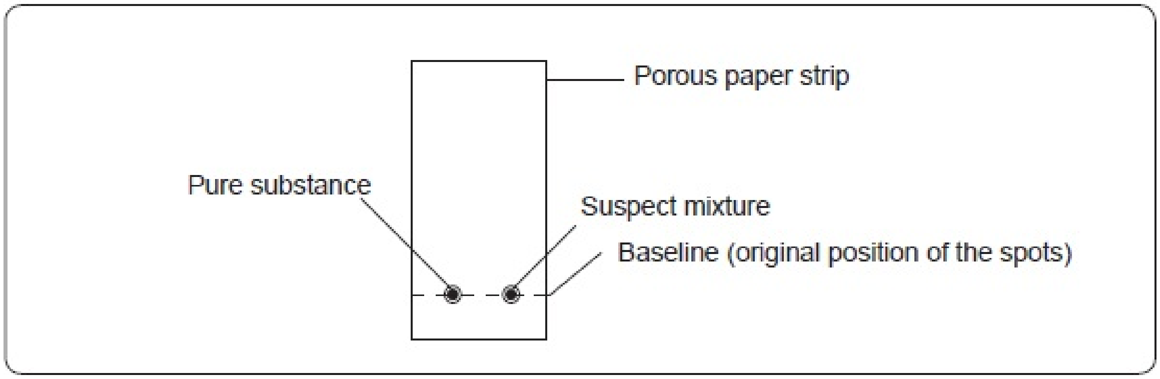


Fig 2.12 : Strip ready for ascending paper chromatography.

The paper strip is then placed in a beaker containing a solvent as shown in figure 2.13 (a).

The solvent is allowed to ascend to the top **(solvent front)** and the paper is then removed and allowed to dry. The position of the spots from the mixture and the pure substance are noted and compared.

If any of the spots in the mixture moves the same distance as the spot in the pure substance, then the mixture contains the pure substance as one of the components. This procedure is known as ascending paper chromatography. Figure 2.13 (b) shows a sample chromatogram.

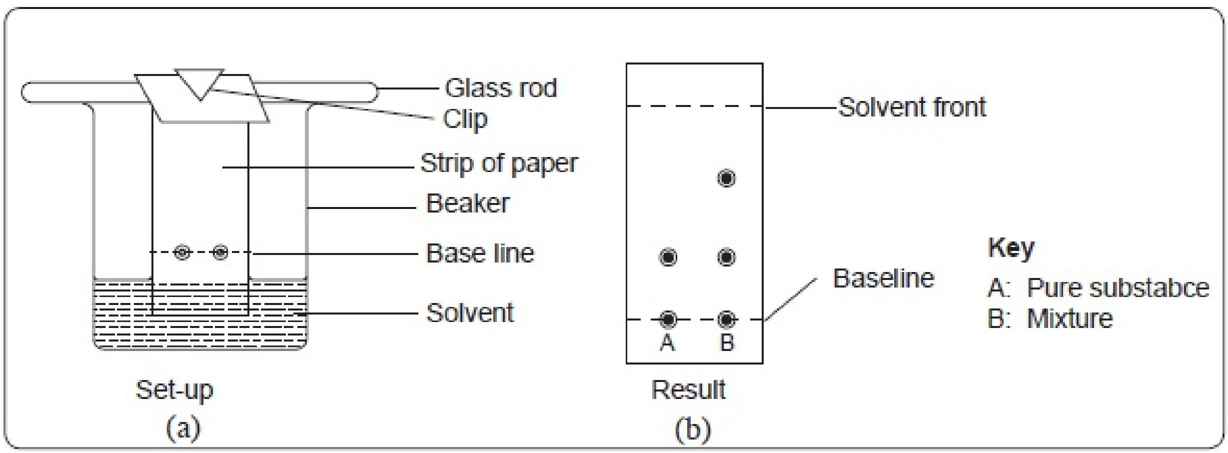


Fig 2.13 Ascending paper chromatography.

Applications

1. In sports, chromatography, is used to identify banned substances, for example steroids in urine or in blood samples.
2. In the pharmaceutical industry, to test purity of drugs.
3. In food industry, to identify contaminants in food and drinks.
4. In the cosmetics industry, to identify harmful substances.

Effect of Heat on Substances

States of Matter

Matter is anything that has mass and occupies space. Matter exists in three physical states namely solid, liquid and gas. A substance can exist in any of the three states depending on the prevailing temperature.

Experiment 2.13: What happens when naphthalene is heated?

Place a spatulaful of naphthalene in a boiling tube. Note the temperature of the naphthalene. Place the boiling tube in a beaker of water as illustrated in figure 2.13. Heat the beaker. Record the temperature of the naphthalene every half a minute throughout the experiment until all the solid has melted. Continue heating for about two more minutes. Record the temperature readings as in table 2.2 and plot a graph of temperature against time.

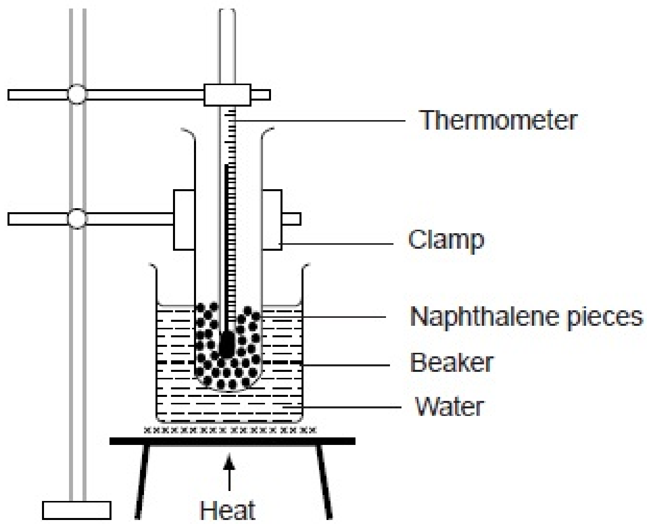


Fig 2.14: Heating pure sample of naphthalene.

Table 2.2: Effect of heat on naphthalene.

Time in minutes	0.0	0.5	1.0	1.5	2.0	2.5	3.0	3.5	4.0	4.5
Temperature in °C										

Answer the following questions

1. What happens to naphthalene when it is heated? Explain.
2. Is there a temperature change as naphthalene is melting?
3. How does heating cause melting?
4. Explain how the change observed in the experiment can be reversed.

Discussion

When naphthalene is heated, its temperature rises steadily until it starts to melt. At this point, the temperature remains constant until all the naphthalene has melted then the temperature starts rising again. Heating in a water bath does not allow the boiling point of naphthalene (218°C) to be attained and therefore protects the thermometer from breaking.

NOTE

If a thermometer of 0 - 360°C range is used, naphthalene can be heated directly.

A sketch of the graph of temperature against time obtained when naphthalene is heated to melting given in figure 2.15 (a). Such a graph is called a **heating curve.**

If liquid naphthalene is allowed to cool, the reverse can be obtained. It is then referred to as **cooling curve.**

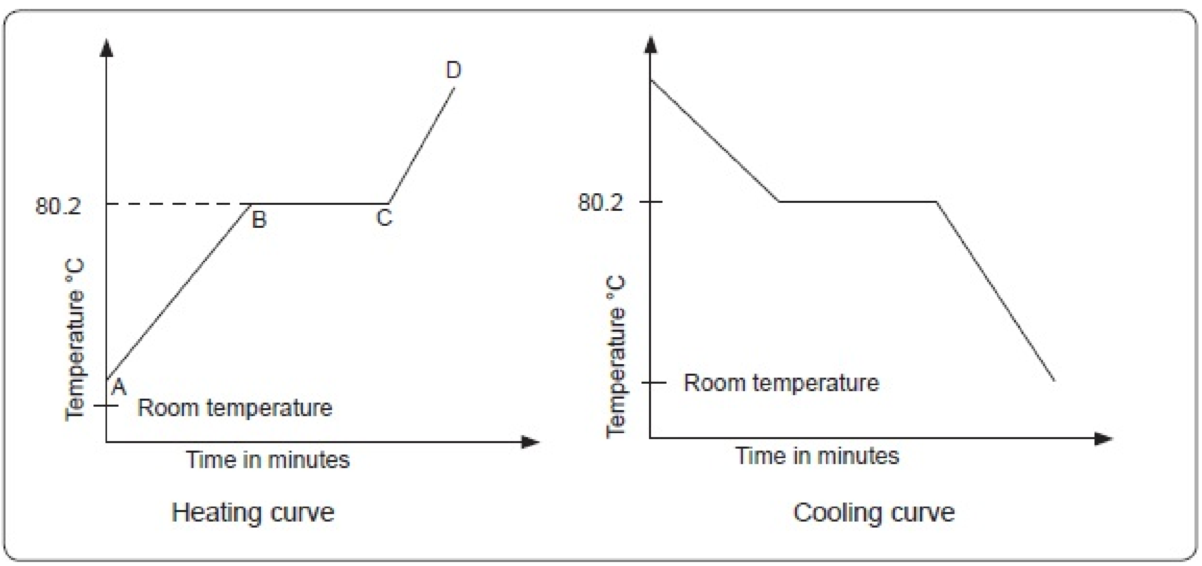


Fig 2.15 (a) Graphical representation of heating and cooling curves of naphthalene

Region AB: The temperature increases steadily as the naphthalene absorbs heat energy. The heat absorbed increases the kinetic energy of the particles and they vibrate more vigorously.

Region BC: The temperature remains constant until all the naphthalene melts. Here, the heat supplied is used to weaken the forces of attraction holding the particles of naphthalene together. The particles therefore move far apart. As a result, the naphthalene changes its state from solid to liquid.

Region CD: Temperature rises steadily as the liquid naphthalene absorbs heat energy. The heat supplied increases further the kinetic energy of the particles causing them to move fast. Figure 2.15 (b) represents a heating curve of a solid which is heated from room temperature to boiling point.

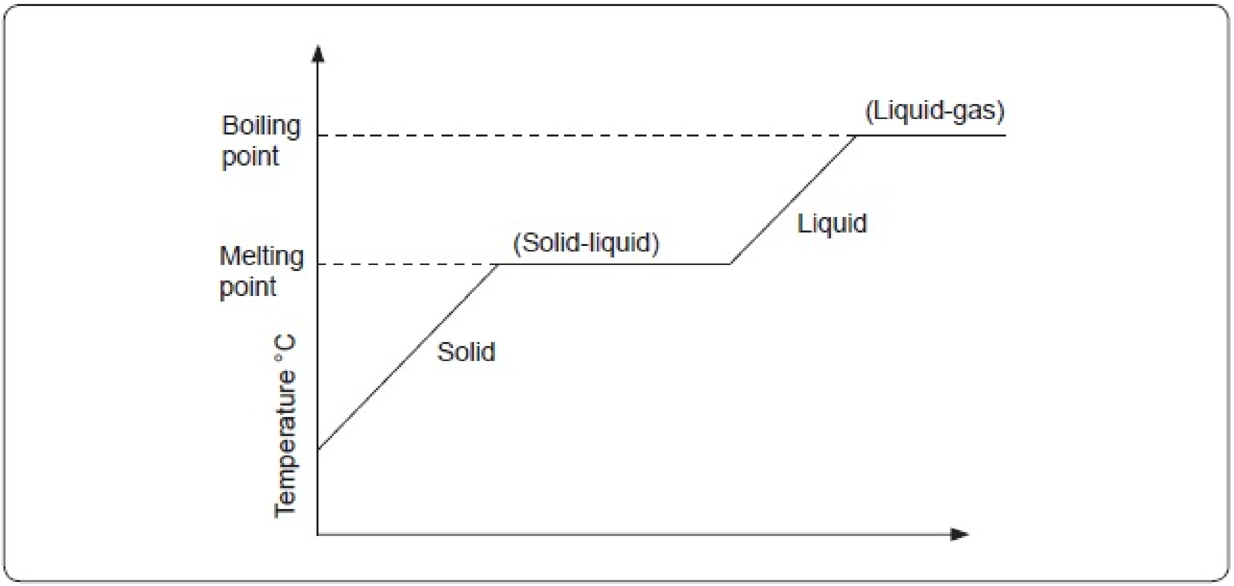


Fig 2.15 (b): Heating curve of a pure substance.

Experiments 2.14: What happens when water is heated?

Put 10 cm^3 of distilled water in a boiling tube. Measure and record the temperature of the water. Arrange the apparatus as shown in figure 2.16. Heat the water gently and record its temperature every 30 seconds until it boils. Continue heating the water after it has boiled for two more minutes. Record your observations as in table 2.3. Draw a graph of temperature against time.

NOTE

The glass tubing ensures that there is no pressure build up in the boiling tube. Increased pressure causes an increase in boiling point of the water. Water may not boil at 100˚C because of variation in pressure from place to place due to changes in altitude.

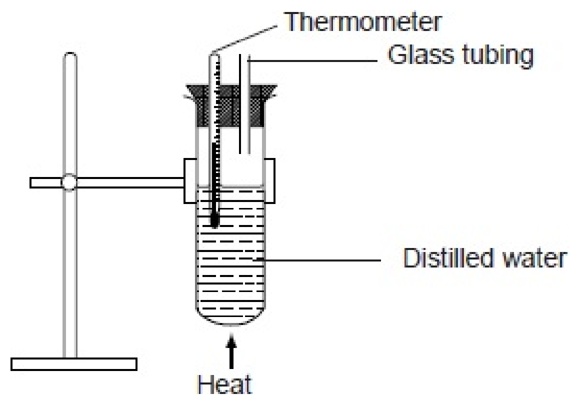


Fig 2.16 Heating water to boiling point.

Table 2.3: Effect of heat on water

Time in seconds	0	30	60	90	120	150	180	210	240
Temperature in ^0C									

Answer the following questions

1. What is observed when water is heated? Explain.
2. Is there a temperature change as water is boiling? Explain.
3. How does heating cause boiling of liquids?

NOTE
Boiling occurs through out a liquid at a fixed temperature. Evaporation. occurs only at the surface of a liquid and at any temperature.

Discussion

When water is heated its temperature steadily increases as molecules absorb heat energy, which increases their kinetic energy. The temperature of the water continues to rise untill the water starts to boil.

The temperature of the water remains constant as it boils. This is because the heat energy supplied is used to break the forces of attraction holding the particles together. As a result some particles break free and the water changes from liquid to gaseous state. The graph of temperature against time obtained when water is heated to boiling is given in figure 2.17.

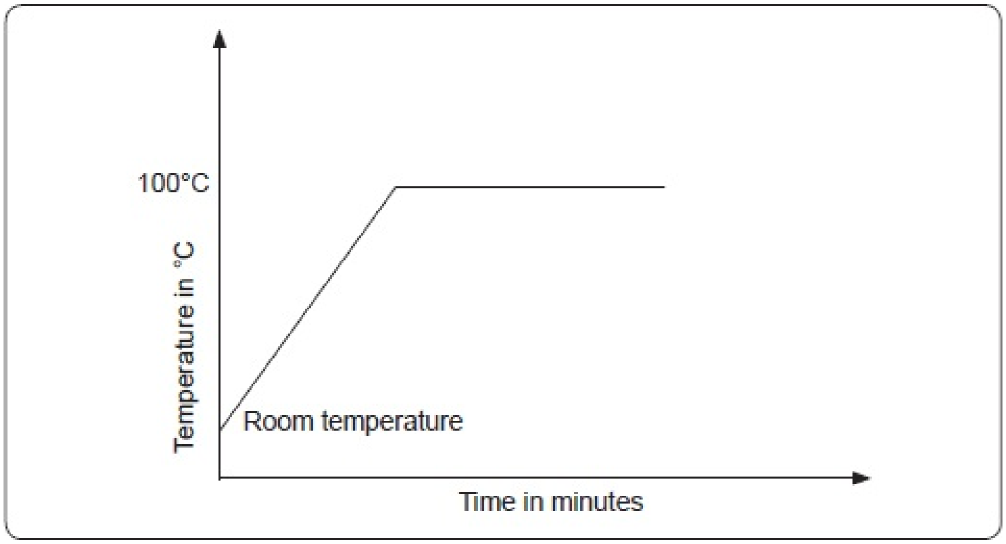
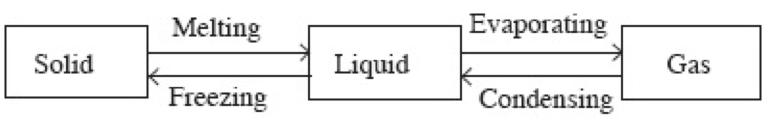


Fig 2.17 Graph showing temperature changes when water is heated to boiling point.

Change of state from solid to liquid and liquid to gas can be reversed by cooling. On cooling, the gas condenses into liquid and finally the liquid freezes into solid. These changes are shown in the flow diagram below.



The arrow signs $\rightleftharpoons$ means a reversible change.

The use of the thermometer enables us to observe that changes in the states of matter are not accompanied by temperature changes. Some forces of attraction hold together the particles which make up matter. Heating provides the energy required to overcome these forces.

Some substances such as iodine do not undergo the above changes because they sublime.

$$\underset{\text{(dark cystals)}}{\text{Iodine}} \xrightleftharpoons[\text{cool}]{\text{heat}} \underset{\text{(purple vapour)}}{\text{Iodine}}$$

The explanation for the behaviour of matter in terms of the movement of particles is provided by the kinetic theory of matter.

The Kinetic Theory of Matter

According to the theory, matter is made up of particles which are in a continuous state of motion. The kinetic theory forms the basis of the theoretical model of matter. Figure 2.18 shows the model in which circles represent small particles of matter.

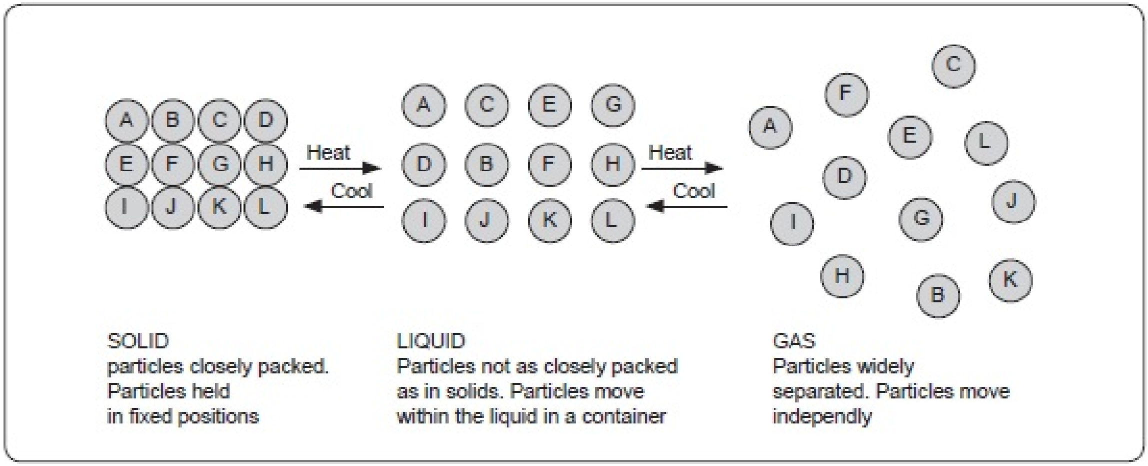


Fig 2.18: A theoretical model of matter.

In the solid state, the particles are closely packed together and can only vibrate within fixed positions. They do not move from one point to another because there are forces that hold them in these positions. When a solid is heated, the kinetic energy of the particles increases and they begin to vibrate more vigorously. At a certain temperature which is fixed for particular substance, the forces holding the particles are weakened enough to allow the particles to change position as a result the solid changes into a liquid. This temperature is known as the **melting point.**

NOTE
The rate of evaporation of a liquid is highest at its boiling point.

In the liquid state, particles are not as close together as they are in the solid state. They can move from one position to another within the liquid. This explains why a liquid has no definite shape and will take the shape of the container. However, a liquid has a definite volume. The particles exert some attraction on one another and these forces of attraction make them to stay together.

When a liquid is heated, the particles move more rapidly as the forces of attraction are further weakened. The weakening continues until the particles gain enough energy to overcome the forces between them. At this point, the liquid boils as particles break free and enter the gaseous phase/state. *The constant temperature at which a pure liquid boils is called the* ***boiling point*** and is fixed for a particular substance. The temperature at which a liquid boils depends on the external atmospheric pressure.

In the gaseous state the particles are far apart and free to move randomly in all directions. This is why a gas does not have definite shape or volume but occupies the whole space within a container.

When a gas is cooled the particles lose kinetic energy and hence slow down. As they slow down they easily attract their neighbouring particles and move close to form a liquid. This process is called **condensation.** Condensation occurs at the same temperature as evaporation.

On further cooling of the liquid, the kinetic energy of the particles continues to decrease and the particles take up fixed positions as the liquid solidifies. This process is called **freezing.** Freezing takes place at the same temperature as melting.

Effect of Impurities on Melting and Boiling Points of Substances

Experiment 2.15: What is the effect of impurities on melting point of naphthalene?

Place a spatulaful of naphthalene on a clean piece of paper. Add a small amount of camphor and mix them thoroughly. Put the mixture into a boiling tube.

Place the boiling tube in a beaker of water as shown in figure 2.19. Heat the water gently and stir it continuously. Record the temperature after every half-minute throughout the experiment. Note the temperature at which the mixture starts to melt and when all of it has melted. Record the temperature readings as shown in table 2.4 and plot the graph of the temperature against time.

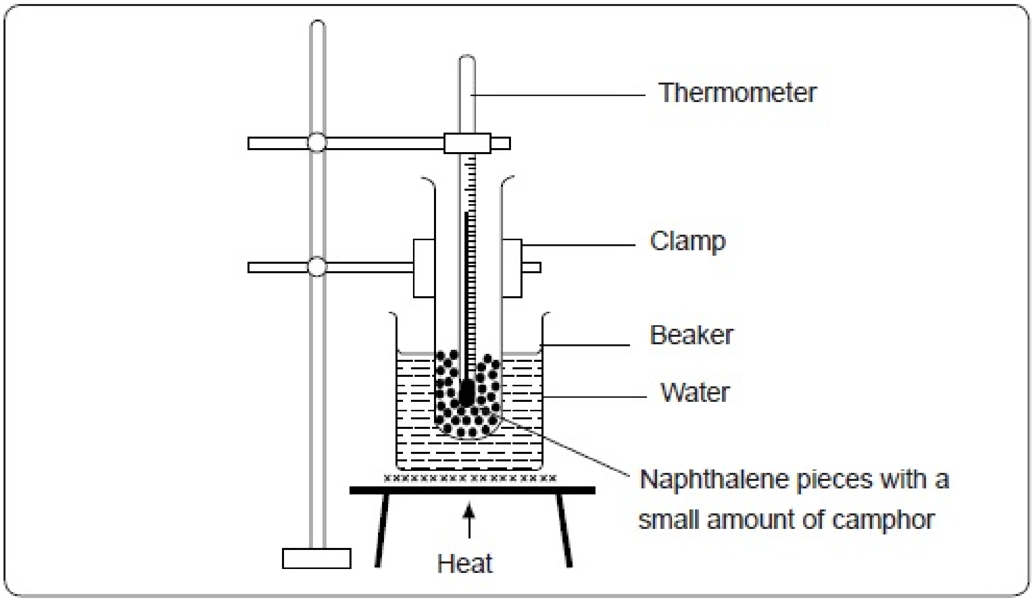


Fig 2.9: Heating an impure sample of Naphthalene.

Table 2.4: Effect of impurities on melting and boiling points

Time in minutes	0.0	0.5	1.0	1.5	2.0	2.5	3.0	3.5	4.0
Temperature in °C									

Answer the following questions

1. At what temperature;
   (a) does the impure naphthalene start melting?
   (b) is the melting of the impure naphthalene complete?
2. Compare the temperature at which pure naphthalene and impure naphthalene melts.
3. What effect do impurities have on the melting point?

Discussion

The melting point of pure naphthalene is 80.2°C. If camphor or another substance is added as an impurity, the naphthalene melts over a temperature range that is lower than the melting point of pure naphthalene. Impure naphthalene does not have a sharp melting point. Impurities therefore lower the melting point of substances.

The graph of temperature against time obtained for impure naphthalene when heated to melting is given in figure 2.20.

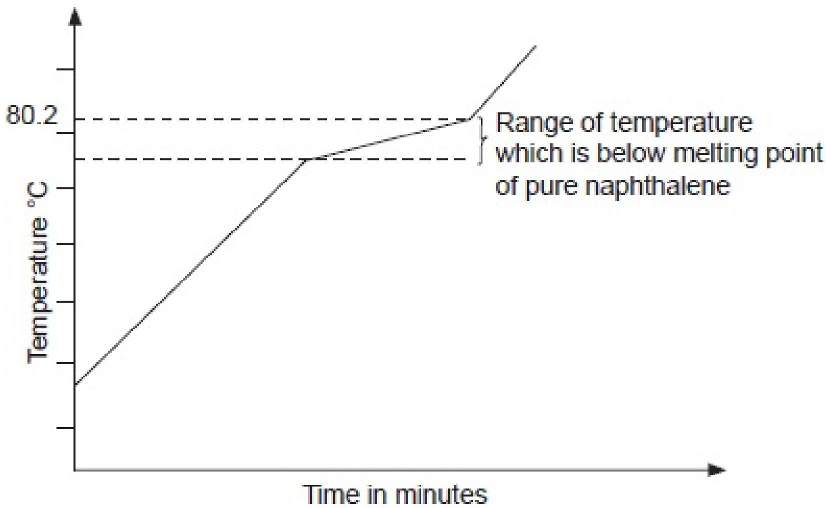


Fig 2.20: Heating curve of impure naphthalene.

Applications

The lowering of melting point by impurities is applied in:

1. Clearing of ice from roads in temperate countries by spreading common salt on the ice.
2. Extraction of metals from their molten compounds. For example, calcium chloride is added to rock salt during the extraction of sodium from sodium chloride.

Experiment 2.16: What is the effect of impurities on the boiling point of water?

Put 5cm^3 of distilled water into a boiling tube as shown in figure 2.21. Heat the water gently and record the temperature at which it boils. Allow it to cool. Add a spatulaful of sodium chloride (impurity) to the water. Measure and record the temperature of the solution.

Repeat the experiment. Record the temperature at which the solution begins to boil. Continue heating for a further two minutes recording after every half a minute.

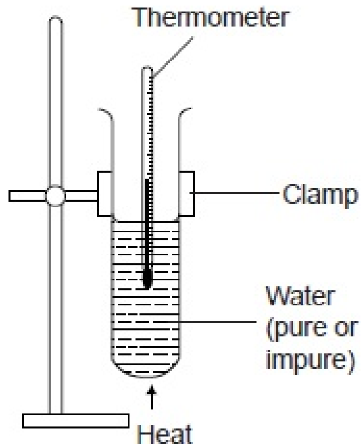


Fig 2.21: Effect of impurities on the boiling point of water.

Answer the following questions

1. At what temperature does:
   (a) pure water boil?
   (b) impure water boil?
2. What effect do impurities have on the boiling point?

Discussion

The boiling point of pure water is 100°C at sea level. The impure water starts to boil at a temperature above 100°C. The temperature continues to rise as the impure water boils. Thus the impure water boils over a range of temperature. Impurities raise the boiling point of a liquid. The heating curves for pure and impure water are shown in figure 2.22 (a) and (b).

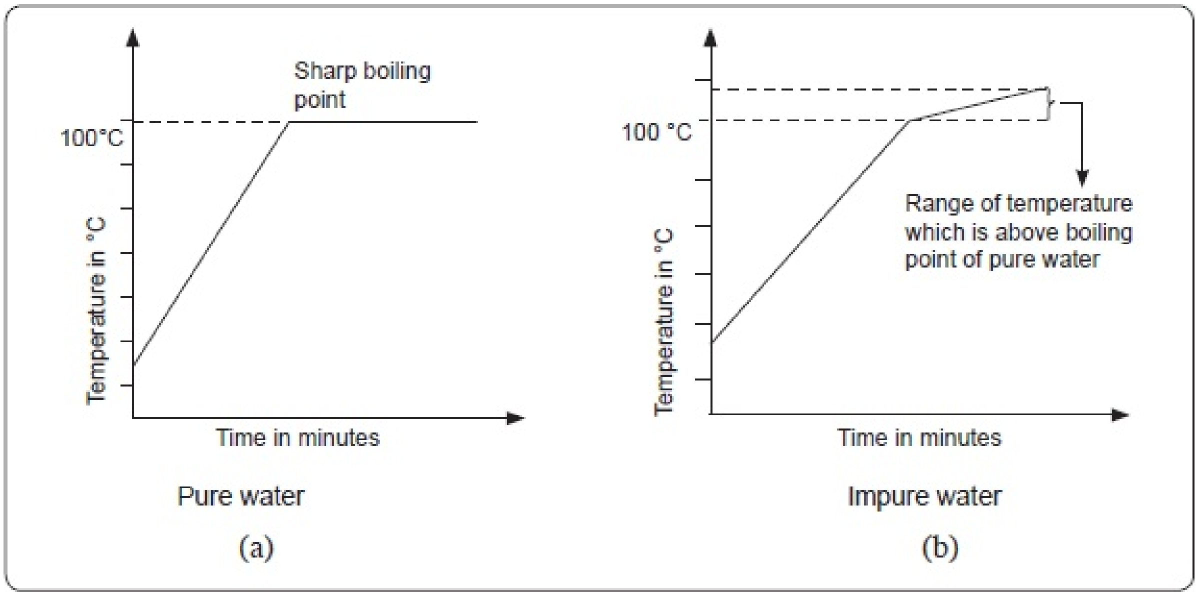


Fig 2.22: Heating curves of pure and impure water.

Criteria of Purity

Pure substances melt and boil at constant temperatures that are specific for a particular substance. Melting and boiling points are therefore used for determining purity of substances.

Temporary and permanent changes

Substances undergo various changes when subjected to different conditions of temperature. This section classifies and defines changes that take place in substances.

Experiment 2.17: What happens when different solids are heated?

Put half a spatulaful of zinc oxide in a test-tube. Heat the test-tube until no further change occurs. Allow it to cool. Record your observations as in table 2.5. Repeat the experiment using wax, and iodine.

Table 2.5: Effect of heat on various substances

Solid	**Observation**	
	When heated	*On cooling*
Zinc oxide		

Wax		
Iodine		

Discussion

When zinc oxide is heated, its colour changes from white to yellow. On cooling, the yellow solid turns white. This change can be represented as:

$$\underset{\text{(white)}}{\text{Zinc oxide}} \underset{Cool}{\overset{Heat}{\rightleftharpoons}} \underset{\text{(yellow)}}{\text{Zinc oxide}}$$

Wax melts on heating. When cooled, liquid wax changes back to solid.

$$\text{Solid wax} \underset{Cool}{\overset{Heat}{\rightleftharpoons}} \text{Liquid wax}$$

When iodine is heated, the shiny black solid turns to a purple vapour. When cooled, the purple vapour changes back to the shiny black solid.

$$\underset{\text{(shinny black)}}{\text{Solid iodine}} \underset{Cool}{\overset{Heat}{\rightleftharpoons}} \underset{\text{(purple)}}{\text{Iodine vapour}}$$

Heating zinc oxide, wax and iodine does not result in the formation of a new substance. Cooling, reverses the changes these substances undergo. A change which can easily be reversed and in which no new substance is formed is called a **temporary physical change.** The following are the characteristics of temporary physical changes.

(a) They are easily reversible.

(b) No new substance is formed.

(c) The mass of the substance does not change.

(d) They are not accompanied by net heat change.

Experiment 2.18: What happens when copper (II) sulphate crystals are heated?

Put a spatulaful of copper (II) sulphate crystals in a dry boiling tube and set-up the apparatus as shown in figure 2.23. Heat the copper (II) sulphate gently until there is no further change. Disconect the delivery tube, continue heating for a while, then stopper the boiling tube. Allow it to cool and record your observations. Divide the solid into two portions. Put the portions into two separate test-tubes. Put a thermometer into the first test tube. Measure and record the temperature of the solid. Using a dropper, add about three to four drops of tap water into the test tube. Take the temperature of the resultant mixture. Place a thermometer into the second portion, measure and record the temperature of the solid. Add two to three drops of the liquid collected during heating. Take the temperature of the resultant mixture. Repeat the experiment using cobalt (II) chloride crystals.

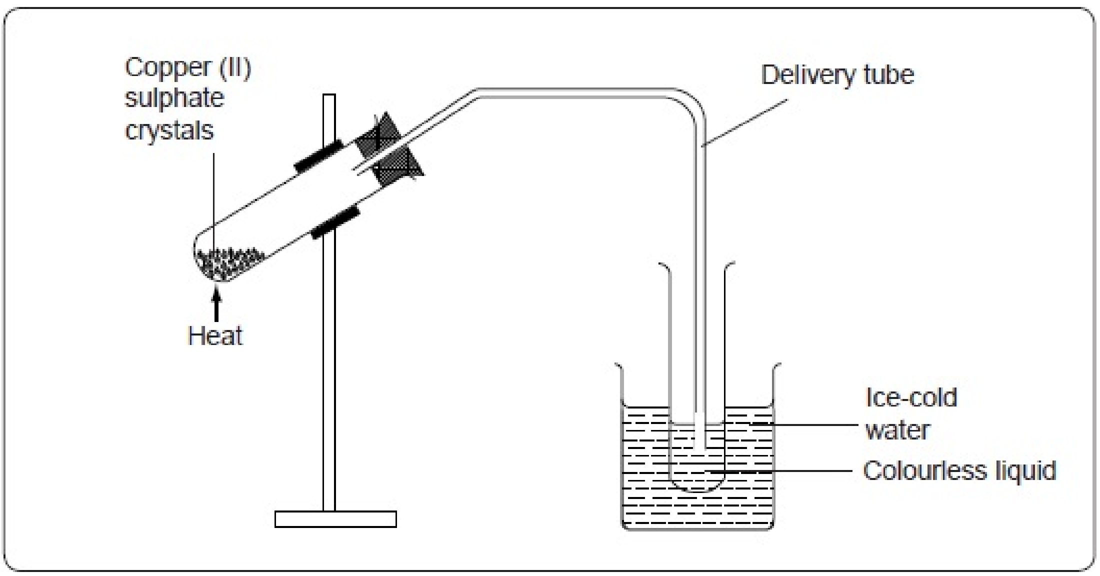


Fig 2.23: Heating of copper sulphate crystals.

Answer the following questions

1. What is observed when:
   (a) copper (II) sulphate crystals are heated?
   (b) cobalt (II) chloride crystals are heated?
2. What is observed:
   (a) when a few drops of tap water are added to the solid products in 1(a) and (b)?
   (b) the liquid collected in 1(a) and (b) is added to the residue in each case respectively?
3. Identify the liquid collected in the above experiment.

4. Classify the changes.

NOTE
A temporary change may be physical or chemical. In a temporary physical change, no new substance is formed. In a temporary chemical change new substances are formed. However the reactions are reversible if the conditions are reversed.

Discussion

Crystals of copper (II) sulphate contain water of crystallisation. It is said to be **hydrated.** When heated it decomposes to produce white copper (II) sulphate powder and water. The white copper (II) sulphate powder does not contain water of crystallisation and is said to be **anhydrous.**

$$\underset{\text{(blue)}}{\text{Hydrated copper (II) sulphate}} \xrightarrow{\text{Heat}} \underset{\text{(white)}}{\text{Anhydrous copper (II) sulphate}} + \text{Water}$$

The white anhydrous copper (II) sulphate does not regain the original blue colour on cooling.

$$\text{Anhydrous copper (II) sulphate} \xrightarrow{\text{Cool}} \text{Anhydrous copper (II) sulphate}$$

Similarly pink cobalt (II) chloride decomposes to form blue anhydrous cobalt (II) chloride and water vapour.

$$\underset{\text{(pink)}}{\text{Hydrated cobalt (II) chloride}} \xrightarrow{\text{Heat}} \underset{\text{(blue)}}{\text{Anhydrous cobalt (II) chloride}} + \text{water}$$

The blue cobalt (II) chloride does not regain the pink colour on cooling.

$$\underset{\text{(blue)}}{\text{Anhydrous cobalt (II) chloride}} \xrightarrow{\text{Cool}} \underset{\text{(blue)}}{\text{Anyhdrous cobalt (II) chloride}}$$

The decomposition of a substance when it is heated is referred to as **thermal dissociation.**

Heat is evolved when a little water is added to the anhydrous copper sulphate or the blue anhydrous cobalt (II) chloride. The anhydrous substances become hydrated and regain their original colour.

White copper (II) sulphate + water ⟶ Blue copper (II) sulphate
(anhydrous) (hydrated)

Blue cobalt (II) chloride + water ⟶ Pink cobalt (II) chloride
(anhydrous) (hydrated)

Heating hydrated copper (II) sulphate or hydrated cobalt (II) chloride results into a **temporary chemical change.** The characteristics of temporary chemical change are:

(a) A new substance is formed.

(b) Heat energy is evolved or absorbed.

(c) There is change in mass.

(d) The change can be reversed.

CAUTION

Small quantities of copper (II) Nitrate and Potassium Manganate (VII) should be used as the fumes produced on heating could be harmful when inhaled.

Experiment 2.19: What happens when copper (II) nitrate is heated?

Weigh a dry test tube. Put a half spatulaful of copper (II) nitrate crystals in the test-tube and weigh. Heat the copper (II) nitrate crystals gently and then strongly. Record your observations. Test the gas evolved with a glowing splint. Allow the product to cool and then re-weigh. Repeat the experiment using potassium manganate (VII).

Answer the following questions

1. What is observed when,
   (a) copper (II) nitrate is heated?
   (b) potassium manganate (VII) is heated?
2. Explain why there are changes in mass after heating in both cases.

Discussion

When copper (II) nitrate is heated it decomposes to form a black solid and a mixture of gases. The black solid is copper (II) oxide. The mixture of gases consists of a red brown gas and another gas which relights a glowing splint. The red brown gas is nitrogen (IV) oxide while the gas which relights the glowing splint is oxygen.

$$\underset{\text{(blue)}}{\text{Copper (II) nitrate}} \xrightarrow{\text{Heat}} \underset{\text{(black)}}{\text{Copper (II) oxide}} + \underset{\text{(red-brown)}}{\text{Nitrogen (IV) oxide}} + \underset{\text{(colourless)}}{\text{Oxygen}}$$

The mass of copper (II) oxide is found to be less than that of copper (II) nitrate because the gaseous products escaped into the atmosphere.

Potassium manganate (VII) decomposes to form a black-green solid and a colourless gas, which relights a glowing splint. The black-green solid is a mixture of black manganese (IV) oxide and green potassium manganate (VI). The black-green solid weighs less than the original potassium manganate (VII).

$$\underset{\text{(Purple)}}{\text{Potassium manganate (VII)}} \xrightarrow{\text{Heat}} \underset{\text{(green)}}{\text{Potassium manganate (VI)}} + \underset{\text{(black)}}{\text{Manganese (IV) oxide}} + \text{Oxygen}$$

The decomposition of copper (II) nitrate and potassium manganate (VII) are examples of **permanent chemical changes.** Permanent chemical changes have the following characteristics.

(a) New substances are formed.

(b) The change is irreversible.

(c) The change is accompanied by change in mass.

(d) Heat energy is released or absorbed.

Constituents of Matter

In the preceeding discussions, matter has been classified as solid, liquid or gas. Pure substances can be classified as **elements** or **compounds.**

Elements

Elements are pure substances which can not be split into simpler substances by chemical means. Examples of elements include oxygen, hydrogen, copper, sulphur, carbon and iron. There are over a hundred known elements.

Elements are made up of **atoms.** *The atom is defined as the smallest particle of an element, which can take part in a chemical change.* Atoms of the same element are similar.

The atoms of some elements can not exist independently but join together to form small groups of atoms. These discreet particles are called **molecules.** A molecule is defined as *the smallest particle of an element or compound, which can exist separately*.

Compounds

Experiment 2.20: What happens when a mixture of iron and sulphur is heated?

Place half a spatulaful of fresh iron filings on a piece of paper. Take another piece of paper and place on it a spatulaful of powdered sulphur. Examine the two substances and note their colours. Mix them on one piece of paper and note the colour of the mixture.

Hold a magnet above the mixture and observe what happens.

Place a small amount of this mixture in a crucible and heat strongly. Allow it to cool. Observe its colour. Hold a magnet over it. Record your observations.

Answer the following questions

1. What is observed when the mixture is being heated?
2. What is the:
   (a) colour of the new substance formed in the experiment?
   (b) effect of the magnet on the new substance?

Discussion

Iron filings are grey in colour while sulphur is a yellow solid. When the two are mixed, they form a mixture in which the elements (iron and sulphur) retain their individual colours. When a magnet is held above the mixture, the iron filings are attracted.

On heating the mixture strongly, a red glow spreads through the mixture. This is due to the heat produced as a result of a chemical reaction-taking place. A black solid is formed. The solid is not attracted by a magnet.

The black solid contains sulphur and iron chemically combined. When two or more elements combine chemically they form a **compound.** *A compound is a pure substance made up of two or more elements chemically combined*. The compound formed when iron and sulphur chemically combine is called iron (II) sulphide.

It is possible to tell which elements are present in a compound from its name. Names ending in **–ide** means the compound is composed of two elements only. e.g:

Sodium Chloride is made up of Sodium and Chlorine.

Iron (II) Sulphide is made up of Iron and Chlorine.

Calcium Nitride is made up of Calcium and nitrogen.

Calcium carbide is made up of Calcium and Carbon.

Sodium hydride is made up of Sodium and hydrogen.

An exception to this are the hydroxides.

Names ending in **–ate** means the compound is composed of three elements one of which is oxygen. eg:

Sodium Sulphate is made up of sodium, sulphur and oxygen.

Sodium Carbonate is made up of sodium, carbon and oxygen.

Potassium chlorate is made up of potassium, chlorine and oxygen.

Calcium nitrate is made up of calcium, nitrogen and oxygen.

Exceptions to this rule are the hydrogen carbonates and hydrogensulphates.

Names ending in **–ite** means the compound is made up of three elements one of which is oxygen. However, the amount of oxygen is less than in those compounds whose names end in **–ate** e.g

Sodium Sulphite is made up of sodium, sulphur and oxygen.

Calcium Sulphite is made up of calcium, sulphur and oxygen.

Chemical Symbols

In Chemistry, elements are represented by letters. The letters are referred to as chemical symbols. A chemical symbol of an element is usually the first letter or the first and another letter of the element's English or Latin name.

The first letter of a chemical symbol must always be a capital letter while the second letter is always a small letter. Table 2.6 shows the chemical symbols of some common elements.

Table 2.6 Symbols derived from first letter of name of element

Element	*Symbol*
Carbon	C
Nitrogen	N

Oxygen	O
Hydrogen	H

Symbols of several elements may begin with the same letter. It is therefore necessary to represent some of these elements with two letters as shown in the following table. The second letter in a chemical symbol is always a small letter.

Table 2.7 Symbols derived from first and second letter of name of element

Elements	*Symbol*
Calcium	Ca
Cobalt	Co
Chlorine	Cl
Magnesium	Mg
Manganese	Mn

In some cases, the symbol of the element is derived from the element's Latin name as shown in the table below.

Table 2.8 Symbols derived from latin name of element

Element	*Latin name*	*Symbol*
Potassium	Kalium	K
Sodium	Natrium	Na
Silver	Argentum	Ag
Gold	Aurum	Au
Iron	Ferrum	Fe
Lead	Plumbum	Pb

Mercury	Hydragyrum	Hg
Copper	Cuprum	Cu

Chemical Equations

Iron and sulphur combine chemically to form iron (II) sulphide. In this process iron and sulphur are referred to as the **reactants** whereas the iron (II) sulphide is referred to as the **product.**

Chemical combination of elements is known as a reaction.

$$\underset{\text{(reactants)}}{\text{Iron + Sulphur}} \longrightarrow \underset{\text{(product)}}{\text{Iron (II) sulphide}}$$

When a reaction is represented as shown, the representation is called a **word equation.** In a chemical equation reactants are written on the left hand side of the arrow sign $\longrightarrow$ while the products are written on the right hand side.

NOTE
When charcoal burns the mostly carbon material combines with oxygen to form carbon dioxide, the ash left behind constitutes impurities in the charcoal. An example of a useful chemical reaction.

The forward arrow sign $\longrightarrow$ is used where the reactions are permanent and proceed only in one direction. Two arrow signs in opposite directions $\rightleftharpoons$ are used where the reactions are reversible. This means the reaction can proceed in either direction.

NOTE
The plus sign on the reactants side means “react with.” The plus sign on the product side means “and”
The arrow sign between the reactants and the products means to “yield”.

1. Iron reacting with sulphur

$$\text{Iron + Sulphur} \longrightarrow \text{Iron (II) sulphide}$$

2. Sulphur reacting with oxygen

$$\text{Sulphur + Oxygen} \longrightarrow \text{Sulphur (IV) oxide}$$

3. Carbon reacting with oxygen

$$\text{Carbon + Oxygen} \longrightarrow \text{Carbon (IV) oxide}$$

4. Sublimation of Iodine

$$\text{Iodine} \rightleftharpoons \text{Iodine vapour}$$

Chemical reactions can also be represented using chemical symbols. This will be studied in Pupils' Book 2.

Summary

1. Matter exists in three physical states; solid, liquid and gas.
2. Liquids which do not mix with each other are said to be immiscible. Those which mix are said to be miscible.
3. Fractional distillation is a method of separating miscible liquids with different but close boiling points.
4. Oil in nuts and seeds can be obtained by solvent extraction.
5. A mixture of coloured substances is separated by chromatography.
6. Melting is the change from solid to liquid and freezing is the change from liquid to solid. Both processes occur at the same temperature.
7. The kinetic theory of matter states that matter is made up of particles which are in a continuous state of motion.
8. The particles in solids and liquids are held together by cohesive forces.
9. Melting and boiling points can be used to determine the purity of substances.
10. An element is a substance made of one kind of atoms and can not be split into a simpler substance by chemical means.
11. An atom is the smallest particle of an element, which can take part in a chemical change.
12. A compound is a substance which is made up of two or more chemically combined elements.
13. Changes in physical state are temporary changes. They are easily reversed.

Revision Exercise

1. Explain the differences between solid, liquid and gaseous states using the

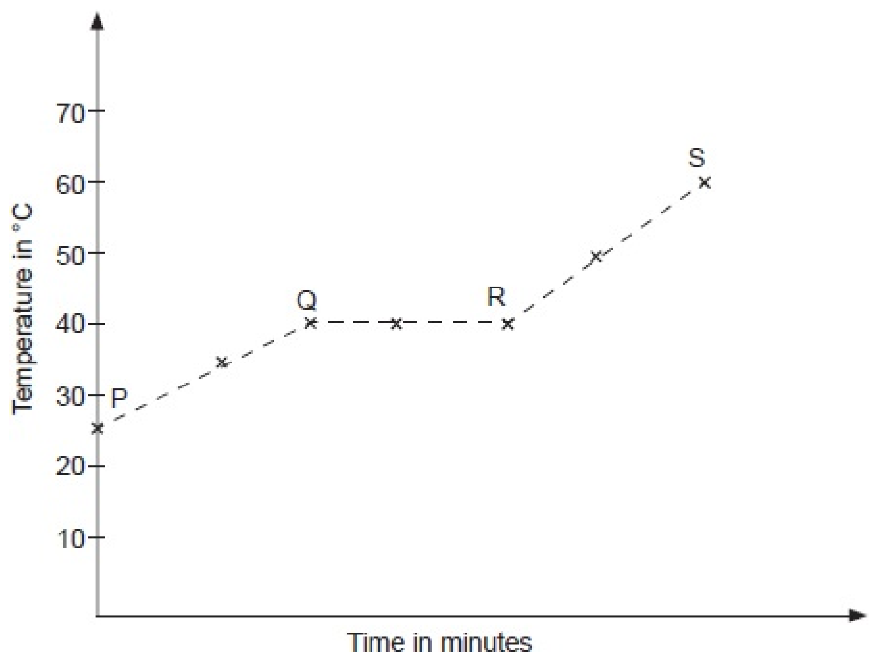
theoretical model of matter in terms of the kinetic theory.

2. The graph below shows the shape of the curve obtained by a student when solid X was heated to boiling.

(a) (i) Determine the melting point of solid X.

(ii) State and explain what portions PQ, QR and RS represent.

(b) If candle wax was used in this experiment, the portions QR would not be horizontal. What does this tell us about candle wax.



3. Name four examples of;

(a) mixtures

(b) compounds

(c) elements

4. Name the elements present in the following compounds.

(a) Sodium bromide

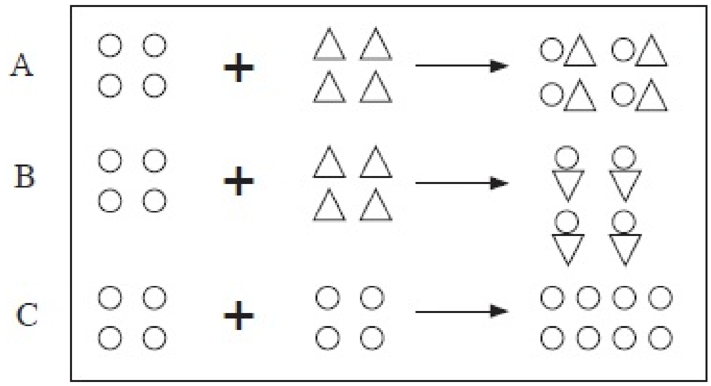
(b) Zinc sulphide

(c) Lead oxide

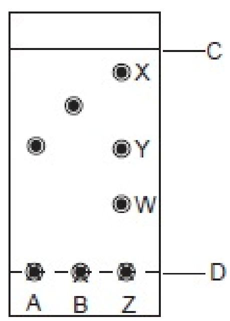
(d) Magnesium nitride

(e) Potassium iodide

5. Which of the following best shows two elements combining to form a compound? Explain.



6. Name a solvent that can be used for the following:
   (a) removing ink from clothes
   (b) cleaning greasy hands
7. How can one separate and collect the solvent from a salt solution?
8. Explain how Elianto oil is obtained from maize seeds.
9. Give two examples of:
   (i) temporary physical change.
   (ii) temporary chemical change.
   (iii) permanent change.
10. List three differences between temporary and permanent changes.
11. Spots of pure pigments A and B, and a mixture Z were placed on a filter paper and allowed to dry. The paper was then dipped in a solvent. The results obtained were as on the paper chromatogram.



   (a) Which is the:
      (i) base line?
      (ii) solvent front?
   (b) Which of the pure pigments was a component of Z? Explain.
   (c) (i) Name a solvent that is used in paper chromatography.

(ii) Why is water not a suitable solvent in paper chromatography?

12. Write a word equation for the reaction between:

(i) carbon and oxygen

(ii) sodium and sulphur

(iii) copper and chlorine

CHAPTER THREE

ACIDS, BASES AND INDICATORS

Some of the foods that man eats are sour while others are bitter. The sour taste is due to the presence of acids whereas the bitter taste is due to the presence of bases. Examples of substances that contain acids are fruits such as oranges, lemons and sour milk. Examples of substances that contain bases include anti-acid tablets and wood ash solution.

Using taste to classify substances as acids or bases is not an accurate method. There are however certain substances, which show different colours when in acids or bases. Such substances are called **indicators** and they are used to classify various substances as either acids or bases.

By the end of this Chapter, you should be able to:

- Define indicator, acid, base and pH.
- Prepare and use plant extracts as acid–base indicators.
- Describe and use the pH scale.
- Use indicators to identify acids and bases.
- State properties of acids and bases.
- Name the uses of acids and bases.
- State and explain the effects of acids.

Simple Acid-Base Indicators

Experiment 3.1: Can flower extracts be used as acid-base indicators?

Crush the flower petals provided using mortar and pestle. Add a small amount of ethanol or propanone. Continue crushing until a deep coloured solution of the extract is obtained. Decant the extract into a beaker or a boiling tube. Place $2cm^3$ of each of the following substances into separate test-tubes and label appropriately: hydrochloric acid, sodium hydroxide, sulphuric acid, potassium hydroxide, lemon juice, orange juice, lime water, wood ash, baking powder and sugar. Add two to three drops of the flower extract into each test-tube and classify the substances in the test-tube as acids or bases. Record your observations as shown in table 3.1. Leave the mixtures to stand until the end of the lesson.

Answer the following questions

1. (a) Does the flower extract give the same colour in all the solutions tested?

   (b) Classify the solutions tested according to the colours observed on addition of the flower extract.

   (c) Do the colours remain the same after sometime? Explain.

The sharp, sour taste in lemons is due to citric acid.

Table 3.1: Use of flower extracts as acid-base indicators

Substance	*Colour of flower extract in the substance*	*State whether acid or base*
Water		
Hydrochloric acid		
Sodium hydroxide		

Discussion

Solutions of hydrochloric acid, sulphuric acid, orange juice and lemon juice give similar colour changes with the same flower extract. Lime water, solutions of sodium hydroxide potassium hydroxide, wood ash and baking powder give a similar but different colour. Water and sugar solution have no effect on the flower extract. They are neutral substances.

The extract from this common ornamental is an indicator, in some varieties flower colour is a good indicator of soil pH.

The composition of flower extracts continuously changes with time causing the colour of the extract to change. The mixture of the flower extract and acid or base also changes colour with time. Flower extracts therefore give inconsistent results when used as acid-base indicators.

For the best results flower extracts should be used when freshly prepared. Other coloured parts of plants may also be used, for example tradescantia, red cabbage and beetroot. In the Laboratory, commercial indicators, which give more consistent results are used as acid base indicators.

Commercial Indicators

Experiment 3.2 (a): What is observed when litmus solution is added to an acid, a base or a neutral solution?

Place 2cm^3 of sodium hydroxide, hydrochloric acid and water into separate test-tubes. Add two drops of litmus solution into each test-tube and record your observations.

Answer the following questions

What is the colour of the litmus indicator in :

(i) an acid.

(ii) a base.

(ii) a neutral solution.

Discussion

Litmus indicator is red in an acid, blue in a base and retains its purple colour in a neutral solution. Litmus indicator is one of the commonly used commercial acid-base indicators. The indicator is also available in paper form as **litmus paper.** Other commercial acid-base indicators include phenolpthalein and methyl orange.

Experiment 3.2 (b): What is observed when some commercial indicators are added to different solutions?

Take three test tubes and put 2cm^3 of hydrochloric acid in each. To the first test-tube, add two drops of litmus solution. To the second and third add two drops of phenolphthalein and methylorange respectively. Record your observation as shown in table 3.2.

Repeat the experiment using water, lemon juice, solutions or suspensions of the following: soap, wood ash, baking powder, anti-acid tablets, toothpaste, sour milk, ammonia, ammonium sulphate, sodium chloride, sodium hydroxide, carbon (IV) oxide (carbon dioxide), sulphur (IV) oxide (sulphur dioxide), sulphuric acid, nitric acid, calcium hydroxide and magnesium oxide.

Answer the following questions

1. Classify the substances as acidic, basic or neutral.
2. State the colour of each indicator in acid, base and neutral solutions.

Discussion

Phenolphthalein indicator is colourless in acidic, pink in basic and colourless in neutral solution. Methyl orange indicator is pink in acidic, yellow in basic and orange in neutral solution.

The colours obtained when the indicators used in the experiment are added to solutions of various substances are shown in table 3.3 (a).

Table 3.2: Effect of various solutions on indicators

INDICATOR / *SOLUTION TESTED*	*Litmus*	*Phenolph-thalein*	*Methyl orange*	*State whether acid or base*
Hydrochloric acid				
Water				
Lemon juice				
Soap				
Wood ash				

Table 3.3 (a): Effect of various solutions on indicators

Indicator	*Colour in*		
	Acid	*Base*	*Neutral*
Litmus	Red	Blue	Purple
Phenolphthalein	Colourless	Pink	Colourless
Methyl orange	Pink	Yellow	Orange

The substances are classified as acidic, basic and neutral as shown in table 3.3 (b).

Table 3.3 (b): Various solutions classified as acidic, basic or neutral

Acidic solutions	*Basic solutions*	*Neutral solutions*
Hydrochloric acid	Soap solution	
Lemon juice	Wood ash solution	
Sour milk	Baking powder	
Ammonium sulphate	Anti-acid tablet solution	Water
Carbon (IV) solution	Toothpaste	Sodium chloride

Sulphur dioxide solution	Ammonia solution	
	Sodium hydroxide solution	
Sulphuric acid	Calcium hydroxide solution	
Nitric acid	Magnesium oxide solution	

Universal Indicators and the pH Scale

The acid-base indicators discussed so far provide no information about the strength of an acid or a base. Some acids are strong while others are weak. Likewise some bases are strong while others are weak.

The universal indicator is a mixture of several indicators. It exhibits a range of colours in acids and in bases depending on the strength of the solution. These shades of colours are related to a continuous acid-base scale called the pH scale. The pH scale has values that range from 0 to 14.

Experiment 3.3: What are the pH values of some solutions?

Place $1cm^3$ of dilute sulphuric acid into a test-tube. Add one drop of the universal indicator and observe the colour of the solution. Place the test-tube and its contents against the pH chart and determine the pH value.

Repeat the experiment using distilled water, dilute solutions of ethanoic acid, lemon juice, nitric acid, hydrochloric acid, sodium chloride, sodium hydrogen carbonate, ammonia solutions, calcium hydroxide, sodium hydroxide, wood ash solution and soap solution respectively. Record the results as shown in table 3.4.

Table 3.4: Effects of acids and bases on the universal indicators

Solution	*Colour of universal indicator in the solution*	*pH value*
Sulphuric acid		
Distilled water		

Answer the following questions

1. What is the pH value of;

(a) distilled water?

(b) hydrochloric acid?

(c) sodium hydroxide?

2. Which substances are:

(i) acidic.

(iii) basic.

(ii) neutral.

3. Give the range of pH values for:

(a) acidic solutions.

(b) basic solutions.

Most pH indicator papers are sold with a pH colour Chart as the one shown here.

Discussion

The pH values of acids range from zero to values just less than seven. Solutions of sulphuric acid, hydrochloric acid and nitric acid have pH values which range between 0 to 4 and are **strong acids.**

Lemon juice and ethanoic acid have pH value which range between 4 and 7, and are **weak acids.**

As the pH values decreases from 7 to 0, the strength of the acids increases. Distilled water and sodium chloride solution have a pH value of 7. They are neither acidic nor basic and are **neutral.**

The pH values of bases range between 7 and 14. Solutions of ammonia, calcium hydroxide and sodium hydrogen carbonate have pH values between 7 and 10. They are **weak bases.**

Solutions of wood ash, soap, and sodium hydroxide have pH values ranging from 10 to 14. They are **strong bases.** The pH scale is shown in figure 3.1.

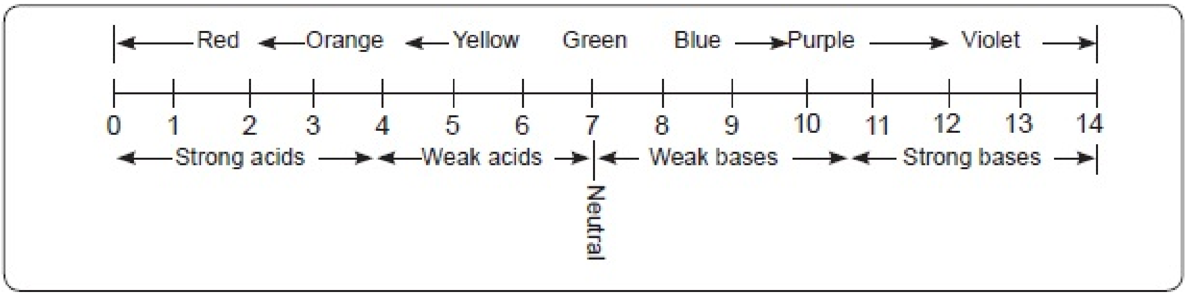


Fig 3.1: The pH scale.

Properties of Acids

Reactions of Acids with Metals

Experiment 3.4: What happens when dilute acids come into contact with metals?

Put a granule of zinc in two test-tubes. Add $2cm^3$ of dilute hydrochloric acid to the first test-tube as shown in figure 3.2.

Record your observations as shown in table 3.5. Repeat the procedure using dilute sulphuric acid in place of dilute hydrochloric acid. Repeat the experiment using clean magnesium ribbon, aluminium foil, iron filings, lead and copper turnings instead of zinc.

CAUTION

Potassium, sodium and calcium should not be reacted with acid.

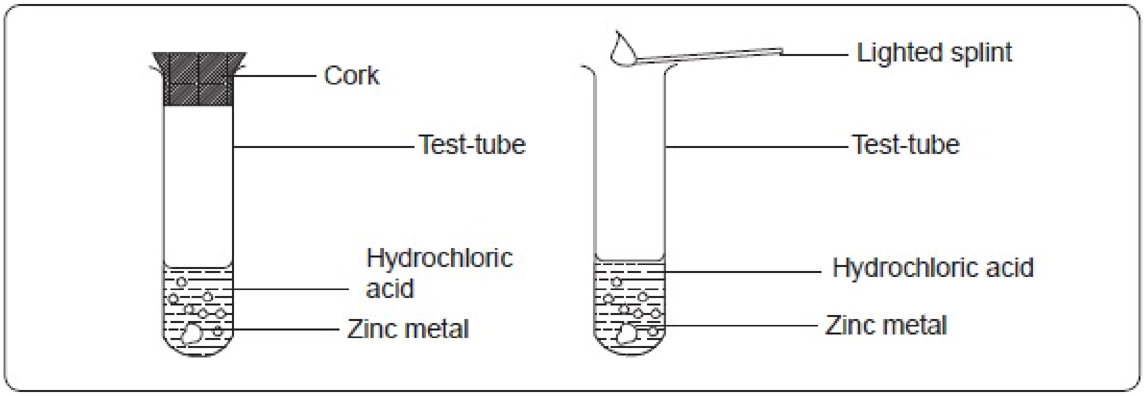


Fig 3.2: Reactions of acids with metals.

Answer the following questions

1. What is observed when dilute acids react with some metals?
2. How is the gas identified?
3. Name the metal or metals that did not react with both acids.
4. Write word equations for the reactions that took place in the experiment.

Table 3.5: Effects of acids on metals

Metal	*Observation with dilute*	
	Hydrochloric acid	*Sulphuric acid*
Zinc		
Magnesium		

NOTE
The intensity of the "pop" sound depends on the purity and amount of the gas mixed with air.

Discussion

When the dilute acids are added to zinc, magnesium, aluminium and iron, bubbles of a colourless gas are evolved. The production of bubbles of a gas is referred to as **effervescence.** The gas produced is hydrogen gas. The gas is tested using a burning splint. A mixture of hydrogen and air burns with a 'pop' sound. This is the test for hydrogen gas.

When dilute sulphuric acid is added to zinc granules, hydrogen gas and zinc sulphate are produced. The reaction can be represented using a word equation.

$$\underset{\text{(metal)}}{\text{Zinc}} + \text{sulphuric acid} \longrightarrow \underset{\text{(salt)}}{\text{Zinc sulphate}} + \underset{\text{(gas)}}{\text{hydrogen}}$$

Similarly, zinc reacts with dilute hydrochloric acid to form zinc chloride and hydrogen gas.

$$\underset{\text{(metal)}}{\text{Zinc}} + \text{hydrochloric acid} \longrightarrow \underset{\text{(salt)}}{\text{Zinc chloride}} + \underset{\text{(gas)}}{\text{hydrogen}}$$

Magnesium and aluminium react more vigorously with the acids than zinc.

Magnesium + hydrochloric acid ⟶ Magnesium chloride + Hydrogen
(metal) (salt) (gas)

Aluminium + hydrochloric acid ⟶ Aluminium chloride + hydrogen
(metal) (salt) (gas)

An acid is a compound that reacts with metals to form a salt and hydrogen gas.

Metal + Acid ⟶ Metal salt + Hydrogen

Very reactive metals like potassium, sodium and calcium react violently with acids. The reaction of calcium with dilute sulphuric acid slows down and eventually stops due to the formation of insoluble calcium sulphate. The insoluble salt coats the metal and prevents further reaction.

Lead reacts slowly with both hydrochloric and sulphuric acids but each reaction eventually stops due to the formation of insoluble coating of lead chloride and lead sulphate respectively. Copper does not react with either dilute hydrochloric acid or dilute sulphuric acid. The name of the salt can easily be derived from the name of the acid as shown in table 3.6.

Table 3.6: Salts produced by magnesium with different acids

Dilute acid	*Metal*	*Name of salt*
Hydrochloric acid		Magnesium chloride
Sulphuric acid		Magnesium sulphate
Nitric acid	Magnesium	Magnesium nitrate
Carbonic acid		Magnesium carbonate
Phosphoric acid		Magnesium phosphate

Reaction of acids with carbonates and hydrogen carbonates

Experiment 3.5: How do dilute acids react with carbonates and hydrogen carbonates

Add a spatulaful of sodium carbonate into a test-tube containing about 5cm^3 of dilute hydrochloric acid. Set up the apparatus as shown in figure 3.3.

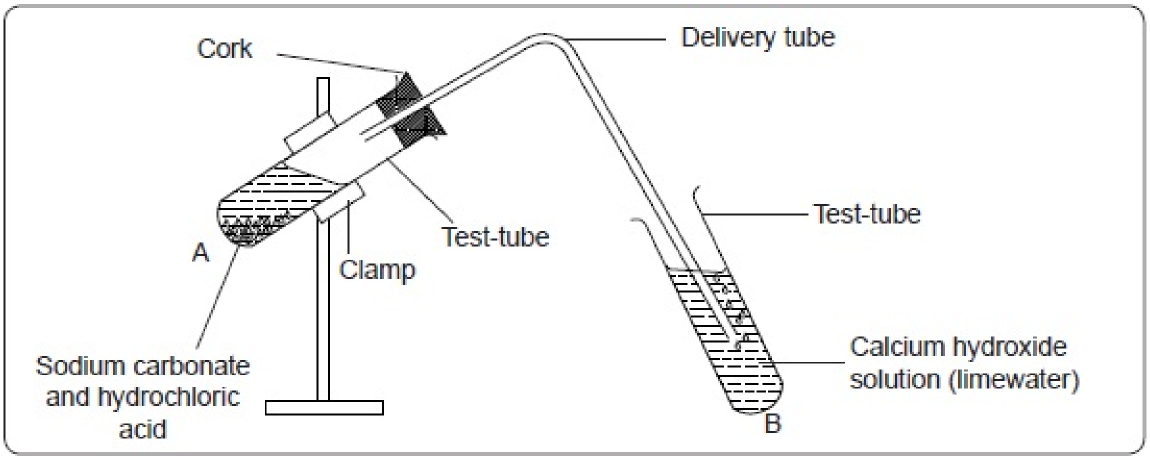


Fig 3.3: Action of acids on carbonates and hydrogen carbonates.

Record the observation as shown in table 3.7. Repeat the procedure using dilute sulphuric acid and dilute nitric acid in place of hydrochloric acid. Record the observations. Repeat the procedure using calcium carbonate and sodium hydrogen carbonate in place of sodium carbonate.

Table 3.7: Action of acids on carbonates and hydrogen carbonates

Metal carbonate or hydrogen carbonate	*Reaction with:*		
	Hydrochloric acid	Nitric acid	Sulphuric acid
Sodium carbonate			

Answer the following questions

1. State the observations made in test tube A and B.
2. Which gas is involved in each case.
3. Explain the observations made when dilute sulphuric acid reacts with calcium carbonate.
4. Write word equations for each of the reactions taking place.

Discussion

When carbonates or hydrogen carbonates are added to dilute acids, effervescence occurs as the solid is used up. A colourless gas is produced. The colourless gas produced is carbon (IV) oxide. When the carbon (IV) oxide gas is bubbled through calcium hydroxide solution (lime water), a white precipitate is formed. This is the test for carbon (IV) oxide. The white precipitate is calcium carbonate.

NOTE

If carbon (IV) oxide gas is bubbled through calcium hydroxide for a long time, the white precipitate dissolves. This is due to the formation of the soluble calcium hydrogen carbonate.

Acids react with carbonates and hydrogen carbonates to produce a salt, water and carbon (IV) oxide (carbon dioxide). The following word equations represent some of the reactions between acids and carbonates or hydrogencarbonates.

Sodium carbonate + hydrochloric acid $\longrightarrow$ Sodium chloride + water + carbon (IV) oxide

Calcium carbonate + nitric acid $\longrightarrow$ Calcium nitrate + water + carbon (IV) oxide

Copper carbonate + Sulphuric acid $\longrightarrow$ Copper Sulphate + water + carbon (IV) oxide

Sodium hydrogen carbonate + hydrochloric acid $\longrightarrow$ Sodium chloride + water + carbon (IV) oxide

The general equations are;

Acid + carbonate $\longrightarrow$ Salt + Water + Carbon (IV) Oxide

Acid + Hydrogen carbonate $\longrightarrow$ Salt + Water + Carbon (IV) Oxide

The reaction between sulphuric acid and calcium carbonate stops after a short while due to the formation of an insoluble layer of calcium sulphate which stops further reaction.

Calcium carbonate + Sulphuric acid $\longrightarrow$ Calcium sulphate (solid) + water + carbon (IV) oxide

Reactions of Acids with Bases

Experiment 3.6: How do dilute acids react with bases?

(a) Measure 10cm^3 of dilute sodium hydroxide solution and put it in a clean conical flask. Add two to three drops of phenolphthalein indicator. Add dilute hydrochloric acid drop by drop, while shaking the conical flask, count the number of drops until the indicator just changes colour. Repeat without adding the indicator. Add the same number of drops of hydrochloric acid to 10 cm^3 of dilute sodium hydroxide. Put the resulting solution in an evaporating dish and heat the solution to saturate it. Allow the saturated solution to cool. Record your observations.

(b) Place a small sample of calcium oxide into a test-tube. Add to it $5cm^3$ of dilute nitric acid. Shake the mixture. Repeat the procedure using the following solids in different test-tubes. Zinc oxide, copper (II) oxide, magnesium oxide and lead (II) oxide. Record your observation as shown in table 3.8. Repeat the experiment using dilute hydrochloric acid in place of dilute nitric acid.

NOTE

Warming increases rate of reaction between the reactants. Therefore where a reaction does not seem to occur, the mixture can be warmed gently to encourage reaction.

Table 3.8: Reaction of acids with bases

Name of solid	*Observation on adding*	
	Dilute nitric acid	*Dilute hydrochloric acid*
Calcium oxide		
Zinc oxide		
Copper (II) oxide		
Magnesium oxide		
Lead (II) oxide		

Answer the following questions

1. (a) Why was it necessary to use an indicator in experiment 3.6 (a)?
   (b) Why was the indicator not used subsequently? Explain.
2. (a) What colour change is observed at the end of the reaction?
   (b) Write a word equation for the reaction.
4. What name is given to the type of reaction between acids and bases?
5. Write word equations for the reactions between metal oxides and the acids.

Discussion

Phenolphthalein indicator is used to determine the end of the reaction, **end point.** The indicator is used to determine the quantity of acid needed to react with $10cm^3$ of dilute sodium hydroxide. It is not used in the next reaction to avoid contamination of the product. The colour of phenolphthalein is pink in basic solution. At the end point, the colour of the indicator in the solution turns from pink to colourless.

NOTE
End point is attained when one of the reactants has been used up.

When acids react with bases, they form a salt and water as the only products. Salt and water are neutral products hence the reaction is referred to as a **neutralisation reaction.**

Sodium hydroxide reacts with dilute hydrochloric acid to form sodium chloride and water only.

$$\underset{\text{(base)}}{\text{Sodium hydroxide}} + \underset{\text{(acid)}}{\text{Hydrochloric acid}} \longrightarrow \underset{\text{(salt)}}{\text{Sodium chloride}} + \text{water}$$

The following are the word equations for the reactions between some metal oxides and acids.

Calcium oxide + Hydrochloric acid $\longrightarrow$ Calcium chloride + water

Zinc oxide + Hydrochloric acid $\longrightarrow$ Zinc chloride + water

Magnesium oxide + nitric acid $\longrightarrow$ Magnesium nitrate + water

Lead (II) oxide + nitric acid $\longrightarrow$ Lead nitrate + water

Copper (II) oxide + sulphuric acid $\longrightarrow$ copper sulphate + water

A base is a substance which when reacted with acids forms salt and water as the only products. Metal oxides, metal hydroxides and ammonia solution are bases. The table below summarises the properties of acids and bases.

Table 3.9: Properties of acids and bases

Acids	*Bases*
Have a sour taste Have pH values below 7	Have a bitter taste Have pH values above 7

Turn litmus red Turn phenolphthalein colourless Turn Methyl orange pink React with bases to form salt and water only React with carbonates to form salt, water and carbon (VI) oxide React with metals to produce salt and hydrogen gas	Turn Litmus blue Turn phenolphthalein pink Turn methyl orange yellow React with acids to form salt and water only.

Effects of Acids on Substances

Industries emit several gases and waste products into the environment leading to environmental pollution. Some of these gases dissolve in rain water to form **acid rain.** This rain reacts with stone work, iron roofs and other metallic surfaces causing damage. This effect is called **corrosion.** When these acidic gases in the atmosphere are inhaled they cause respiratory disorders.

A badly corroded wall, could this be the effect of pollution.

Soil may become acidic due to leaching and water logging. Such soil is unsuitable for the growth of many plants.

Applications of Acids and Bases

Uses of acids

Carbonic Acid is used in aerated drinks to enhance taste.

Hydrochloric acid is used to clean metal surfaces.

Sulphuric acid is used in car batteries, manufacture of ferterlizers, etching of metals, manufacture of paints and detergents.

Nitric acid is used to manufacture of dyes, paints, explosives and fertilizers.

Ethanoic acid and citric acid are used as a flavour in foods.

Table 3.11: Uses of bases

Base	*Use*
Magnesium oxide and hydroxide	Manufacture of anti-acid tablets. Lining of furnaces.
Calcium oxide and hydroxide	Neutralising soil acidity and industrial wastes. Making cement and concrete. Manufacture of toothpaste.
Sodium hydroxide	Manufacture of soap. As a degreasing agent.
Ammonia solution	As a degreasing agent. Manufacture of fertilizers. Manufacture of nitric acid.

Summary

1. An indicator is a substance, which gives a definite colour in an acidic solution and a different definite colour in a basic solution.
2. Universal indicator is a mixture of indicators, which shows strengths of acids or bases.
3. Acids react with some metals to produce salt and hydrogen gas.
4. Ammonia solution, metal oxides and metal hydroxides are bases.
5. Acids react with bases to produce salt and water only. This reaction is called neutralisation.
6. Acids react with carbonates or hydrogen, carbonates to form salt, water

and carbon (IV) oxide.

Revision Exercise

1. (a) What are acid-base indicators?
   (b) Give three examples of commercial acid-base indicators and state the colours they show in acidic and basic solutions.
   (c) What are the advantages of the universal indicators over other acid-base indicators?
2. (a) What is a pH scale?
   (b) State whether solutions with the following pH values are acidic, basic or neutral:
   pH = 3, pH = 6, pH = 2, pH = 12, pH = 7, pH = 8
   (c) Which of the following pH values listed above is of:
   (i) a strong acid?
   (ii) a weak base?
   (iii) a strong base?
   (iv) a weak acid?
3. (a) Write word equations for the reaction between dilute hydrochloric acid and each one of the following:
   (i) zinc metal.
   (ii) calcium hydrogen carbonate.
   (iii) magnesium oxide.
   (iv) potassium hydroxide.
   (b) Which of the reactions are neutralisation reactions?
4. Dilute sulphuric acid was added to a compound of magnesium P. The solid reacted with the acid to form a colourless solution, Q and a colourless gas R which formed a white precipitate when bubbled through lime water.
   (a) Name:
   (i) compound P.
   (ii) solution Q.
   (iii) colourless gas R.
   (b) Write a word equation for the reaction that took place.
   (c) State the observations that would be made if a similar compound of

calcium was used instead of magnesium. Explain.

CHAPTER FOUR
AIR AND COMBUSTION

Air sustains life on earth. Living organisms need air for respiration. Plants need air for photosynthesis. Air is required for combustion of fuels to give energy.

By the end of this chapter , you should be able to:

- State the composition of air.
- Determine experimentally the percentage of oxygen in air by volume.
- Describe fractional distillation of air.
- Define combustion.
- Investigate the conditions for rusting and state the composition of rust.
- State the methods of preventing rusting.
- Prepare and investigate the properties of oxygen.
- Experimentally compare the rates of combustion of elements in air and in oxygen.
- State the nature of the products of burning elements in air and in oxygen.
- State the pollution effects of combustion.
- State the applications of the reactivity series.

Composition of air

Air is a mixture of gases such as oxygen, nitrogen, carbon (IV) oxide and water vapour. Table 4.1 shows the approximate percentage composition by volume of air.

Table 4.1 Approximate percentage composition of air

Component	*Percentage composition*
Nitrogen	78.1
Oxygen	20.9
Carbon (IV) oxide	0.03

Noble gases	0.97
Water vapour	Variable
Dust	Variable

Experiment 4.1 (a): What percentage of air supports combustion

Put dilute sodium hydroxide solution in a trough. Place a small candle on a cork and float it on the solution as shown in figure 4.1 (a) to (c). Cover it with a gas jar. Mark on the gas jar, the level of the solution. Measure the height of the air column and record it.

Remove the jar and light the candle. Gently cover the burning candle with the gas jar. After the candle has gone off leave the apparatus to cool to room temperature. Mark on the gas jar the final level of the solution. Measure and record the height of the air column once more.

Remove the gas jar and measure the change in height. Record all your observations.

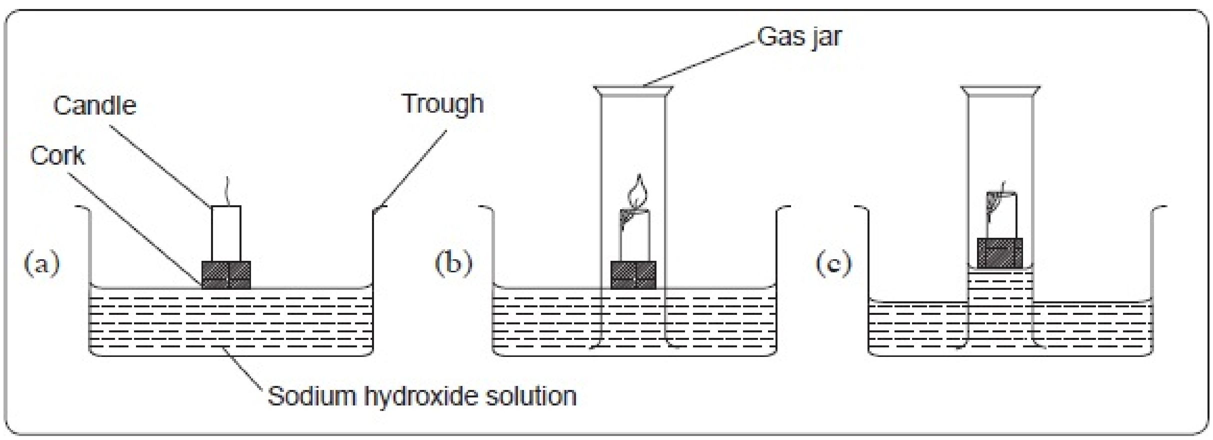


Fig 4.1: Measurement of the amount of air that supports combustion.

Answer the following questions

1. Why does the candle go off after burning for some time?
2. Explain why the level of dilute sodium hydroxide solution rises in the gas jar after the candle goes off.
3. What was the length of the air column in the gas jar before and after burning?

4. Determine the percentage of air used up by the burning candle.
5. Why is sodium hydroxide preferably used in the experiment instead of water?
6. State the sources of errors in the experiment.
7. Why is it necessary to leave the apparatus to cool before taking the final reading?

NOTE
A thin candle of length 3cm produces better results in the experiment above.

Discussion

The candle burns for a while then it goes off. As the candle burns, it uses up the active part of air in the fixed amount of air enclosed in the gas jar. This leaves a partial vacuum in the jar. Greater atmospheric pressure acting on the surface of the sodium hydroxide forces the solution up into the jar. The following are sample results for a similar experiment.

Height of air column before burning = 16.0cm

Height of air column after burning = 12.9cm

Height of air used during burning = 3.1cm

Percentage of air by volume used up–

$$= \frac{\text{Height of air used}}{\text{Initial height of air}} \square 100$$

$$= \frac{3.1}{16.0} \square 100$$

$$= 19.375\%$$

Combustion or burning is a process in which a substance combines with oxygen with the production of heat. The part of air that supports combustion is active air. The active part is oxygen, which forms about 20% of dry air by volume. The part of air that remains in the gas jar does not support combustion. The component of air that is inactive is mainly nitrogen.

The experimental result is not the same as the theoretical value of the percentage of oxygen in air by volume. This is due to experimental error, which may result from:

- The sodium hydroxide solution may not absorb all the carbon (IV) oxide gas.
- The candle may go off before all the oxygen is used up due to the build-up of

carbon (IV) oxide levels.

Dilute sodium hydroxide is preferably used instead of water to absorb carbon (IV) oxide that was initially in the gas jar and that which is produced during combustion.

Heating causes expansion of gases therefore the apparatus should be allowed to cool before the final reading is taken.

NOTE
For this experiment, the combustion tube can be made from an ignition tube.
CAUTION
Always handle glass wool with a pair of tongs.

Experiment 4.1 (b): What proportion of air is used up when copper is heated in a fixed volume of air?

Pack copper turnings in a long hard glass tube about 6cm long. Connect the tube with two glass syringes as shown in figure 4.2 with one syringe containing a specific volume of air while the other is empty.

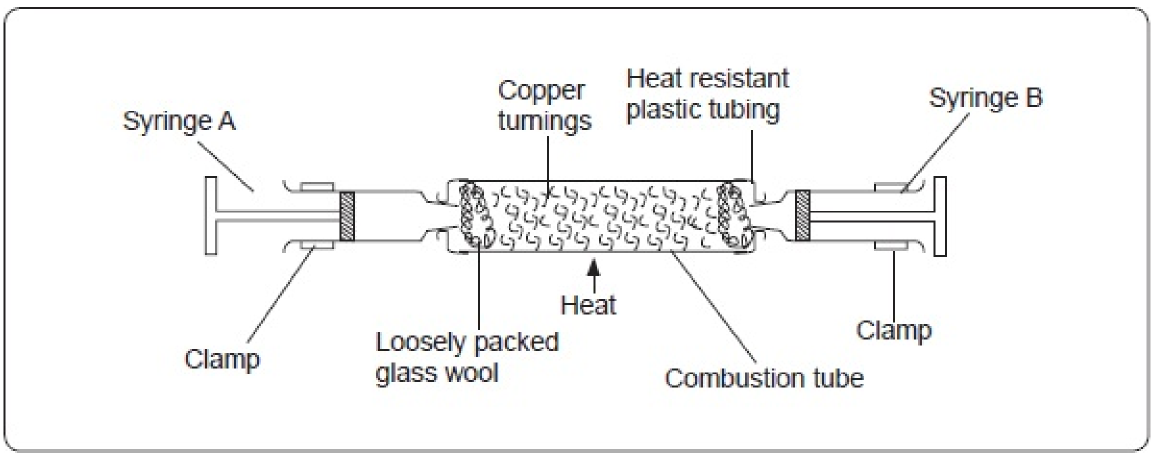


Fig 4.2: Determination of the active part of air using copper.

Heat the copper turnings until they are red hot. Slowly pass the air from syringe A through the hot turnings to syringe B and back. Repeat this process while heating the copper turnings until the new volume of air in syringe A is constant. Allow the glass tube to cool and record the volume of the gas in syringe A.

Answer the following questions

1. What is the use of the glass wool plug in the experiment?

2. Why is the air passed through copper turnings
   (a) slowly?
   (b) repeatedly until there is no further change in volume?
3. (a) What is the volume of the air in syringe A;
      (i) before heating?
      (ii) after heating?
   (b) Use the data obtained to calculate the;
   (i) volume of air used up during the experiment.
   (ii) percentage by volume of air used during the experiment.
4. Why does the volume of air in the syringe decrease?
5. State the sources of error in this experiment.

Discussion

Copper is a red-brown metal. When it is heated in air, it turns black. This is because it combines with oxygen to form black copper (II) oxide. Below is a word equation for the reaction.

$$\underset{\text{(red-brown solid)}}{\text{Copper}} + \underset{\text{(colourless gas)}}{\text{Oxygen}} \longrightarrow \underset{\text{(black solid)}}{\text{Copper (II) oxide}}$$

The following are sample results for the experiment:

Volume of air in syringe A before heating = $7.5cm^3$

Volume of air in syringe A after heating = $6.0cm^3$

Volume of air used up during heating = $1.5cm^3$

$$= \frac{\text{Volume used during heating}}{\text{Original volume of air}} \square 100 = \frac{1.5 \square 100}{7.5} = 20\%$$

About 20% by volume of air is used during combustion and the 80% of air left does not react with heated copper.

In this experiment, the glass wool plug is used to stop the copper turnings from being sucked into the syringes. The air is passed repeatedly over heated copper to ensure that all oxygen in the syringes and tube is used up. The air is passed slowly to allow enough time of contact between the reactants. The gas left in the syringe does not react with copper. It is mainly nitrogen.

The possible sources of error in this experiment include:

(a) The air initially present in the tube is not accounted for.
(b) There is possible leakage of air.
(c) Not all the oxygen may have been used up.

Experiment 4.1 (c): What percentage of air is used when iron filings rust?

Wet a measuring cylinder and sprinkle some iron filings on the wet surface. Remove the excess iron filings. Invert the measuring cylinder in a trough of water. Read the volume of air column in the measuring cylinder. Leave the set up for 48 hours. Read and record the volume of the air column. Record all your observations.

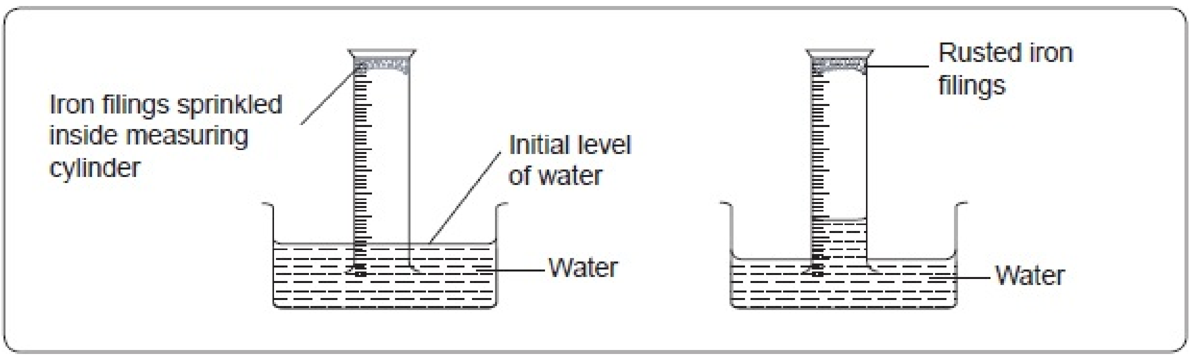


Fig 4.3: Determination of the amount of air used up during rusting of iron.

Answer the following questions

1. Why is the measuring cylinder made wet before sprinkling the iron filings?
2. State the observations made after 48 hours.
3. What percentage volume of air is used up in the experiment?

Discussion

The measuring cylinder is made wet to ensure that the iron filings stick onto the wet surface. When iron filings are left for 48 hours in the measuring cylinder, a brown coating is formed on the filings. The brown coating is rust. Rust is a compound of iron and oxygen. During rusting, oxygen is used and therefore water rises up in the measuring cylinder to replace the volume of air used during rusting. About 20% of air by volume is used up during rusting.

CAUTION
Avoid getting into contact with phosphorous or inhaling its fumes.

Experiment 4.1 (d): What percentage of air is used when white phosphorus smoulders?

Invert an empty measuring cylinder in a trough of water. Record the volume of the air column. Cut a small piece of white phosphorous under water. Attach the piece of white phosphorous to the end of a piece of copper wire. Arrange the apparatus as shown in the diagram, figure 4.4.

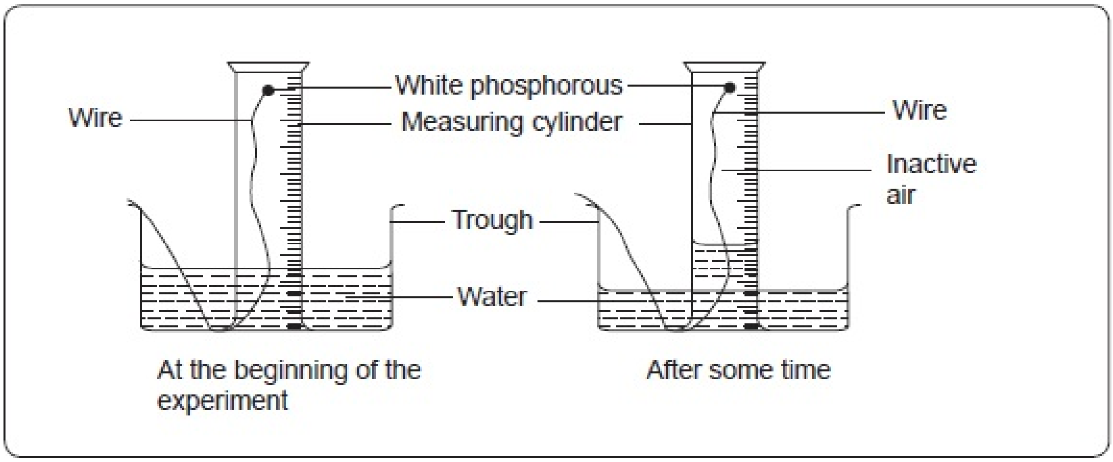


Fig 4.4: Measuring the amount of air used by smouldering white phosphorous.

Leave the set-up overnight and record the volume of the air column. Record your observations at the end of the experiment.

Answer the following questions

1. What is observed when phosphorous is exposed to air?
2. What is the percentage volume of air used up in this experiment?
3. Why is phosphorus stored under water?

Discussion

Yellow and white phosphorus smoulder in air. This is because phosphorus reacts spontaneously with oxygen to form a mixture of oxides. This explains why phoshorous is stored under water as it does not react with water. The reaction can be represented by the following word equations.

$$\text{Phosphorous + oxygen} \longrightarrow \text{phosphorous (III) oxide}$$

$$\text{Phosphorous + oxygen} \longrightarrow \text{phosphorous (V) oxide}$$

After 24 hours, the water level inside the measuring cylinder will have risen to occupy the volume of oxygen used up. The difference in volume can be used to calculate the percentage of oxygen by volume in air.

CAUTION

The white phosphorous should not be allowed to come in contact with the walls of the measuring cylinder because it stops smouldering.

Experiment 4.2: How can the presence of Carbon (IV) oxide and water in air be established?

(a) Place $2cm^3$ of fresh calcium hydroxide solution (lime water) in a boiling tube. Pass water slowly from a tap into an aspirator as shown in figure 4.5. Record your observations.

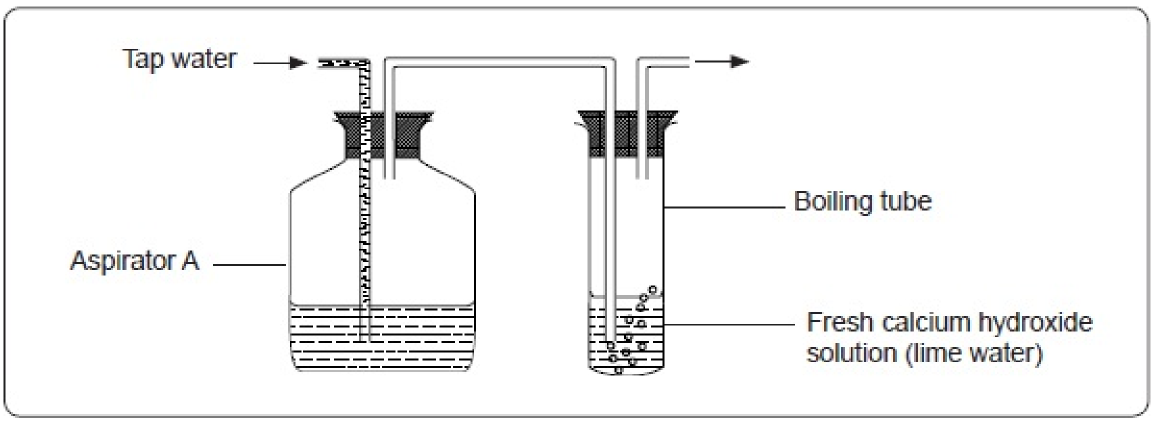


Fig 4.5: Demonstration of the presence of carbon (IV) oxide in air.

(b) Pack the bottom of a U-tube with anhydrous calcium chloride. Arrange as shown in figure 4.6. Pass air through the U-tube by means of an aspirator or a suction pump. Record your observations.

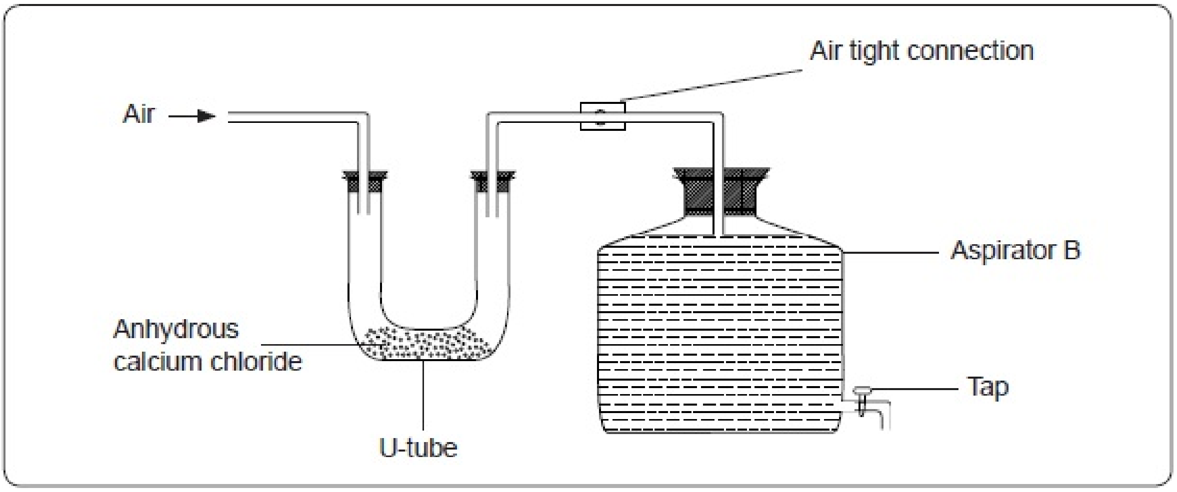


Fig 4.6: Demonstration of the presence of water vapour in air.

Answer the following questions

1. Why is it necessary to allow water to flow into the aspirator in experiment 4.2?
2. Explain what was observed when air was bubbled through lime water?
3. Why is it necessary to allow the water to flow out of the aspirator in experiment 4.2?
4. What happens when air is passed through anhydrous calcium chloride?

NOTE

A deliquescent substance is one which absorbs water from the atmosphere and dissolves.

Discussion

Water is allowed to flow into aspirator A to drive out air and bubble it through the calcium hydroxide solution. When the stream of air is passed through calcium hydroxide solution, a white precipitate is formed. This indicates that carbon (IV) oxide gas is present in air. Water is allowed to flow out of aspirator B in order to create a suction force which draws air through the U-tube. Anhydrous calcium chloride is a white solid. When air is passed through the U-tube the anhydrous calcium chloride absorbs water vapour from the air and becomes wet. It may form a colourless solution depending on the amount of moisture in the air. Substances, which absorb moisture from the air to form a solution are called **deliquescent substances.** Other deliquescent substances are anhydrous iron (III) chloride, magnesium chloride and zinc chloride.

Fractional Distillation of Liquefied Air

As already established, air is a mixture of many gases. It can be separated into its constituent gases by fractional distillation.

The air is first purified by passing it through filters to remove dust. The dust-free air is then passed through a solution of concentrated sodium hydroxide to remove carbon (IV) oxide gas. The remaining part of air is then cooled to -25°C to remove water vapour, which solidifies out as ice. The remaining part of air is then compressed to a pressure of 200 atmospheres and allowed to expand. Repeated compression and expansion of the air cools it to liquid at -200°C.

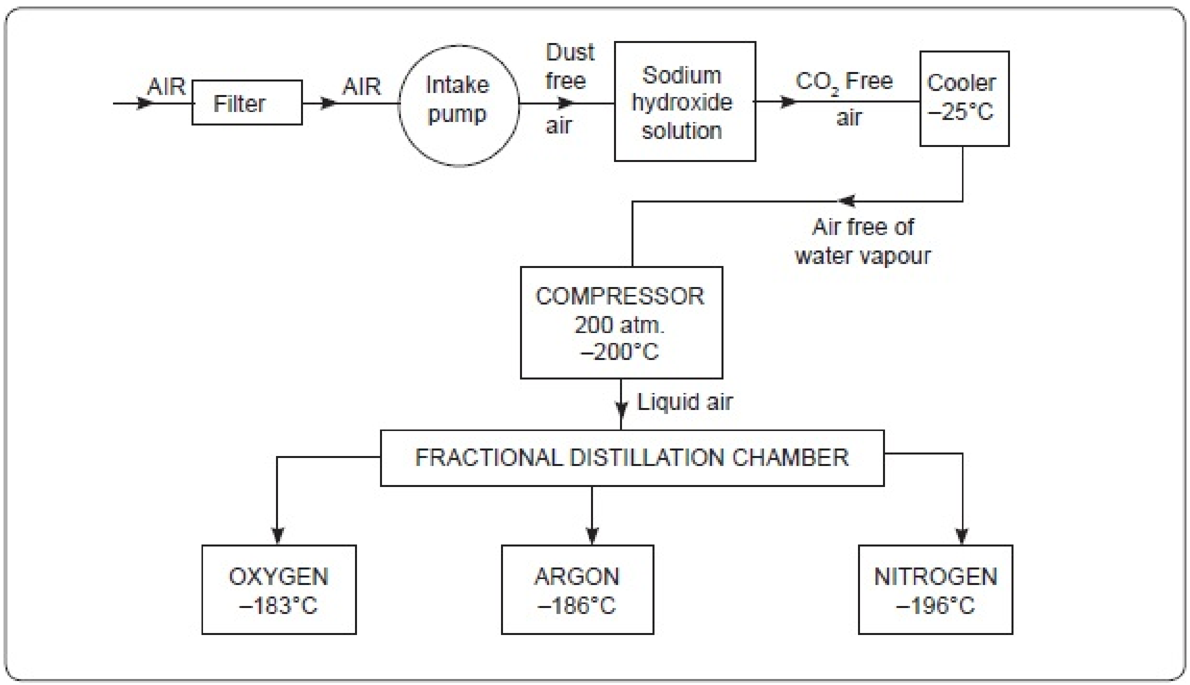


Fig 4.7: Summary of fractional distillation of air.

The liquid air consists of oxygen, nitrogen and noble gases. Since these gases have different boiling points, they can be separated by fractional distillation. Liquid oxygen boils at -183°C and nitrogen at -196°C. Nitrogen distils out first because it has a lower boiling point. The other gases, made of mainly argon, boil at -186°C. They form the second fraction. The argon can be separated from oxygen by further distillation. This process is known as fractional distillation of liquefied air.

NOTE

Corrosion of iron is called rusting. Corrosion of other metals is not refered to as rusting.

Rusting

Have you ever known why old iron nails or sheets look different from new ones? Old iron sheets are brown and dull while new ones are shiny. Many other metals change after sometime due to some chemical reactions, which take place on their surfaces. These reactions are referred to as corrosion. Corrosion of iron due to its reaction with atmospheric oxygen and moisture is called rusting. Rust forms a brown coating on the surface of iron material. Because rust is porous, once an object starts to rust, the process continues until the object is completly destroyed.

Experiment 4.3: What are the conditions necessary for rusting?

Label five boiling tubes 1, 2, 3, 4 and 5 respectively. Put two clean nails in each of the boiling tubes. To the first tube add $10cm^3$ of tap water. To the second tube, add $10cm^3$ of boiling water followed by about $3cm^3$ of oil. To the third tube, push a piece of cotton wool half-way and place some anhydrous calcium chloride on it and cork the tube. To the fourth tube, add salted water. The fifth boiling tube contains nails only. Observe the nails after three days and record your observations.

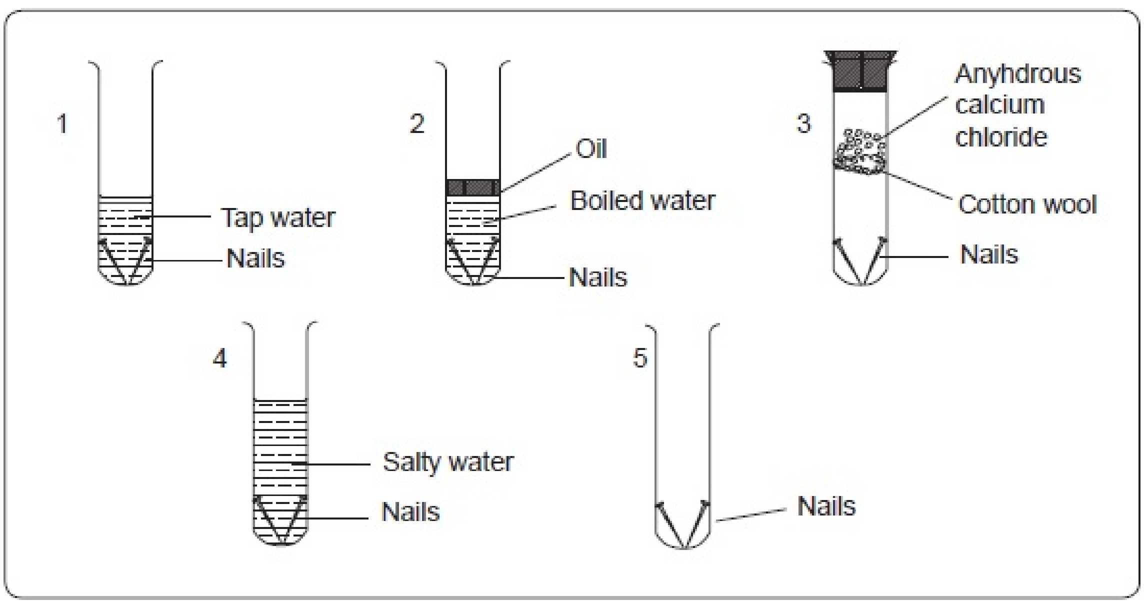


Fig 4.8: Finding out what conditions favour rusting.

Answer the following questions

1. What is observed in each of the tubes after three days?
2. Why is the water in the second tube

(a) boiled?

(b) covered with oil?

3. What is the purpose of anhydrous calcium chloride?
4. Why is it necessary to cork the third tube?
5. List the conditions necessary for rusting to occur.
6. Suggest methods by which rusting can be prevented.

Discussion

It is observed that the nails in boiling tube 1 would have rusted after three days. The nails in tube 5 would have rusted to a smaller extent. The rusting in tube 4 is more intense. No rust is observed in tubes 2 and 3.

Tap water contains dissolved oxygen. The iron nails combine with oxygen in the presence of water to form hydrated iron oxide.

$$\text{Iron + oxygen} \longrightarrow \text{Iron (III) oxide}$$

$$\text{Iron (III) oxide + water} \longrightarrow \text{hydrated iron (III) oxide (brown)}$$

Rusting occurred in tube 1 because both water and oxygen were present. Some rusting occurred in tube 5 since there was some moisture in the air. Rusting was more intense in tube 4 due to presence of dissolved sodium chloride.

When water was boiled, all the dissolved gases in it were expelled. The layer of oil covering the boiled water prevented re-entry of air. There was no oxygen in tube 2 and therefore the nails did not rust. Anhydrous calcium chloride absorbs moisture from the air, thus air in tube 3 is dry. It was necessary to cork tube 3 to prevent entry of water vapour from the atmosphere. The nails in tube 3 do not rust because there was no water in it.

The presence of water and oxygen are thus necessary for iron to rust. Rusting occurs faster in salty or acidic surroundings. For example, cars rust faster in Mombasa than in Nairobi.

A rusted shell of a vehicle consigned to the scrap yard.

Methods of Preventing Rusting

Rusting destroys machinery, equipment and roofs made of iron. Rusting can be very expensive. Prevention of rusting is therefore of great importance. The basis of rust prevention is to keep iron out of direct contact with water and oxygen.

The following methods are widely used to prevent rusting of iron.

(i) Painting e.g. cars, roofs, marine vessels etc.

(ii) Coating with other metals. This can be done through galvanisation or electroplating.

(iii) Alloying: This involves the mixing of iron with one or more metals to produce a substance, which does not rust.

(iv) Oiling and greasing: This method is used in moving engine parts where other methods can not be used due to friction.

(v) Sacrificial protection: In this arrangement, a more reactive metal such as zinc or magnesium is attached to the iron structure. The more reactive metal corrodes instead of iron. The method is applied in ships, water and oil pipes

Oxygen

Oxygen exists freely in the atmosphere as a gas. Its chemical symbol is O. Two atoms of oxygen combine to form a molecule with a chemical formula of O_2. Oxygen is also found combined with other elements such as hydrogen in water and metals in metal oxides. It is the most active component in air.

Experiment 4.4: How is oxygen prepared in the laboratory?

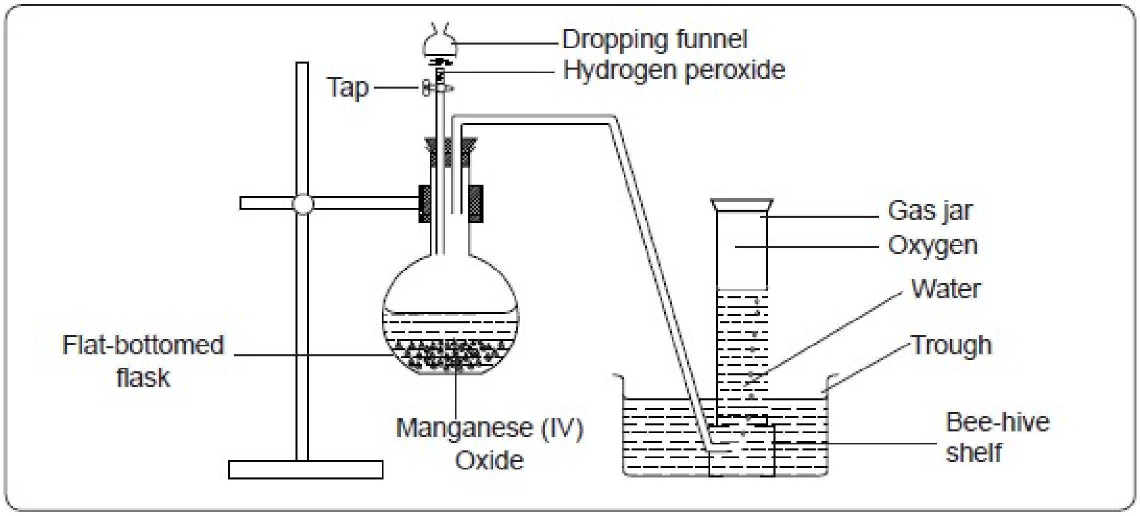


Fig 4.9: Laboratory preparation of oxygen.

Set the apparatus as shown on figure 4.9 and remove the lid. Put about 2g of manganese (IV) oxide into a flat-bottomed flask. By means of a dropping funnel add hydrogen peroxide drop-wise into the flask. Let the first few bubbles escape then collect the gas as shown in figure 4.9.

Note the colour and smell of the gas collected. Lower a glowing splint into a gas jar containing the gas. Record all your observations.

CAUTION

Handle hydrogen peroxide with care it is corrosive.

Answer the following questions

1. Why are the first few bubbles of oxygen produced not collected?
2. What is the colour and smell of oxygen?
3. Why is it possible to collect oxygen over water?
4. What is the function of manganese (IV) oxide in the experiment?
5. Describe how oxygen is tested.

NOTE
Catalysts will only catalyse specific reactions and act best when in powder form.

In the absence of manganese (IV) oxide, the hydrogen peroxide can be warmed to speed up the reaction.

Discussion

Hydrogen peroxide decomposes slowly to produce oxygen and water under normal conditions. On adding manganese (IV) oxide the rate of decomposition is speeded up. Manganese (IV) oxide acts as a **catalyst.** *A catalyst is a substance that alters the rate of a reaction.*

Manganese (IV) oxide is used as a catalyst in the decomposition of hydrogen peroxide.

$$\text{Hydrogen peroxide} \xrightarrow{\textit{Manganese (IV) oxide}} \underset{\text{(gas)}}{\text{Oxygen}} + \text{Water}$$

The first few bubbles of oxygen are not collected because the gas is mixed with air which was originally in the flask. Oxygen is a colourless, odourless gas with a low boiling point of -183°C. It is slightly solube in water and so it can be collected over water.

Oxygen relights a glowing splint. This is the test for oxygen. Oxygen can also be prepared in the laboratory by:

(i) Adding water to Solid Sodium Peroxide using the set-up in figure 4.9

Sodium peroxide + water $\rightarrow$ Sodium hydroxide + Oxygen

(ii) Heating Potassium Manganate (VII) (Potassium Permanganate)

Potassium Heat $\xrightarrow{\textit{Heat}}$ Potassium + Manganese (IV) oxide + Oxygen
Manganate (VII) Manganate (VII)

Burning of Substances in Air

The most familiar chemical reaction of air is burning. There are many substances, which burn in air. It has been established that oxygen is the active part of air, which supports burning.

Experiment 4.5: How do metals burn in air and in oxygen?

Warm a piece of sodium in a deflagrating spoon until it begins to burn. Lower it into a gas jar of air as shown in figure 4.10. Record your observations.

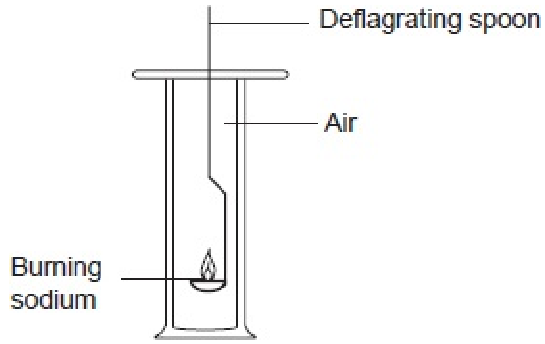


Fig 4.10: Burning metals in oxygen.

NOTE

Use a pair of tongs for heating magnesium, iron and copper.

Allow the gas jar to cool, add some water to the product and shake the mixture. Test any gas given out with moist red and blue litmus paper. Test the solution in the gas jar using litmus papers and record your observations. Repeat the experiment using oxygen instead of air.

Repeat the whole procedure using calcium, magnesium, iron and copper in place of sodium.

Record your observation as shown in table 4.2

Table 4.2: Burning Metals in Oxygen

Metal	*How it burns in air*	*How it burns in Oxygen*	*Appear-ance of product*	*Name of product*	*Solubility of the product in water*	*Effect on litmus paper*
Sodium						
Calcium						
Magnesium						
Iron						
Copper						

Answer the following questions

1. State the difference in the way substances burn in air and in oxygen.
2. Write a word equation for the reaction taking place for each of the metals.
3. Arrange the metals in order of their reactivity.
4. Suggest the name of the products formed when sodium and magnesium burn in air. Explain what happens when water is added to the products.

Discussion

Many metals burn in air and in oxygen at different rates. They burn faster in oxygen than in air. Nitrogen is the component of air, which slows down the rate of burning. When metals burn in oxygen they form metal oxides.

$$\text{Sodium} + \text{Oxygen} \longrightarrow \text{sodium oxide}$$

$$\text{Calcium} + \text{Oxygen} \longrightarrow \text{calcium oxide}$$

$$\text{Magnesium} + \text{Oxygen} \longrightarrow \text{Magnesium oxide}$$

$$\text{Iron} + \text{Oxygen} \longrightarrow \text{Iron (III)oxide}$$

Sodium reacts most vigorously with oxygen while copper is the least reactive. A summary of the observations made when metals burn in air and in Oxygen is given in table 4.3.

Table 4.3: Products of Burning Metals in Oxygen

Metal	*How it burns in air*	*How it burns in Oxygen*	*Appearance of product*	*Name of product*	*Solubility of the product in water*	*Effect on litmus paper*
Sodium	Few sparks. Burns with a yellow-orange flame	Very vigorous Yellow-orange flame	White solid	Sodium oxide and Sodium Nitride	Soluble Alkaline gas evolved	Turns blue
Calcium	Brick-red glow	Burns vigorously with a red flame	White solid	Calcium oxide and Calcium Nitride	Slightly soluble Alkaline gas evolved	Turns blue
Magnesium	Burns brightly with a white flame	Burns more brightly with a white glaring flame	White powder	Magnesium oxide and Magnesium Nitride	Slightly soluble alkaline gas produced	Turns blue
Iron	No apparent burning, glows red with few sparks	Glows more brightly and sparks	red-brown solid	Iron (III) oxide	Insoluble	No effect
Copper	No apparent burning, surface turns black	Burns with a blue flame; surface turns black	Black solid	Copper (II) oxide	Insoluble	No effect

NOTE
Sodium burns in excess oxygen to form sodium peroxide which is a yellow solid.

Reactive metals such as sodium, calcium and magnesium react with nitrogen in the air to form nitrides.

$$\text{Sodium + Nitrogen} \longrightarrow \text{Sodium nitride}$$

$$\text{Calcium + Nitrogen} \longrightarrow \text{Calcium nitride}$$

$$\text{Magnesium + Nitrogen} \longrightarrow \text{Magnesium nitride}$$

When the nitrides react with water, ammonia gas is given out.

$$\text{Sodium nitride + Water} \longrightarrow \text{Sodium hydroxide + Ammonia}$$

$$\text{Calcium nitride + Water} \longrightarrow \text{Calcium hydroxide + Ammonia}$$

$$\text{Magnesium nitride + Water} \longrightarrow \text{Magnesium hydroxide + Ammonia}$$

The reactions in which elements combine with oxygen are referred to as **oxidation.**

The substance to which the oxygen is added is said to have been **oxidised.** The metals can be arranged in order of their rates of reaction with oxygen from the most reactive to the least reactive. This arrangement is referred to as a reactivity series of metals.

Mercury, silver and gold are less reactive than copper and are not easily oxidised.

The following is part of the Reactivity Series for some metals.

Potassium	Most reactive
Sodium	
Calcium	
Magnesium	
Aluminium	
Zinc	
Iron	
Lead	
Copper	
Mercury	
Silver	
Gold	Least reactive

CAUTION

Products of burning sulphur and phosphorus are poisonous and therefore the experiments should be done in a fume cupboard. Phosphorus is also highly flammable.

Experiment 4.6: What happens when non-metallic elements burn in oxygen?

Heat sulphur in a deflagrating spoon until it begins to burn. Lower it into a gas jar of oxygen and record your observations. Add a little water to the product and shake the mixture. Test the resulting solution with litmus paper. Repeat the experiment using carbon (charcoal) and red phosphorus in place of sulphur. Record your observations as shown in table 4.4.

Table 4.4: Burning non-metals in oxygen

Non-Metal	*How it burns in oxygen*	*Name of product*	*Appearance of the product*	*Effect of solution on litmus paper*
Sulphur				
Carbon				
Phosphorus				

Answer the following questions

(i) Write word equations for each of the reactions of the non-metallic elements with oxygen.

(ii) State whether the products of burning the non-metals are acidic or basic. Explain

Discussion

Sulphur burns in oxygen to give a gaseous product which has a choking irritating smell. The product is sulphur (IV) oxide.

$$\text{Sulphur + Oxygen} \longrightarrow \text{Sulphur (IV) oxide}$$

A solution of sulphur (IV) oxide in water is acidic and turns blue litmus paper red. The acid is called sulphuric (IV) acid (sulphurous acid).

$$\text{Sulphur (IV) oxide + Water} \longrightarrow \text{Sulphurous acid.}$$

Oxides which dissolve in water to form acidic solutions are referred to as **acidic oxides.** Table 4.5 shows the summary of the effect of burning some non-metals in oxygen.

Table 4.5: Summary of effect of burning non-metals in oxygen

Non Metal	*How it burns in oxygen*	*Name of product*	*Appearance of product*	*Effect of solution on litmus paper*
Sulphur	Burns with a blue flame	Sulphur (IV) oxide	White fumes	Turns red
Carbon	Burns with a yellow flame	Carbon (IV) oxide	White fumes	Turns red
Phosphorous	Burns with a White flame	Phosphorous (V) oxide and Phosphorous (III) oxide	White fumes	Turns red

The following equations represent the reactions of the non-metals with oxygen.

$$\text{Carbon} + \underset{\text{(excess)}}{\text{Oxygen}} \longrightarrow \text{Carbon (IV) oxide}$$

$$\text{Carbon} + \underset{\text{(limited supply)}}{\text{Oxygen}} \longrightarrow \text{Carbon (II) oxide}$$

$$\text{Phosphorous} + \underset{\text{(excess)}}{\text{Oxygen}} \longrightarrow \text{Phosphorous (V) oxide}$$

$$\text{Phosphorous} + \underset{\text{(limited supply)}}{\text{Oxygen}} \longrightarrow \text{Phosphorous (III) oxide}$$

Some non-metallic elements form oxides which are neither acidic nor basic. These oxides are referred to as **neutral oxides.** Carbon (II) oxide and water (hydrogen oxide) are examples of neutral oxides.

Experiment 4.7: Is there change in mass when a substance burns in air?

Weigh about 1g clean magnesium ribbon in a crucible. Set up the apparatus as shown in figure 4.11.

Heat the crucible, occasionally lifting the lid to let air in. Do not allow any contents to escape from the crucible. When all the magnesium has burned, allow the crucible to cool. Weigh the cool crucible and its contents again. Record your observations as follows:

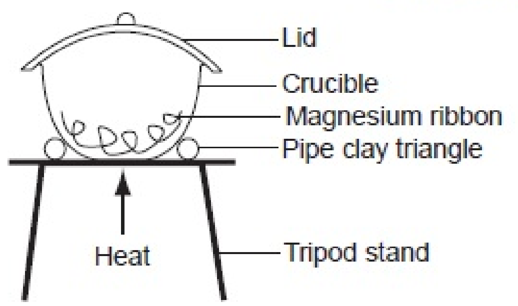


Fig 4.11: Burning substances in air

Mass of crucible + Magnesium before burning = Xg

Mass of crucible + contents after burning = Yg

Change in Mass = (Y–X)g

Answer the following questions

1. Why is it necessary to occasionally allow air into the crucible?
2. Does the mass increase or decrease after the burning? Explain.

Discussion

When magnesium is burned in a closed crucible, most of the oxygen inside is consumed. It is therefore necessary to allow air in so that burning can continue. The mass of the product is more than the original mass of magnesium. This shows that as it burns, magnesium combines with air to form a new product.

When substances burn in air, they combine with oxygen to form oxides. If the product is a solid there is increase in mass. When the product is gaseous there is decrease in mass. The decrease in mass is because the products, being gaseous escape into the air.

Experiment 4.8: How do elements compete for combined oxygen?

Place a spatulaful of magnesium oxide in a bottle top. Add magnesium powder and mix well. Heat and record your observations. Repeat the experiment using zinc oxide, iron (III) oxide, Lead oxide and copper (II) oxide. Repeat the experiment using lead, copper, zinc and iron in place of magnesium. Record the results shown in table 4.6.

Table 4.6: Competition for oxygen

Metal / *Metal Oxide*	*Magnesium*	*Zinc*	*Iron*	*Lead*	*Copper*
Magnesium oxide					
Zinc oxide					
Iron (III) oxide					

Answer the following questions

1. (a) Name the metals displaced from their oxides by:

(i) magnesium (ii) zinc (iii) iron (iv) lead (v) copper

(b) Arrange the metals in order of their decreasing reactivity with combined oxygen.

Discussion

Magnesium combines with oxygen more readily than copper. Therefore, magnesium removes combined oxygen in copper (II) oxide to form magnesium oxide. Copper is said to have been displaced by magnesium. Copper on the other hand does not remove combined oxygen from the oxides of magnesium, lead, zinc and iron. This is due to the fact that copper reacts with oxygen less readily than these metals. It is the least reactive.

A more reactive metal removes combined oxygen from a metal oxide of a less reactive metal. More reactive metals displace less reactive metals from their oxides. Table 4.7 shows a summary of the competition for combined oxygen.

Table 4.7: Summary of the competition for combined oxygen by elements

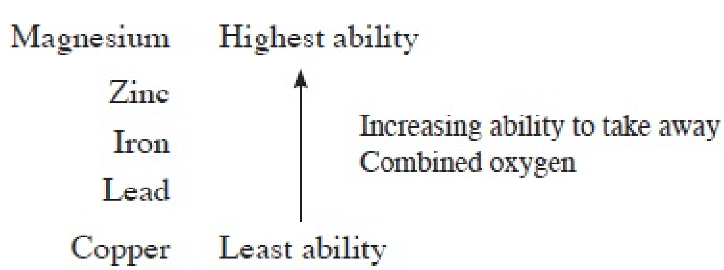
	Magnesium	*Lead*	*Iron*	*Zinc*	*Copper*
Magnesium oxide (white)	No reaction	No reaction	No reaction	No reaction	No reaction
Copper (II) oxide (black)	Magnesium oxide and copper formed	Lead (II) oxide and copper formed	Ion (II) oxide and Copper formed	Zinc oxide and copper formed	No reaction
Lead (II) oxide (yellow)	Magnesium oxide and lead formed	No reaction	Iron (III) oxide and Lead formed	Zinc oxide and Lead oxide	No reaction
Zinc oxide (white)	Magnesium oxide and Zinc formed	No reaction	No reaction	No reaction	No reaction
Iron (III) oxide	Magnesium oxide and Iron formed	No reaction	No reaction	Zinc oxide and Iron formed	No reaction

From the table, magnesium displaces four metals from their oxides. Zinc displaces three while lead and iron displaces two and one respectively. Copper displaces none. Therefore, magnesium is the most reactive while copper is the least reactive.

By considering the number of metals displaced, a reactivity series similar to the one developed earlier can be obtained.

NOTE

A reaction in which both reduction and oxidation occur simultaneously is called a REDOX reaction.



Removal of oxygen from a substance is called **reduction.** When a metal oxide loses oxygen, it is said to have been **reduced.** The metal, which gains oxygen is said to have been **oxidised.**

$$\text{Zinc + Copper (II) oxide} \longrightarrow \text{Zinc oxide + Copper}$$

In the above equation, zinc is oxidised while copper oxide is reduced. Both reduction and oxidation take place simultaneously.

Application

The extraction of metals from their ores uses the concept of reduction. The ores that contain the metal oxides are reduced by more reactive metals. For example, Aluminium is used to reduce iron (III) oxide by the thermite process.

Carbon, a non-metal can remove combined oxygen from some metal oxides such as iron (III) oxide and copper (II) oxide.

$$\text{Carbon + Copper (II) oxide} \longrightarrow \text{Carbon (IV) oxide + Copper}$$

The ability of carbon to reduce some metal oxides is applied in the extraction of metals such as copper and zinc from their ores.

NOTE
A pollutant is a substance (contaminant) or form of energy which has harmful effects to the enviroment.

Atmospheric Pollution

Human activities have changed the composition of air in some places. Gases such as carbon (IV) oxide, carbon (II) oxide, sulphur (IV) oxide and phosphorous (V) oxide, are examples of harmful substances emitted into the atmosphere mainly from the combustion of fossil fuels. These gases cause pollution of the atmosphere. For example, sulphur (IV) oxide dissolves in rain water and is converted to sulphurous acid, which forms “acid rain”. Acid rain destroys plants and aquatic life. It also corrodes iron sheets, zinc roofing and buildings.

Uses of Oxygen

1. Air enriched with oxygen is used in hospitals by patients with breathing difficulties.
2. When mixed with helium it is used by mountain climbers and deep-sea

divers.

3. Oxygen is used to burn fuels such as those used for propelling rockets.
4. A mixture of oxygen and acetylene burns to produce a very hot flame used in welding and for cutting metals.
5. During steel making, oxygen is used to remove iron impurities.
6. Oxygen is used as one of the reactants in fuel cells.

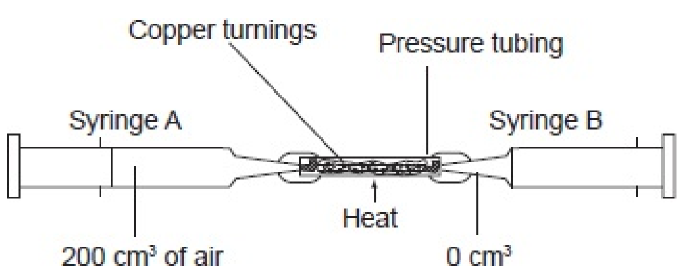
Deep sea divers inspect fish and corals.

Summary

1. Air is a mixture of gases.
2. Oxygen and water are necessary for iron to rust. Rust is mainly hydrated Iron (III) oxide.
3. Atmospheric pollution is caused by presence of harmful gases in the air.
4. A catalyst alters the rate of a reaction.
5. (a) Most metals burn in oxygen to form basic oxides.
   (b) Most non-metals burn in oxygen to form acidic oxides.
   (c) Some non-metals burn in oxygen to form neutral oxides.
6. (a) Addition of oxygen to a substance is called oxidation.
   (b) Removal of oxygen from a substance is called reduction.
7. The ability of a metal to compete for combined oxygen depends on the position of that metal in the Reactivity Series.

Revision Exercise

1. Explain what happens when anhydrous calcium chloride or anhydrous copper (II) sulphate are exposed to the atmosphere for about two days.
2. Describe an experiment to show that there is increase in mass when magnesium is burned in air.
3. Is air a mixture or a compound? Explain.
4. Explain why cars in the coastal city of Mombasa rust faster than cars in Kisumu City.
5. State one advantage and three disadvantages of rusting.
6. List some industrial plants in Kenya and indicate the gaseous pollutants they emit.
7. Describe an experiment for the preparation and collection of oxygen from sodium peroxide.
8. The apparatus below were used to determine the volume of oxygen in air. About 200cm^3 of air were passed repeatedly from syringe A to syringe B over heated copper turnings as shown in the diagram.



After sometime, the volume of air in syringe A was 160 cm^3 and syringe B 0 cm^3.

(a) Calculate the percentage of oxygen in the initial sample of air.

(b) Write down a word equation for the reaction that took place in the combustion tube.

(c) What are the possible sources of error in the experiment?

CHAPTER FIVE
WATER AND HYDROGEN

Water is the most abundant substance on earth. About 71% of the earth's surface is covered by water. The sources of water are seas, lakes, rivers and oceans. Water has diverse uses and hence its study is important.

By the end of this chapter you should be able to:

- Describe experiments to show that water is a product of burning organic matter and that it contains hydrogen.
- State the products of reactions of metals with cold water and steam then derive the reactivity series of the metals.
- Prepare hydrogen and investigate its properties and uses.

Experiment 5.1: What products are formed when candle wax burns in air?

Set the apparatus as shown in figure 5.1. Light the candle and turn on the pump. Allow the candle to burn for about 15 minutes. Observe and record what happens in tube A and B.

Answer the following questions

1. What observations are made in:
   (a) test-tube A?
   (b) test-tube B? Explain.
2. Explain why test-tube A is dipped in cold water.

Divide the contents of test-tube A into two portions. To one portion, add anhydrous copper (II) sulphate or blue cobalt (II) chloride paper. Record your observations.

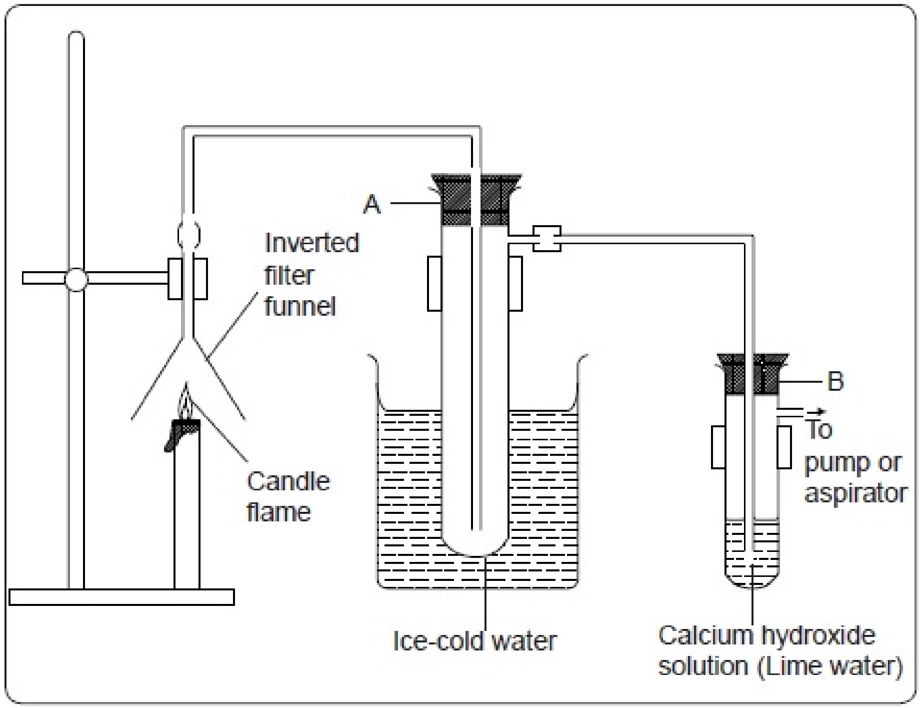


Fig 5.1: Products of burning a candle in air.

3. What is the effect of the contents of test-tube A on:
   (i) anhydrous copper (II) sulphate?
   (ii) anhydrous cobalt (II) chloride?
4. Identify the products formed when candle wax burns in air.
5. Which elements are present in candle wax?

Discussion

Candle wax burns in air to form a colourless liquid, which turns white anhydrous copper (II) sulphate to blue and blue cobalt (II) chloride paper to pink. The liquid formed is water. A colourless gas that forms a white precipitate with lime water is also produced. This gas is carbon (IV) oxide.

Candle wax is a compound of carbon and hydrogen. It belongs to a group of compounds called **hydrocarbons.** It burns in air to form carbon (IV) oxide and water. Other hydrocarbons include kerosene, petrol and diesel.

Carbon and hydrogen are the major components of organic matter. When organic matter burns in air carbon (IV) oxide and water are produced.

Reactions of Water with Metals

Experiment 5.2: What happens when some metals are in contact with cold water?

(a) Cut a very small piece of sodium metal, the size of a rice grain. Drop it into a trough containing cold water, as shown in figure 5.2. Test the solution in the trough using pieces of red and blue litmus papers. Record your observations.

CAUTION
The piece of sodium used in this experiment should be not more than the size of a rice grain!

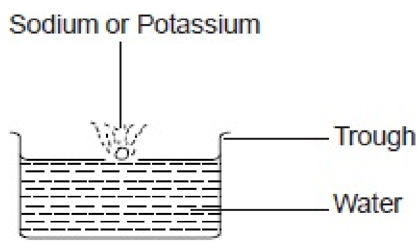


Fig 5.2: Reaction of sodium metal with cold water

(b) Drop a piece of calcium metal into water in a trough. Invert a funnel over the calcium. Fill a test-tube with water and invert it over the funnel. Collect the gas given off as shown in figure 5.3. Cork the test-tube to prevent the gas from escaping. Lift the test-tube from water and turn it upright. As you remove the cork, bring a burning splint near the mouth of the test-tube. Observe what happens. Test the solution in the beaker using pieces of red and blue litmus paper. Record your observations.

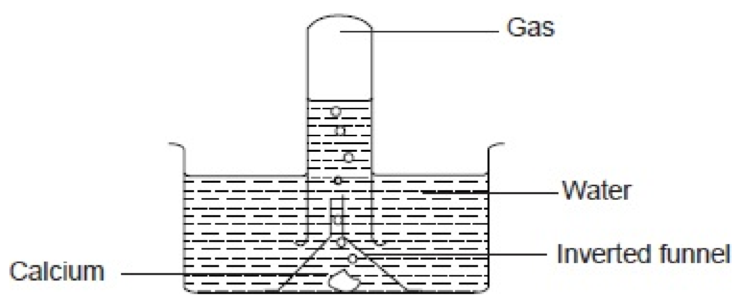


Fig 5.3: Reaction of calcium with cold water

(c) Put a two-centimetre piece of clean magnesium ribbon into a test-tube containing water. Record your observations. Repeat the experiment using clean zinc granules and fresh iron filings respectively.

Answer the following questions

1. What is observed when each of the metals reacts with water? Identify the gas liberated in each case.
2. What is observed when the solution is tested with litmus paper? Explain.
3. Write word equations for the reactions which occurred.
4. Why was it necessary to clean the magnesium ribbon?

NOTE
Sodium and potassium are stored under paraffin because they react with air and moisture.

Discussion

When a piece of sodium metal is placed in water, it melts into a silvery ball as it reacts vigorously darting on the surface of the water with a hissing sound. The resulting solution turns red litmus blue showing that it is basic.

$$\text{Sodium + Water} \longrightarrow \text{Sodium hydroxide + Hydrogen.}$$

When calcium is added to water it sinks and reacts moderately with water producing a steady stream of bubbles. When the gas is tested with a burning splint, it burns with a 'pop' sound indicating the gas is hydrogen. A basic solution of calcium hydroxide is formed. A white suspension is observed because the calcium hydroxide is slightly soluble in water. A water soluble base is called an **alkali.**

$$\text{Calcium + Water} \longrightarrow \text{Calcium hydroxide + Hydrogen}$$

Magnesium reacts with atmospheric oxygen to form a coating of magnesium oxide. The coating has to be removed so that the metal surface comes into contact with the water. Magnesium reacts very slowly with cold water while zinc and iron do not react.

The reaction of potassium with water is explosive. A small piece of potassium placed on water melts into a silvery ball and moves about rapidly on the surface. The reaction generates a lot of heat. As a result, hydrogen gas produced ignites spontaneously.

The flame produced is lilac (purple) due to the presence of potassium vapour produced during the reaction.

$$\text{Potassium + Water} \longrightarrow \text{Potassium hydroxide + Hydrogen}$$

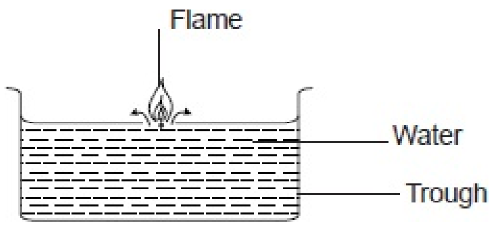


Fig 5.4 Reaction of potassium metal with cold water.

NOTE

- The magnesium ribbon is coiled in order to increase the surface area in contact with the boiling tube.
- The iron fillings must be fresh. Alternatively use steel wool.

Reactions of Metals with Steam

Experiment 5.3: What happens when steam is passed over heated metals?

Place clean, wet sand in a boiling tube. Clean a piece of magnesium ribbon about 6cm long and make it into a spiral coil. Place it in the middle of the boiling tube and arrange the apparatus as shown in figure 5.5. Heat the magnesium strongly and warm the wet sand gently as you continue heating the magnesium. Observe what happens. Remove the delivery tube before you stop heating. Test the gas produced using a burning splint. Record your observations. Repeat the experiment using zinc powder and iron filings.

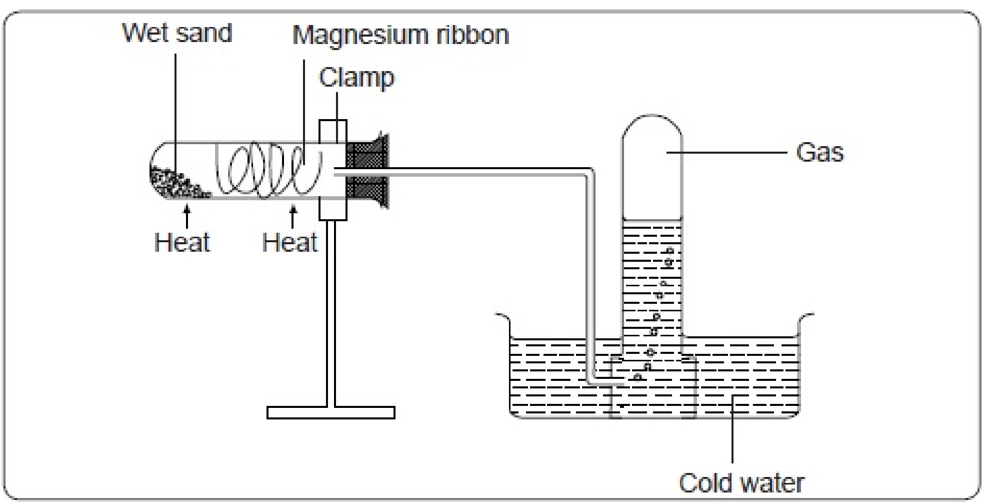


Fig 5.5. Effect of steam on metals.

Answer the following questions

1. Why is the wet sand warmed?

2. What is observed when steam is passed over heated magnesium, zinc or iron?
3. For each reaction;
   (a) name the gas involved.
   (b) the solid left in the test-tubes.
   (c) Write word equations for the reactions which occurred.
4. Why is the delivery tube removed before heating is stopped?
5. Arrange the metals in order of their reactivity with cold water and steam.

Discussion

Wet sand is warmed to generate steam, which reacts with the metal. Some metals, do not react with cold water but react with steam. Magnesium burns brightly in steam to form white magnesium oxide powder and hydrogen gas.

$$\text{Magnesium} + \text{Steam} \longrightarrow \text{Magnesium oxide} + \text{Hydrogen}$$

Zinc and iron do not burn in steam, they glow. Zinc metal reacts with steam to form a yellow powder of zinc oxide, which turns white on cooling.

$$\text{Zinc} + \text{Steam} \longrightarrow \text{Zinc oxide} + \text{Hydrogen gas}$$

Iron forms a black residue of an oxide called tri-iron tetra oxide and hydrogen.

$$\text{Iron} + \text{Steam} \longrightarrow \text{Tri-iron tetra oxide} + \text{Hydrogen gas}$$

Aluminium reacts with steam but quickly forms a layer of aluminium oxide which prevents further reaction. Lead and copper do not react with either cold water or steam. During this experiment the delivery tube is removed before heating stops to prevent water being sucked into the hot boiling tube as it cools. Table 5.1 summarises the reactions between metals and water or steam.

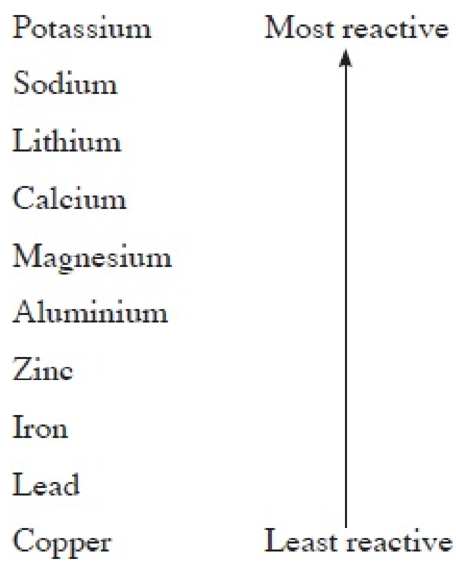
Table 5.1: Summary of reaction of metals with cold water and steam

Metal	*Action of metal on water*	*Action of metal on steam*
Potassium	Violent	Explosive
Sodium	Violent	Explosive
Calcium	Moderate	Violent
Magnesium	Very slow	Rapid
Aluminium	No reaction	Slow
Zinc	No reaction	Slow

Iron	No reaction	Slow
Lead	No reaction	No reaction
Copper	No reaction	No reaction

Metals can be arranged in order of their reactivity with water. The arrangement gives the **Reactivity Series of metals.**

Reactivity Series of Metals



NOTE

The sun is almost entirely made up of hydrogen.

Hydrogen

Hyrogen is the simplest and lightest element. It does not occur as a free element on earth but exists in combined form. Examples of compounds containing hydrogen are water, hydrocarbons and and sugar.

Laboratory Preparation of Hydrogen

Hydrogen gas is prepared in the laboratory by the reaction between dilute acids and some metals. Dilute sulphuric acid and hydrochloric acid react with some metals to liberate hydrogen gas. Zinc is the most suitable metal for the laboratory preparation of hydrogen gas.

Experiment 5.4: How is hydrogen prepared in the laboratory?

Put a few zinc granules into a conical flask and arrange the apparatus as shown in figure 5.6. Add dilute hydrochloric acid using the dropping funnel and collect several test-tubes of the gas produced. If the reaction is slow, add a few crystals of copper (II) sulphate. If the gas is required dry, it should be bubbled through concentrated sulphuric acid and collected by upward delivery as shown in figure 5.7.

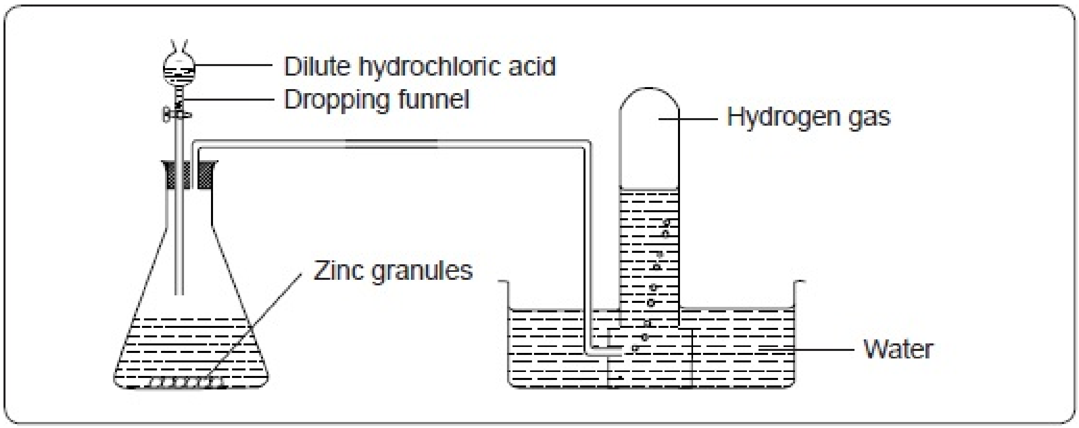


Fig 5.6: Preparation of hydrogen gas.

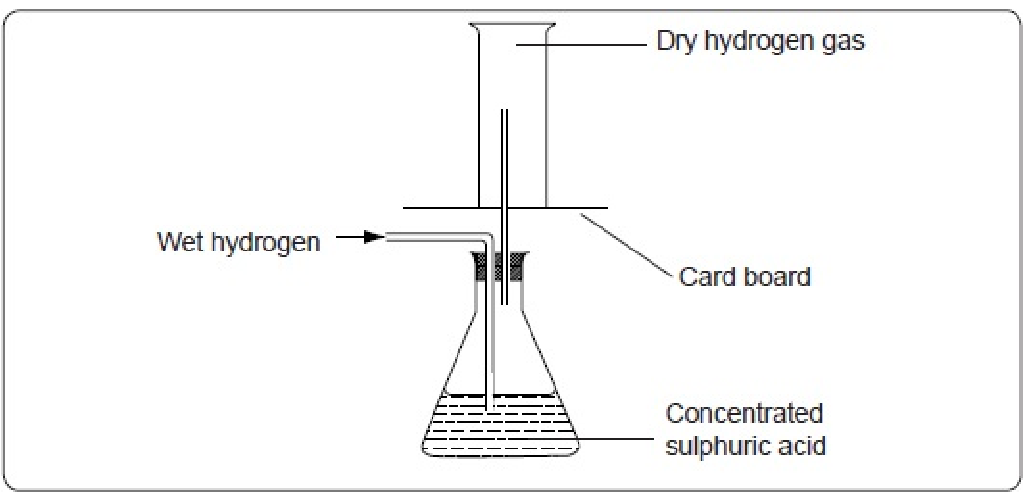


Fig 5.7: Collection of dry hydrogen.

Perform the following tests on the samples of the gas collected. Record your observations.

1. Note the colour and smell of the gas.
2. Test the gas using:
   (i) a lighted splint.
   (ii) wet blue and red litmus paper.

3. Invert an empty test-tube over the one containing the gas for about half a minute as shown in figure 5.8. Take a lighted splint to the mouth of the upper test-tube as you turn it upright.

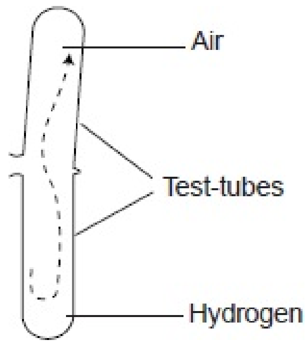


Fig 5.8: The density of hydrogen gas compared to that of air.

Answer the following questions

1. Why is it possible to collect the gas;
   (a) by downward displacement of air?
   (b) over water?
2. What is the;
   (a) colour of hydrogen gas?
   (b) smell of hydrogen gas?
3. What is the use of copper (II) sulphate in the reaction?
4. What is the test for hydrogen?
5. Write the word equation for the reaction generating hydrogen gas.

Discussion

When zinc granules are added to dilute hydrochloric acid there is effervescence and hydrogen gas is evolved. A small amount of copper (II) sulphate crystals may be added to speed up the reaction. Copper (II) sulphate acts as a catalyst.

$$\text{Zinc + Hydrochloric acid} \longrightarrow \text{Zinc chloride + Hydrogen}$$

Dry hydrogen gas can be obtained by passing wet hydrogen through anhydrous calcium chloride in a U-tube or passing it through concentrated sulphuric acid.

Nitric acid is not used to prepare hydrogen gas because the hydrogen formed is oxidised to water. However, very dilute nitric acid liberates hydrogen with magnesium metal.

Potassium, sodium, lithium and calcium react explosively with dilute acids hence **must** not be used. Magnesium could be used for the laboratory preparation of hydrogen but it is expensive. Aluminium forms a protective layer of aluminium oxide, which should be washed using concentrated hydrochloric acid before the metal can react with dilute acids. Zinc is preferably used because it reacts moderately with dilute acids. Impure iron gives a mixture of gases including the bad smelling hydrogen sulphide when it reacts with dilute acids.

Hydrogen is a colourless, odourless gas and is insoluble in water, so it is collected over water. Hydrogen is less dense than air and it can also be collected by **upward delivery.** This method is also called **downward displacement of air.**

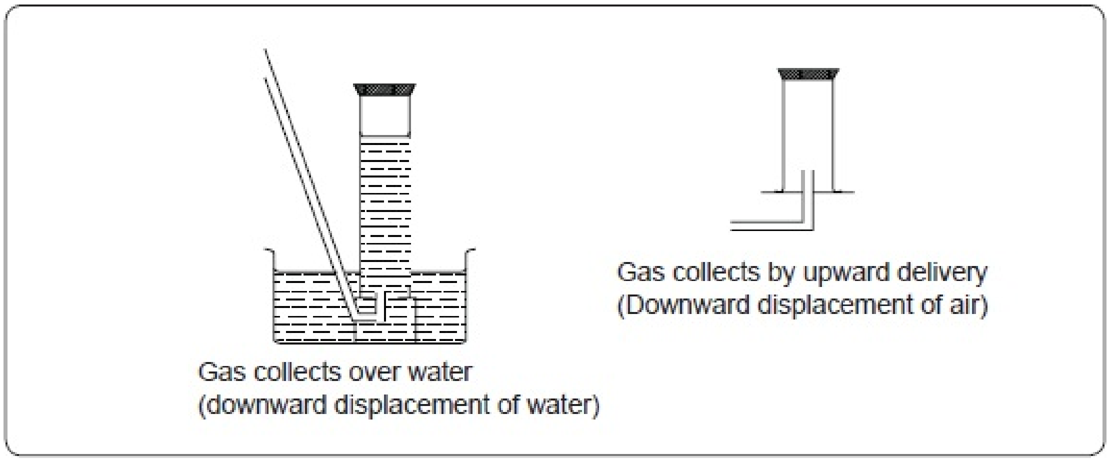


Fig. 5.9: Two methods of collecting hydrogen gas.

Hydrogen has no effect on litmus paper. It is neutral. When mixed with air, hydrogen burns with a 'pop' sound. This is the test for hydrogen gas. The intensity of the pop sound diminishes as the purity of hydrogen increases. Pure hydrogen burns quietly with a blue flame. Hydrogen does not support combustion but it burns in air.

Experiment 5.5: What happens when hydrogen is passed over heated copper (II) oxide?

Place a spatulaful of copper (II) oxide in a porcelain boat or on a piece of aluminium foil. Place the boat or aluminium foil in a combustion tube. Set the apparatus as shown in figure 5.9. Pass dry hydrogen through the tube for some time to drive out the air from the apparatus. Collect a sample of the gas from the jet in a test-tube and test it with a burning splint. Continue collecting and testing until a sample of the gas burns silently (without a 'pop' sound).

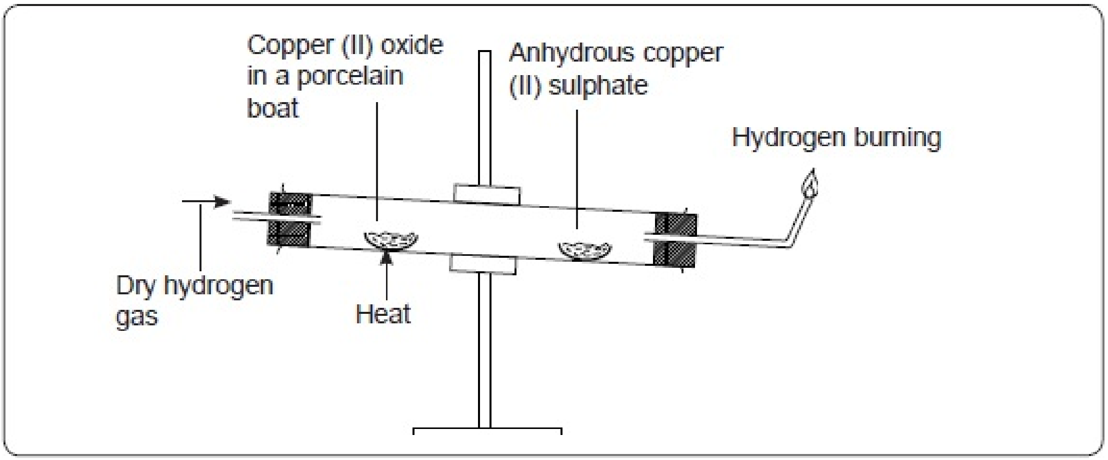


Fig. 5.10: Action of hydrogen gas on heated copper oxide.

Light the gas at the jet and start heating copper (II) oxide as shown in figure 5.9. Ensure a steady flow of hydrogen gas. Continue heating until there is no further change. Observe what happens. Allow the apparatus to cool as you continue passing the stream of hydrogen gas over the residue. Disconnect the apparatus and test the colourless liquid in the cooler parts of the combustion tube.

Answer the following questions

1. Why is it important to drive out all the air from the combustion tube before lighting the jet?
2. State and explain the observations made in the combustion tube.
3. Why is excess hydrogen burned and not allowed to escape into the air?
4. Why is the supply of hydrogen gas continued while the apparatus cools?
5. Write a word equation for the reaction between hydrogen and copper (II) oxide.
6. Define the terms reduction, oxidation, reducing agent and oxidising agent.

Discussion

Air is driven out of the apparatus to ensure the hydrogen being burned at the jet is pure to avoid explosion. On passing a stream of dry hydrogen gas over hot copper (II) oxide the black copper (II) oxide changes to red-brown. At the same time a colourless liquid condenses and collects on the cooler parts of the combustion tube.

Hydrogen combines with oxygen from the copper (II) oxide to form water which is the colourless liquid. The red-brown solid is copper metal. *Removal of oxygen from a compound is known as reduction while the addition of oxygen is known as oxidation*. In this reaction copper (II) oxide is reduced to copper metal while hydrogen is oxidised to water.

$$\text{Copper (II) oxide} + \text{Hydrogen} \longrightarrow \text{Copper metal} + \text{Water}$$

Hydrogen removes oxygen from copper (II) oxide and is therefore referred to as a reducing agent. Copper (II) oxide loses oxygen to hydrogen hence it is an oxidising agent. Hydrogen will also reduce the oxides of lead and iron.

$$\text{Lead (II) oxide} + \text{Hydrogen} \longrightarrow \text{Lead metal} + \text{Water}$$

$$\text{Iron (III) oxide} + \text{Hydrogen} \longrightarrow \text{Iron metal} + \text{Water}$$

Hydrogen does not remove oxygen from oxides of more reactive metals.

The supply of hydrogen gas is continued while the apparatus cools to avoid the reoxidation of the hot metal by oxygen from the air. Excess hydrogen gas is burned because its mixture with oxygen (air) is explosive when ignited. The excess gas is therefore not allowed to escape into the air for safety reasons.

Experiment 5.6: What is formed when hydrogen burns in air?

Arrange the apparatus as shown in figure 5.10. Pass a stream of hydrogen gas through anhydrous calcium chloride. Test the gas for purity by collecting samples over the jet and testing with a burning splint until the gas burns silently. Light the gas at the jet and turn on the pump. Draw in the products of the burning hydrogen using the pump for about 15 minutes. Test the liquid in the test-tube using white anhydrous copper (II) sulphate.

Record your observations

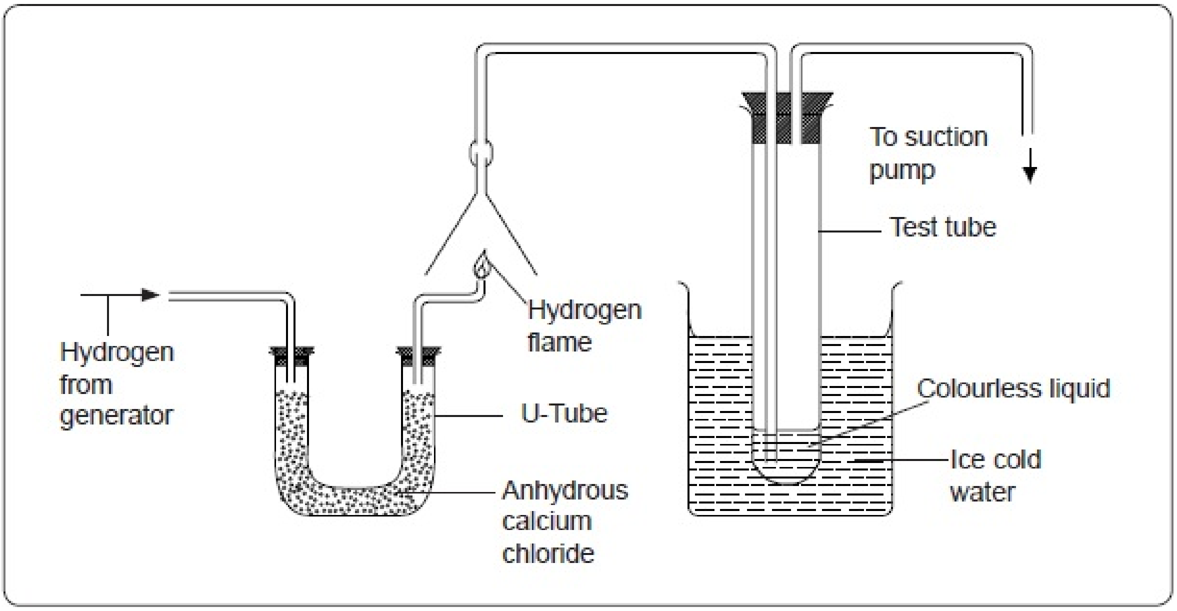


Fig 5.10: Products of burning hydrogen in air.

Answer the following questions

1. (a) What is the effect of the liquid collected in the test tube on anhydrous copper (II) sulphate? Explain.
2. What is the use of anhydrous calcium chloride in this experiment?
3. What is the use of the ice-cold water in the beaker?
4. Write a word equation for the burning of hydrogen in air.

Discussion

When hydrogen burns in air a colourless liquid which turns white anhydrous copper (II) sulphate blue is formed. The liquid is water. Hydrogen combines with oxygen from the air to form water. Water is therefore an oxide of hydrogen.

$$\text{Hydrogen} + \text{Oxygen} \longrightarrow \text{Hydrogen oxide (water)}$$

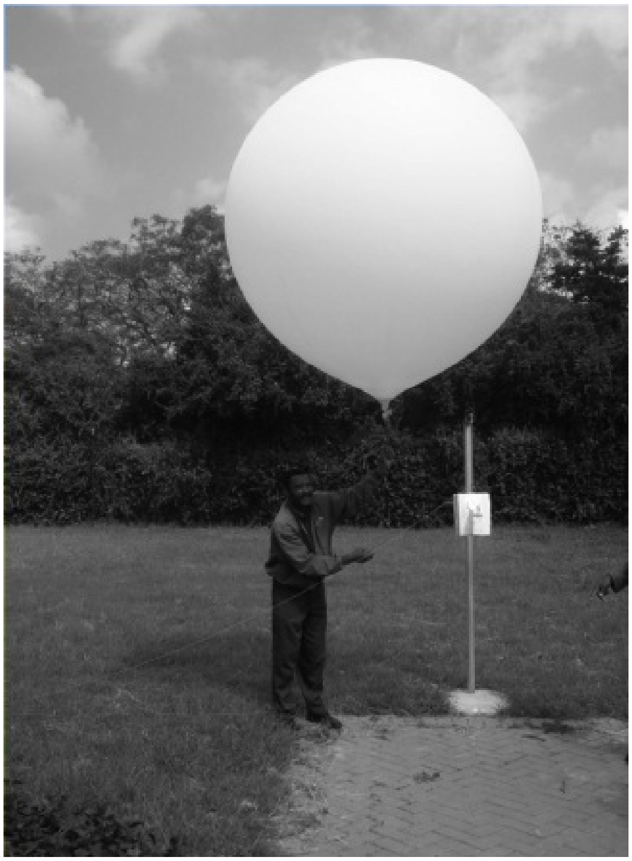
Anhydrous calcium chloride is used to dry the gas. The ice cold water condenses the steam to form liquid water.

Uses of Hydrogen

1. Hydrogen is used in the large-scale manufacture of ammonia in a process known as the **Haber process.**
2. The gas is also used during the hardening of oils to form fats. When hydrogen is passed through liquid oil in the presence of nickel catalyst, the

oil takes up hydrogen and is converted into fat. This process is called **hydrogenation.** It is used in the manufacture of margarine.

3. Hydrogen is used in balloons because it is lighter than air. A balloon filled with hydrogen floats in air. A light radio instrument can be connected to the balloon to collect information from the atmosphere by meteorologists who study weather conditions.



A technician prepares to launch a radio sonde.

4. A mixture of hydrogen and oxygen burns to produce a very hot flame, the oxy-hydrogen flame, which has a temperature of up to 2000°C. The flame is used in welding and for cutting metals.
5. Hydrogen is used in rocket as fuel.
6. Hydrogen is used in the manufacture of hydrochloric acid.
7. Hydrogen is used as a fuel in fuel cells.

Summary

1. Water is the most abundant substance on the surface of the earth. About 71% of the earth's surface is covered by water.
2. Hydrogen is prepared in the laboratory by the action of zinc on dilute hydrochloric acid or dilute sulphuric acid.
3. Hydrogen burns in oxygen to produce water. Water is an oxide of hydrogen.
4. Active metals react with cold water to produce hydrogen gas and the hydroxide of the metal in solution. Less active metals react with steam to produce hydrogen and the oxide of the metal. Copper and lead do not react with water.
5. Hydrogen is a reducing agent. It removes combined oxygen from metal oxides of the less reactive metals.
6. Reduction is the loss of oxygen from a compound. Oxidation is the gain of oxygen by a substance.
7. A reducing agent is a substance which removes oxygen from another substance. An oxidising agent is a substance which gives out oxygen to another substance.

Revision Exercise

1. (a) State the chemical tests for presence of water.
   (b) State the test, which is used to show that water is pure.
2. Describe an experiment to show that water is an oxide of hydrogen.
3. State the precautions that must be taken when carrying out experiments with hydrogen.
4. Why is it not advisable to use iron in making steam boilers?
5. Write a word equation for a reaction in which hydrogen acts as a reducing agent.
6. Name the products formed when kerosene is burned in air.
7. State what is observed when a small piece of potassium is placed in water. Write a word equation for the reaction.
8. Draw a labelled diagram to show how a reaction between steam and magnesium should be carried out.
9. Describe how dry hydrogen is prepared in the laboratory.

This Fifth Edition of ***Secondary Chemistry Form One Students' Book*** published by Kenya Literature Bureau is the first of the Chemistry series to be completely revised, repackaged and reissued in the new larger format. It covers all the topics in Form One as per syllabus requirements.

It incorporates the following features which learners will find useful:

- An attractive layout which greatly enhances readability.
- Employs a simple language with clear introduction to the topics followed by easy to understand explanations and precise instructions for practical activities.
- A well applied two colour scheme with clear diagrams and high quality photographs to aid comprehension.
- Well chosen end of topic revision questions and topic summaries to assist recall and self evaluation.
- The book has retained the popular enquiry approach to presentation of content.
- Use of local examples and infusion of emerging issues to enliven the learning of Chemistry.
- A totally new section of examination style questions and answers.

The book is accompanied by a Teachers' Guide, which gives pointers on time allocation, practicals, problematic areas and answers to revision exercises.

Cover Photograph: *Students assist to assemble fractional distillation apparatus.* ***Courtesy of Nembu Secondary School.***

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