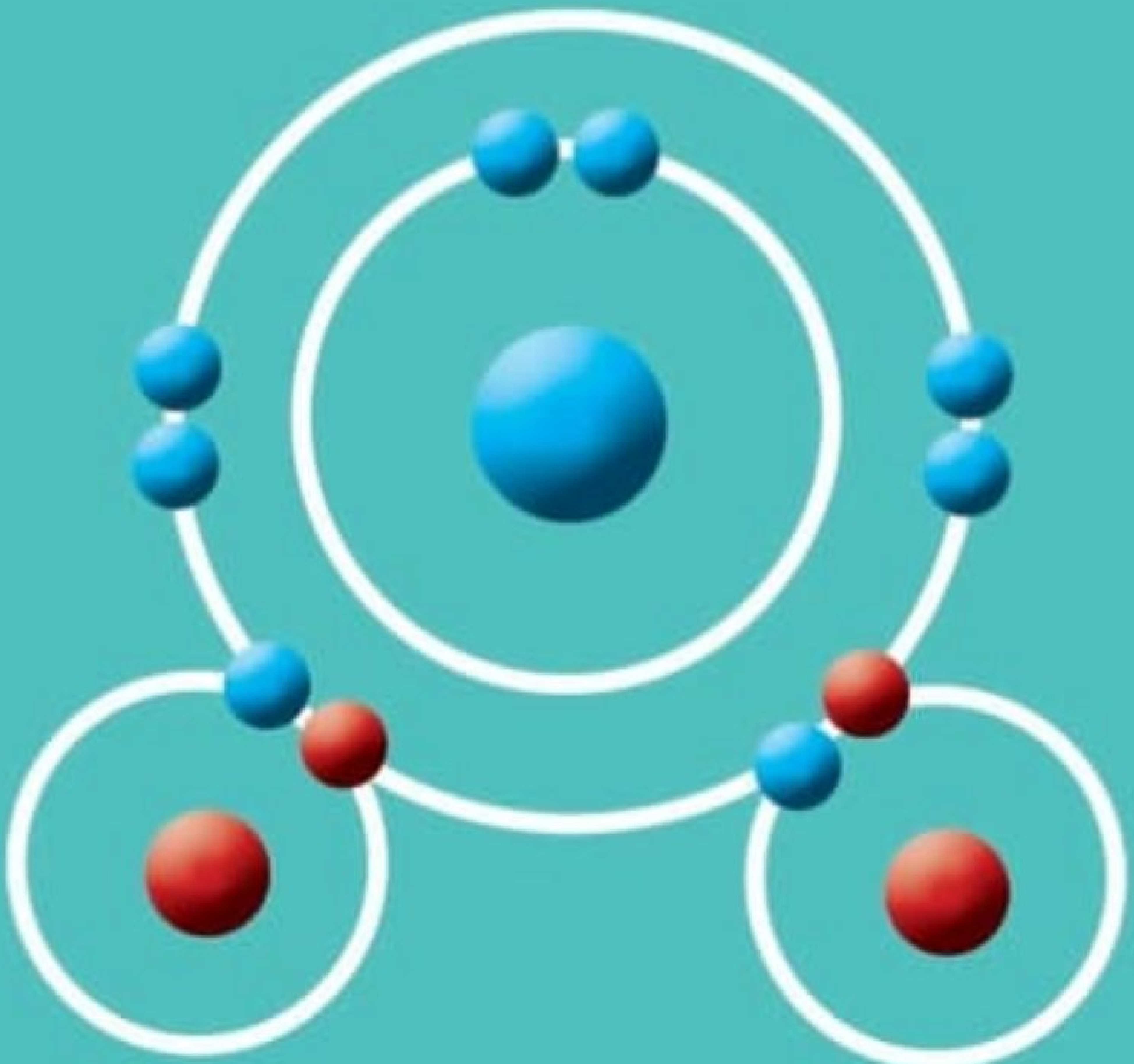
Chemistry
for Secondary Schools
Student's Book
Form Two



Tanzania Institute of Education

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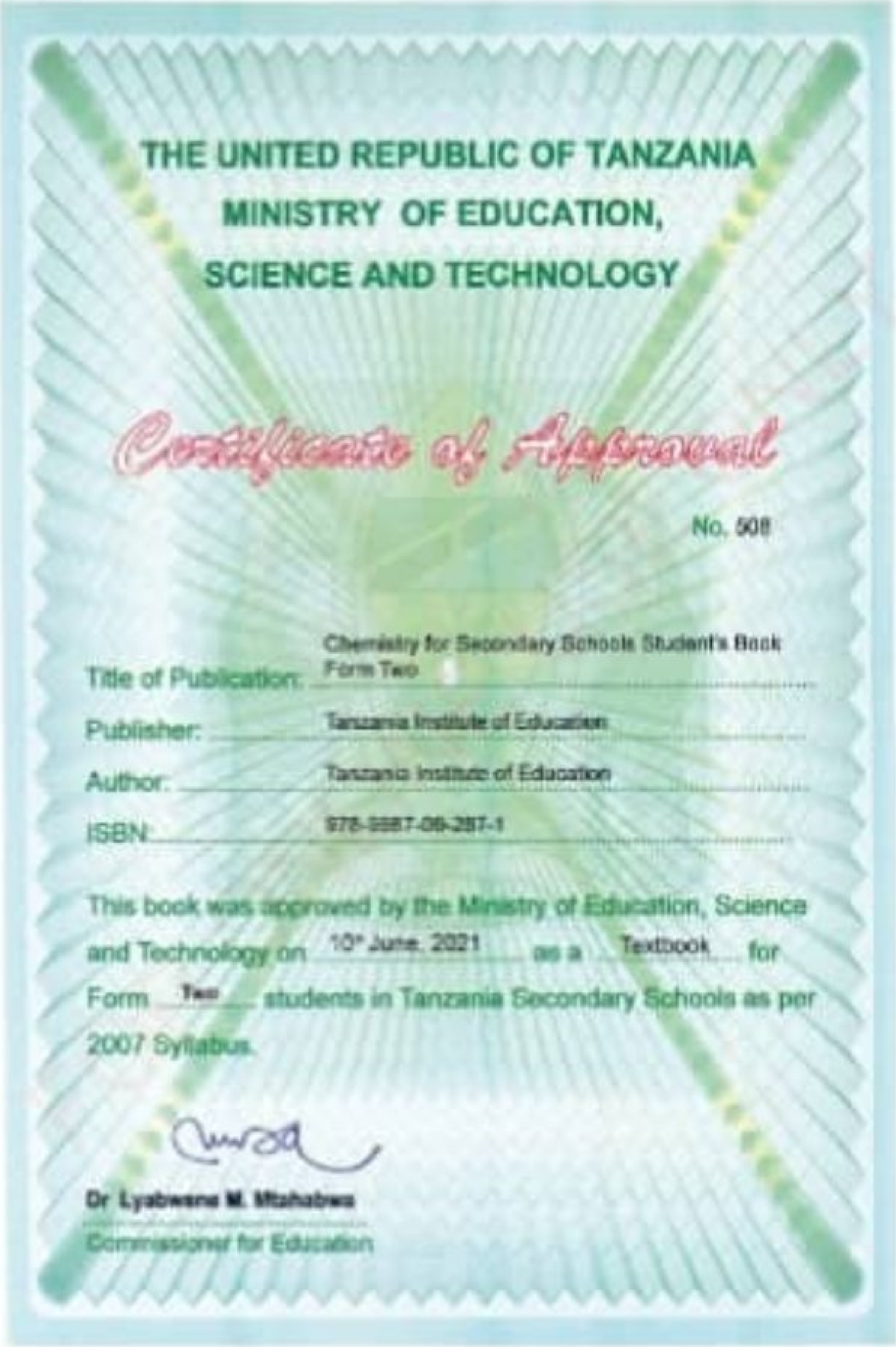

THE UNITED REPUBLIC OF TANZANIA
MINISTRY OF EDUCATION,
SCIENCE AND TECHNOLOGY

Certificate of Approval

No. 508

Title of Publication: Chemistry for Secondary Schools Student's Book Form Two

Publisher: Tanzania Institute of Education

Author: Tanzania Institute of Education

ISBN: 978-9987-09-287-1

This book was approved by the Ministry of Education, Science and Technology on 10th June, 2021 as a Textbook for Form Two students in Tanzania Secondary Schools as per 2007 Syllabus.

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Commissioner for Education

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Published 2021

ISBN: 978–9987–09–287–1

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Acknowledgements

The Tanzania Institute of Education (TIE) would like to acknowledge the contributions of all the organisations and individuals who participated in designing and developing this textbook. In particular, TIE wishes to thank the University of Dar es Salaam (UDSM), Dar es Salaam University College of Education (DUCE), Mkwawa University College of Education (MUCE), Marian University College (MARUCo), the Open University of Tanzania (OUT), the State University of Zanzibar (SUZA), school quality assurance offices, teacher colleges and secondary schools.

Besides, the following individuals are also acknowledged:

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Furthermore, TIE extends its sincere appreciation to the United States Agency for International Development (USAID)-Tanzania, for granting permission to use materials from the 2008 Chemistry for Secondary Schools, Forms 1 & 2 (First Edition) textbook.

TIE also appreciates the contribution of secondary school teachers and students who participated in the trial phase of the manuscript.

Likewise, the Institute would like to thank the Ministry of Education, Science and Technology for facilitating the writing and printing of this textbook.

Dkt. Aneth A. Komba
Director General
Tanzania Institute of Education

Preface

This textbook, *Chemistry for Secondary Schools*, is written specifically for Form Two students in the United Republic of Tanzania. It is prepared in accordance with the 2007 Chemistry Syllabus for Secondary Schools, Form I–IV, issued by the then, Ministry of Education and Vocational Training (MoEVT).

The book consists of seven chapters, namely Oxygen, Hydrogen, Water, Fuels and energy, Atomic structure, Periodic classification, and Chemical bonding, formula and nomenclature. The chapters comprise of illustrations, activities, tasks, projects, and exercises. You are encouraged to do all the activities, projects, and exercises together with any other assignment provided. Doing so, will promote the development of the intended competencies.

Tanzania Institute of Education

Chapter One

Oxygen

Introduction

Oxygen is one of the gases that constitute air. In this chapter, you will learn about the occurrence, laboratory preparation, properties, industrial production and uses of oxygen. The competencies developed will enable you to prepare oxygen, perform simple experiments to demonstrate the properties of oxygen gas, and explain its uses. You will also be able to handle different chemical reactions involving oxygen and help your society in many issues related to it.

Occurrence of oxygen

Oxygen occurs as a gas with an abundance of about 21% by volume of the air. It also occurs in combination with other substances, for example, in water, mineral ores, and other chemical compounds such as protein molecules that make up most of the living things.

Laboratory preparation of oxygen

The preparation of oxygen in the laboratory can be done through thermal decomposition of potassium chlorate, decomposition of hydrogen peroxide and heating some other compounds which are rich in oxygen.

Decomposition of potassium chlorate

Potassium chlorate is decomposed by heating in the presence of manganese(IV) oxide catalyst to produce potassium chloride and oxygen gas. Figure 1.1 shows a schematic diagram for the preparation of oxygen gas by thermal decomposition of potassium chlorate salt.

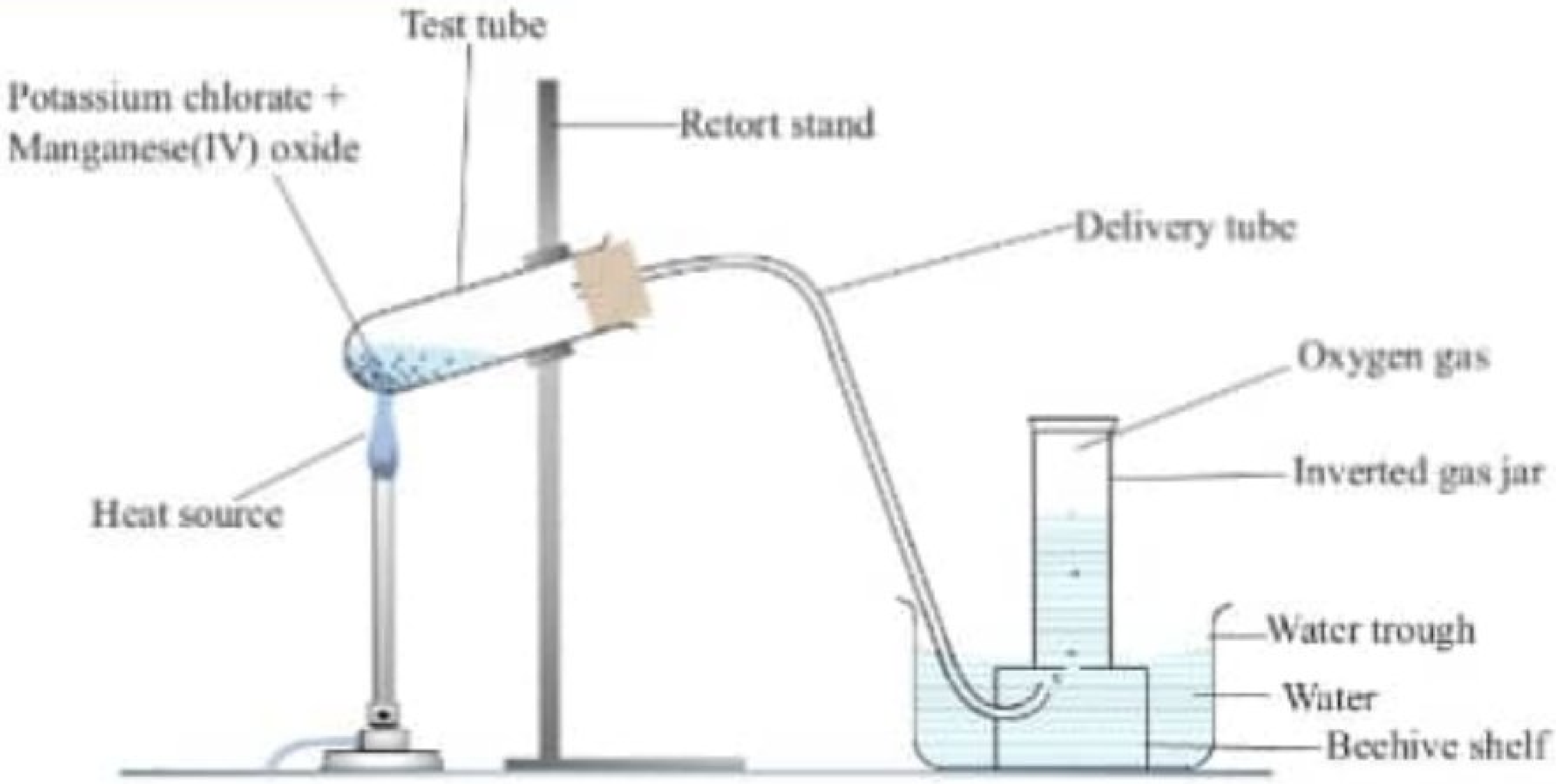


Figure 1.1: *Experimental set-up for the preparation of oxygen by thermal decomposition of potassium chlorate*

The decomposition of potassium chlorate is summarised in the following word and formula equations:

$$\text{Potassium chlorate} \xrightarrow[\textit{Heat}]{\textit{Manganese(IV) oxide}} \text{Potassium chloride} + \text{Oxygen gas}$$

$$2KClO_3(s) \xrightarrow[Heat]{MnO_2} 2KCl(s) + 3O_2(g)$$

Decomposition of hydrogen peroxide

Compared to potassium chlorate, this method is preferred because it produces enough gas without the use of heat. The decomposition of hydrogen peroxide in the presence of manganese(IV) oxide yields oxygen and water as expressed by the following equations:

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

$$2H_2O_2(l) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

Decomposition is the reaction in which a chemical compound breaks down into its constituent elements or simpler compounds. A *catalyst* is a substance that alters the rate of a chemical reaction but remains unchanged at the end of the reaction. Manganese(IV) oxide speeds up the decomposition of potassium chlorate and hydrogen peroxide.

Oxygen gas is collected over water by a process called *downward displacement of water*. Oxygen is slightly soluble in water and lighter than water, thus, during its preparation, it easily displaces water and is collected over water.

Activity 1.1

Aim: To prepare oxygen gas by decomposing hydrogen peroxide.

Requirements: Flat bottomed flask, beehive shelf, delivery tube, water trough, thistle funnel with a tap, gas jars, two-holed rubber bung, hydrogen peroxide, manganese(IV) oxide, and water

Procedure

1. Put about 5 g of manganese(IV) oxide into a flat-bottomed flask.
2. Set the apparatus as shown in Figure 1.2. Make sure that the tap of the thistle funnel is closed.
3. Fill the trough with water to about ¾ full. Put a beehive shelf in it. Fill a gas jar with water and invert it over a beehive shelf.
4. Connect the delivery tube through the shelf in the trough. Ensure that the water in the gas jar has no bubbles.
5. Put 80 cm^3 of hydrogen peroxide in the thistle funnel.
6. Open the tap of the thistle funnel to allow hydrogen peroxide to fall onto manganese(IV) oxide. Ensure that hydrogen peroxide falls drop by drop.

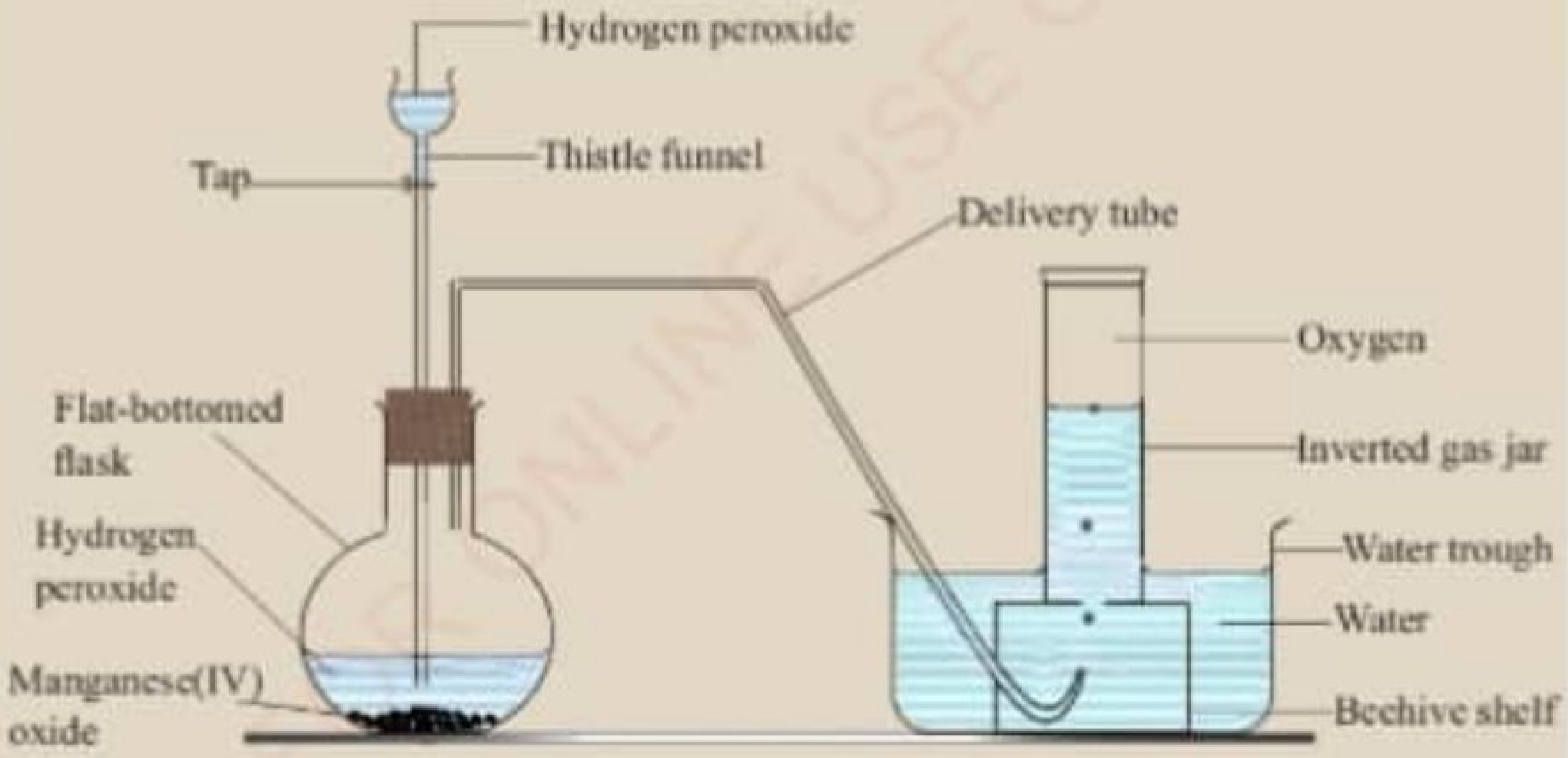


Figure 1.2: *Experimental set-up for the preparation of oxygen by using hydrogen peroxide*

7. Allow the first few bubbles to escape (this ensures purity of the collected gas), then collect the gas into the inverted gas jar. Remove the gas jar from the beehive shelf keeping it well covered with a lid. Collect several gas jars of the gas for further use.

Questions

1. Both manganese(IV) oxide and hydrogen peroxide contain oxygen. Which of them produced the oxygen you collected? Give reason(s).
2. How is it possible that oxygen can be collected over water?
3. Why did you allow the first few bubbles of the gas to escape?

The reaction of other compounds containing oxygen

Other compounds that contain oxygen can be used to prepare oxygen gas, for example, potassium permanganate. However, this method is not commonly used in the laboratory preparation of oxygen due to some associated challenges such as high energy requirement and explosion. For example, potassium permanganate requires high temperature (200-300 °C) to decompose.

Mercury oxide can also be heated to produce oxygen gas. However, the process produces a mixture of highly toxic mercury fumes and oxygen. The fumes irritate the eyes, skin and respiratory tract. These fumes may also have effects on the kidneys and can cause death.

Exercise 1.1

1. Why is hydrogen peroxide preferred over potassium chlorate in the laboratory preparation of oxygen?
2. Why is oxygen gas collected by downward displacement of water?
3. Oxygen gas may be prepared in the laboratory by heating a mixture of potassium chlorate with manganese(IV) oxide. Which of these two compounds produces the required oxygen?

4. During the laboratory preparation of oxygen, manganese(IV) oxide is used as a catalyst in the decomposition reactions.
    (a) What is decomposition?
    (b) What is a catalyst?
    (c) What would happen if the preparations of oxygen were performed without the use of catalysts?

Properties of oxygen

The physical and chemical properties of oxygen gas are as follows:

Physical properties

1. Oxygen gas is colourless.
2. It is odourless.
3. It is tasteless.
4. It is slightly soluble in water.
5. It is slightly denser than air (about 1.1 times).
6. It boils at –183 °C.
7. It freezes at –218 °C.

Chemical properties

1. Oxygen gas supports combustion.
2. It is a strong oxidising agent.
3. It reacts with metals to form basic oxides.
4. It reacts with non–metals to form acidic oxides.

Test for the presence of oxygen

When a glowing wooden splint is lowered into a gas jar containing oxygen gas, it re-lights up (Figure 1.3). This is the test for the presence of oxygen gas. Similarly, when a lit candle is lowered into a gas jar containing oxgen gas, it burns more brightly. This confirms that oxygen supports combustion.

A glowing splint in a gas jar without oxygen

A re-lit splint in a gas jar with oxygen

Figure 1.3: *A glowing wooden splint in a gas jar without oxygen and a re-lit splint in a gas jar with oxygen*

Activity 1.2

Aim: To demonstrate the chemical properties of oxygen gas.

Requirements: Gas jars with oxygen, a candle, wooden splints, metals (sodium, potassium and magnesium ribbon), non-metals (sulphur, carbon and phosphorus), deflagrating spoon and Bunsen burner

Procedure

1. Prepare and collect oxygen gas in several gas jars.
2. Light a wooden splint, let it burn for some time then extinguish the flame to leave a glowing end. Lower the glowing splint into a gas jar containing oxygen. Record your observation.
3. Place a small candle on a deflagrating spoon, light the candle and lower it in a gas jar containing oxygen. Note your observation. Figure 1.4 shows a deflagrating spoon with a burning candle immersed in the gas jar.

Figure 1.4: *Burning candle in a gas jar containing oxygen*

4. Take each of the metals and non-metals, one at a time and place each of them on a deflagrating spoon.
5. Heat each element using a Bunsen burner flame until it is red hot or it catches fire.
6. Put the hot or burning element into a jar containing oxygen. Leave it there until the burning stops.
7. Record your observations in a table, showing how each element burns, the colour, and the name of the products formed.

Questions

1. What happens when:
    (a) a glowing splint is lowered into a gas jar containing oxygen?
    (b) a lit candle is lowered into a gas jar containing oxygen?
2. (a) When the metals and non-metals were heated in oxygen, which ones burnt:
        (i) vigorously?
        (ii) slowly?
    (b) Identify the products formed as a result of burning the metals and non-metals.

Basic nature of metal oxides

Many metals burn in oxygen to produce basic oxides. These oxides are basic because they react with water to form basic solutions, or with acids to form salt and water. Basic solutions turn red litmus paper blue, for example, magnesium burns to form magnesium oxide whose solution turns red litmus paper blue.

Activity 1.3

Aim: To demonstrate the basic nature of the products of burnt metals.

Requirements: Gas jars containing oxygen, metals (potassium, sodium, calcium, magnesium, aluminium, zinc, iron, lead, copper), deflagrating spoon, red litmus paper, and water

Procedure

1. Pour some water into the gas jar containing oxygen.
2. Put a small piece of potassium metal on a deflagrating spoon, burn, and then lower it into the gas jar containing oxygen and water. Make sure that the deflagrating spoon is just above the water level. Let the metal continue burning in the gas jar until the powder is formed and the fire goes off.
3. Shake the powder in order to mix it with water, then dip a red litmus paper into the solution.
4. Follow the same procedure (steps 1 to 3), now using magnesium ribbon or any other metal.
5. Tabulate your observations in steps 3 and 4.

Figure 1.5 shows the flame colours of some metals burning in oxygen.

Caution: Sodium reacts vigorously with oxygen and moisture present in the air, and thus catches fire.

Potassium *Sodium* *Calcium* *Magnesium*

Figure 1.5: *Flame colours of some metals burning in oxygen*

Table 1.1 summarises the reactions and properties of products formed when metals burn in oxygen.

Table 1.1: *Properties of products formed when some metals burn in oxygen*

Metal	Colour of the metal	How it burns	Colour and nature of product	Name of product	Action of their aqueous solutions on red litmus paper
Potassium	Silvery white	Melts easily and burns with a lilac flame	White powder	Potassium oxide	Turns blue
Sodium	Silvery white	Burns vigorously with a yellow flame	Pale yellow solid	Sodium oxide	Turns blue
Calcium	Silvery white	Burns with a brick red flame	White solid	Calcium oxide	Turns blue
Magnesium	Grey/ Silvery	Melts and burns with a bright white flame	White powder	Magnesium oxide	Turns blue
Zinc	Bluish grey	Burns slowly with a dull red flame	Yellow/green flakes that are white when cool	Zinc oxide	No action
Iron	Silvery grey	Glows red hot	Reddish brown	Iron(III) oxide	No action
Copper	Orange red	Turns orange then the surface of the product turns black	Black solid	Copper(II) oxide	No action

Acidic nature of non-metal oxides

Properties of the products of burnt non-metals are different from those of metals. Non-metals burn in oxygen to produce acidic oxides. These oxides are acidic because they react with water to form acidic solutions which turn a blue litmus paper red.

Activity 1.4

Aim: To demonstrate the acidic nature of the products of burnt non-metals.

Requirements: Gas jars containing oxygen, non-metals (carbon, phosphorus, silicon, sulphur), deflagrating spoon, and blue litmus paper

Procedure

1. Pour some water into a gas jar containing oxygen gas.
2. Put a piece of carbon on a deflagrating spoon, burn, and then lower it into a gas jar containing oxygen and water. Let it continue burning in the gas jar until a powder forms and the fire goes off.
3. Shake the powder in order to mix it with water, then dip a blue litmus paper into the water.
4. Repeat steps 1 to 3, now using phosphorus and other non-metal elements, each at a time.
5. Tabulate your observations in steps 3 and 4.

Table 1.2 summarises the properties of the products of the reactions of non-metals with oxygen. Flame colours for some non-metals are shown in Figure 1.6.

Table 1.2: *Properties of products of the reactions of non-metals with oxygen*

Non-metal	Nature of the element	How it burns	Colour and nature of product	Name of product	Action on wet blue litmus paper
Carbon	Black solid	Burns slowly with a yellow-white flame	Colourless gas	Carbon dioxide	Turns red
Phosphorus	Yellow solid	Burns brightly to produce clouds of white smoke	White solid	Phosphorus(V) oxide	Turns red
Sulphur	Yellow solid	Melts and burns with a blue flame	Misty (white gas)	Sulphur dioxide	Turns red
Silicon	Dark-grey solid	Burns with dark-brown crystals	Solid whitish yellow	Silicon dioxide	Turns red

Carbon Phosphorus Sulphur Silicon

Figure 1.6: *Flame colours for some non-metals burning in oxygen*

Industrial production of oxygen

Various methods are used for large scale production of oxygen. The main methods are *fractional distillation of liquefied air* and *the electrolysis of water*. Currently, the most common method is the fractional distillation of liquefied air. This produces the highest amount of oxygen than the electrolysis of water.

Fractional distillation of air

Fractional distillation of air starts with the liquefaction of air followed by distillation of the liquid air.

Liquefaction of air

Liquefaction of air involves filtration of air to remove dust, compressing, and then cooling down to –200 °C until it liquefies. During the liquefaction, the following occur:

(a) Water vapour condenses, and is removed by using special filters.

(b) At –78.5 °C, carbon dioxide freezes and is removed.

(c) At –183 °C, oxygen liquefies.

(d) Nitrogen liquefies at –196 °C.

At –200 °C, there is still a mixture of some liquid nitrogen and liquid oxygen. The two liquids are separated by fractional distillation due to their close boiling points.

Distillation

The liquid mixture of nitrogen and oxygen is then separated by fractional distillation into pure oxygen and nitrogen gas. The liquid mixture is passed into a fractionating column from the bottom. Since the column is warmer at the bottom than at the top, the liquid nitrogen boils at the bottom of the column. The gaseous nitrogen rises to the top where it is collected via the pipe to the storage tank, while the liquid oxygen collects at the bottom of the column (Figure 1.7).

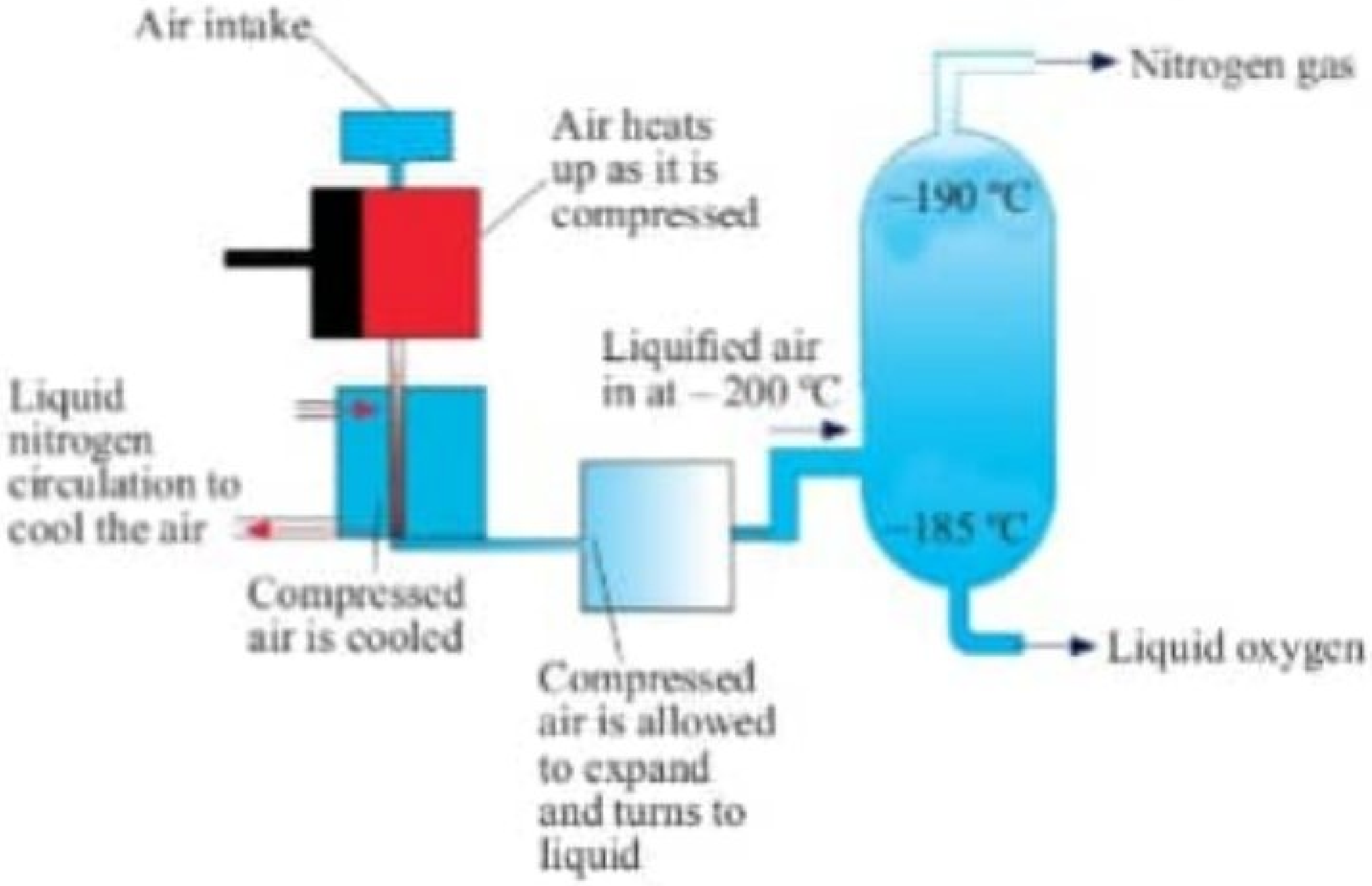


Figure 1.7: *Production of oxygen by fractional distillation of liquefied air*

Fractional distillation is the separation of liquid components with close boiling points from a liquid mixture. Figure 1.8 shows a fractional distillation plant.

Figure 1.8: *Fractional distillation plant*

Uses of oxygen

Life on earth depends on the presence of oxygen gas. You breath-in oxygen without which there would be no survival. Animals and plants need oxygen gas for respiration. Some important industrial processes also need oxygen. There are many uses of oxygen in the universe. The major uses of oxygen are as described in the following paragraphs.

Uses of oxygen in living organisms

Oxygen is used for respiration by organisms. Germinating seeds require oxygen and water, divers in the deep sea carry with them oxygen in cylinders to support them in breathing. Mountain climbers also carry oxygen in the cylinders to help them in breathing at high altitudes where there is low supply of oxygen in the air. In hospitals, oxygen is used to support the breathing of patients with breathing difficulties, and premature babies. Figure 1.9 shows a diver and a mountain climber with oxygen cylinders to support breathing.

A water diver *A mountain climber*

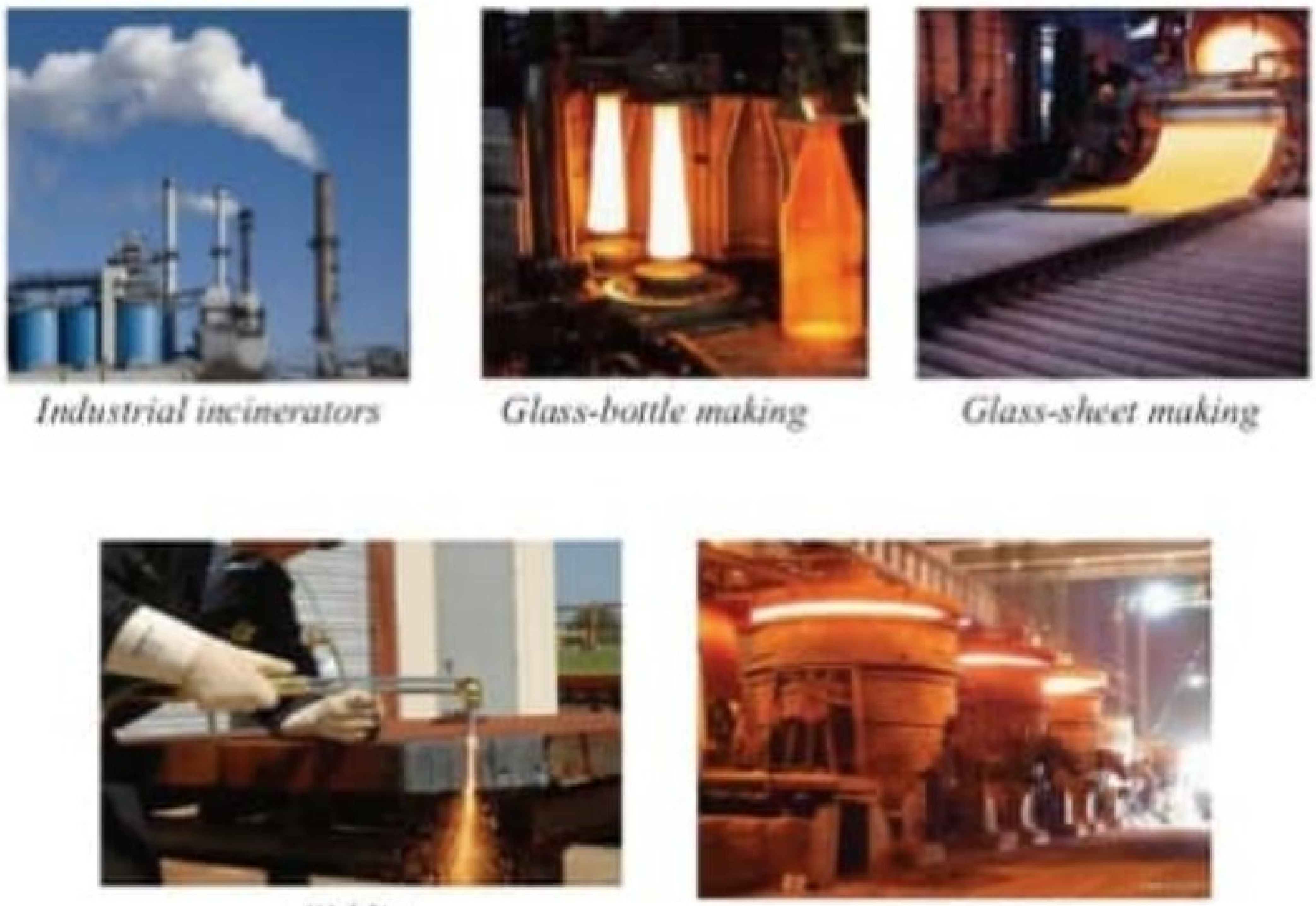
Figure 1.9: *Diver and a mountain climber with oxygen cylinders*

Environment

In environmental conservation, oxygen is used in sewage treatment and replenishment of the ozone layer that protects the Earth from harmful radiations. Oxygen is also useful in destruction of wastes, since it supports burning.

Industrial chemical processes

Oxygen is used in various chemical processes such as manufacturing of chemicals, incineration, manufacturing of synthetic fuels, manufacturing of steel, metal cutting, welding, glass making, pulp, and paper making. Some of these are shown in Figure 1.10.



Industrial incinerators *Glass-bottle making* *Glass-sheet making*

Welding *Steel manufacturing industry*

Figure 1.10: *Some of the chemical processes that use oxygen*

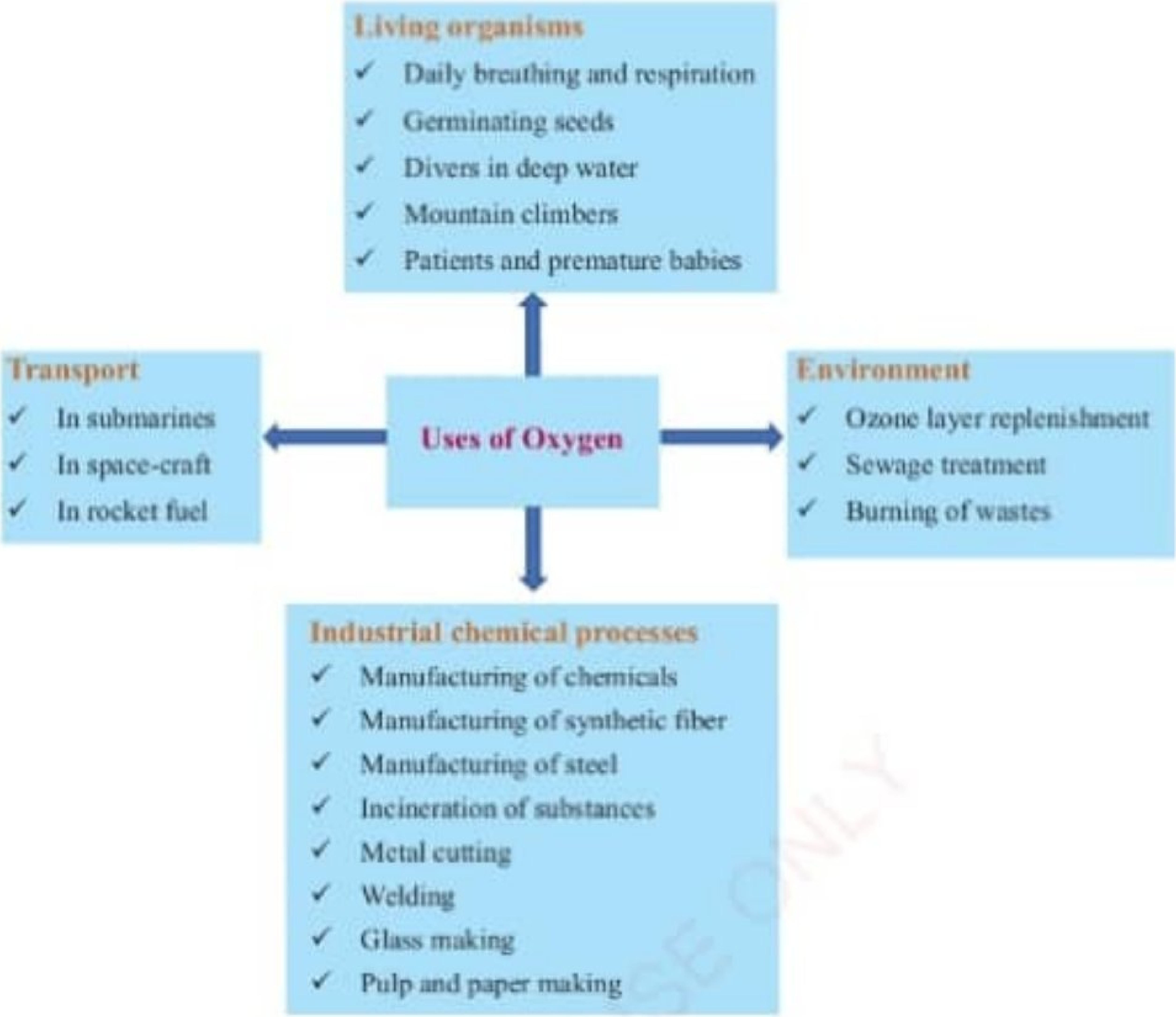
Transport

In transport, oxygen aids respiration for crews in sub-marines and space-craft. Liquid oxygen is also used to burn fuels in rockets. Some of these uses are shown in Figure 1.11.

A submarine *The spacecraft* *A rocket engine*

Figure 1.11: *Uses of oxygen in transport*

The areas in which oxygen is used are summarised in Figure 1.12



Uses of Oxygen

Living organisms
- ✓ Daily breathing and respiration
- ✓ Germinating seeds
- ✓ Divers in deep water
- ✓ Mountain climbers
- ✓ Patients and premature babies

Transport
- ✓ In submarines
- ✓ In space-craft
- ✓ In rocket fuel

Environment
- ✓ Ozone layer replenishment
- ✓ Sewage treatment
- ✓ Burning of wastes

Industrial chemical processes
- ✓ Manufacturing of chemicals
- ✓ Manufacturing of synthetic fiber
- ✓ Manufacturing of steel
- ✓ Incineration of substances
- ✓ Metal cutting
- ✓ Welding
- ✓ Glass making
- ✓ Pulp and paper making

Figure 1.12: *Summary of the uses of oxygen*

Most of the uses of oxygen are related to its properties as follows:

Solubility in water

The usefulness of oxygen to aquatic organisms is due to its solubility in water. Oxygen is slightly soluble in water, therefore, water contains some dissolved oxygen that is used by aquatic living organisms.

Combustibility

Oxygen supports combustion. Therefore, it is used in incineration or burning of substances, welding, and metal cutting. In living organisms, oxygen is used to support burning of food in the body through respiration.

Reactivity with elements

Oxygen reacts with many elements. This makes it useful in industrial chemical processes such as manufacturing of chemicals, glass, pulp, and paper.

Exercise 1.2

1. Explain the physical properties of oxygen.
2. What will happen to a glowing splint if it is lowered into a gas jar containing oxygen gas?
3. What happens when the following elements are burnt in oxygen?
   (a) Calcium (b) Sulphur
4. What do you understand by the following terms?
   (a) Basic oxide (b) Acidic oxide
5. Relate the uses of oxygen to its properties.

Task

1. Collect information on the uses of oxygen from books, newspapers, and other sources.
2. Discuss your findings in groups, and compile them in the form of a detailed report.
3. Present your report to the rest of the class, using pictures and charts, where necessary.

Chapter summary

1. Oxygen constitutes about 21% by volume of the air we breathe, and is essential for combustion and respiration.
2. Oxygen is prepared in the laboratory mainly by the decomposition of hydrogen peroxide with manganese(IV) oxide as a catalyst.
3. A catalyst is a substance that alters the rate of a chemical reaction but remains chemically unchanged at the end of the reaction.
4. Oxygen is colourless, tasteless, odourless and slightly soluble in water.

5. Oxygen supports combustion, and it reacts with many elements to form oxides.
6. Industrial manufacturing of oxygen is mainly by the fractional distillation of liquefied air.
7. The main uses of oxygen are:
   - (a) sustenance of life to living things.
   - (b) in environmental conservation.
   - (c) in transport.
   - (d) in various chemical processes.

Revision exercise 1

1. Write **TRUE** for a correct statement and **FALSE** for an incorrect statement.
   - (a) Manganese(IV) oxide is used in the preparation of oxygen gas.
   - (b) Oxygen is used in deep sea diving.
   - (c) Hydrogen peroxide is a very useful catalyst in the laboratory preparation of oxygen gas.
   - (d) Oxygen combines with metals to form basic oxides.
   - (e) Oxygen supports combustion.
   - (f) Oxygen is prepared in the laboratory mainly through fractional distillation of liquefied air.
2. Choose the word from the box that best matches with each of the following statements:
   - (a) It turns a wet red litmus paper blue.
   - (b) It burns with a brick red flame.
   - (c) It exists as a black solid
   - (d) It alters the speed of a reaction but remains chemically unchanged.
   - (e) It melts and burns with a bright white flame.
   - (f) Its oxide is a black solid.
   - (g) It melts and burns with a blue flame

Catalyst, Calcium, Carbon, Decomposition, Elements, Magnesium, Copper, Non-metal oxide, Metal oxides, Reagent, Zinc, Sulphur

3. How is oxygen prepared using the following chemicals?

   (a) Hydrogen peroxide (b) Potassium chlorate

4. How would you distinguish pure oxygen from ordinary air?

5. What would happen if there were no oxygen in the atmosphere?

6. A student placed a silvery white solid on a deflagrating spoon, ignited it and then lowered the spoon into a gas jar of oxygen. The solid burned with a brick red flame.
   (a) Identify the silvery white substance that burned in oxygen.
   (b) Explain the nature of the product in terms of acidic or basic properties.
   (c) Write the product formed after burning the silvery solid.

7. Oxygen is collected through a downward displacement of water.

   (a) Write a word equation for the preparation of oxygen by the decomposition of hydrogen peroxide using manganese dioxide as a catalyst.

   (b) Is it possible to collect pure oxygen during its preparation? Explain.

   (c) Can all the oxygen formed from hydrogen peroxide be collected into the gas jar? Give reason(s).

8. Oxygen gas can also be prepared by thermal decomposition of potassium chlorate using manganese dioxide as a catalyst.

   (a) Explain the activities which will be done when preparing the gas using this method.

   (b) Draw a well labelled diagram to show how oxygen is prepared using this method.

   (c) Write the word equation for this reaction.

9. Why is hydrogen peroxide preferred to potassium chlorate in the laboratory preparation of oxygen?

10. Most uses of oxygen are dictated by its properties. Explain.

11. Draw a clearly labelled diagram showing the laboratory preparation of oxygen without the application of heat.

Chapter Two

Hydrogen

Introduction

Hydrogen is one of the important and most abundant elements found in the universe. Like oxygen, hydrogen is gaseous in nature and it occurs in different forms. In this chapter, you will learn about the occurrence and nature of hydrogen, laboratory preparation, properties, industrial production, and uses of hydrogen. The competencies developed will enable you to use and manipulate hydrogen and other substances in your daily life activities.

Occurrence and nature of hydrogen

The word *hydrogen* comes from the Greek words *hydro*–meaning "water" and *genes*–meaning "creator" or "generator". This was after a Greek scientist Antoine Lavoisier, who discovered that hydrogen was produced when water was decomposed. Hydrogen is the lightest and most abundant element in the universe. Hydrogen gas is lighter than air, and therefore, it rises high in the atmosphere. This is why hydrogen gas is not found free on its own on the Earth's surface and in the lower atmosphere. Hydrogen gas is colourless and odourless. It can be collected by downward displacement of water because it is lighter than water and slightly soluble in it. If the container containing hydrogen is left open, it can escape because hydrogen is lighter than air. Moreover, hydrogen is a very reactive element, but at room temperature the reaction rates are usually so low as to be negligible. This is why it is found in combination with many other elements, forming different substances. For example, hydrogen is found in large quantities in the form of water on Earth. It is the main element from which the sun and the stars are made.

Hydrogen is found in combination with carbon that leads to the formation of organic compounds such as coal, petroleum, natural gas, and other compounds. It is also present in acids and in some bases. It can be prepared in the laboratory. Hydrogen can be tested by lighting it in air in which it ignites with a "pop" sound explosion.

Laboratory preparation of hydrogen

Hydrogen can be prepared in the laboratory in different ways. There are four main methods for laboratory preparation of hydrogen, which are; the reactions of dilute acids with some metals, the reactions of water with some metals, the reaction of water with hot carbon, and the electrolysis of water. The most common method of preparation of hydrogen is by the actions of dilute acids on metals. An example is the action of dilute hydrochloric acid on zinc. This reaction can be summarised by the following word and formula equations:

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

$$Zn + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(l)$$

Activity 2.1

Aim: To prepare hydrogen in the laboratory by reacting dilute hydrochloric acid with zinc granules.

Requirements: Flat-bottomed flask, thistle funnel, gas jars and their lids, water trough, beehive shelf, two-holed rubber bung, zinc granules, dilute hydrochloric acid, and water

Procedure

1. Put some zinc granules into a flat-bottomed flask.
2. Fill a gas jar with water and invert it over the beehive shelf in the water trough.
3. Set up the rest of the apparatus as shown in Figure 2.1.
4. Add dilute hydrochloric acid to the zinc granules.
5. Collect the gas over water, ensuring that you only remove the gas jar when it is full, and that you keep the jar tightly closed with a lid.

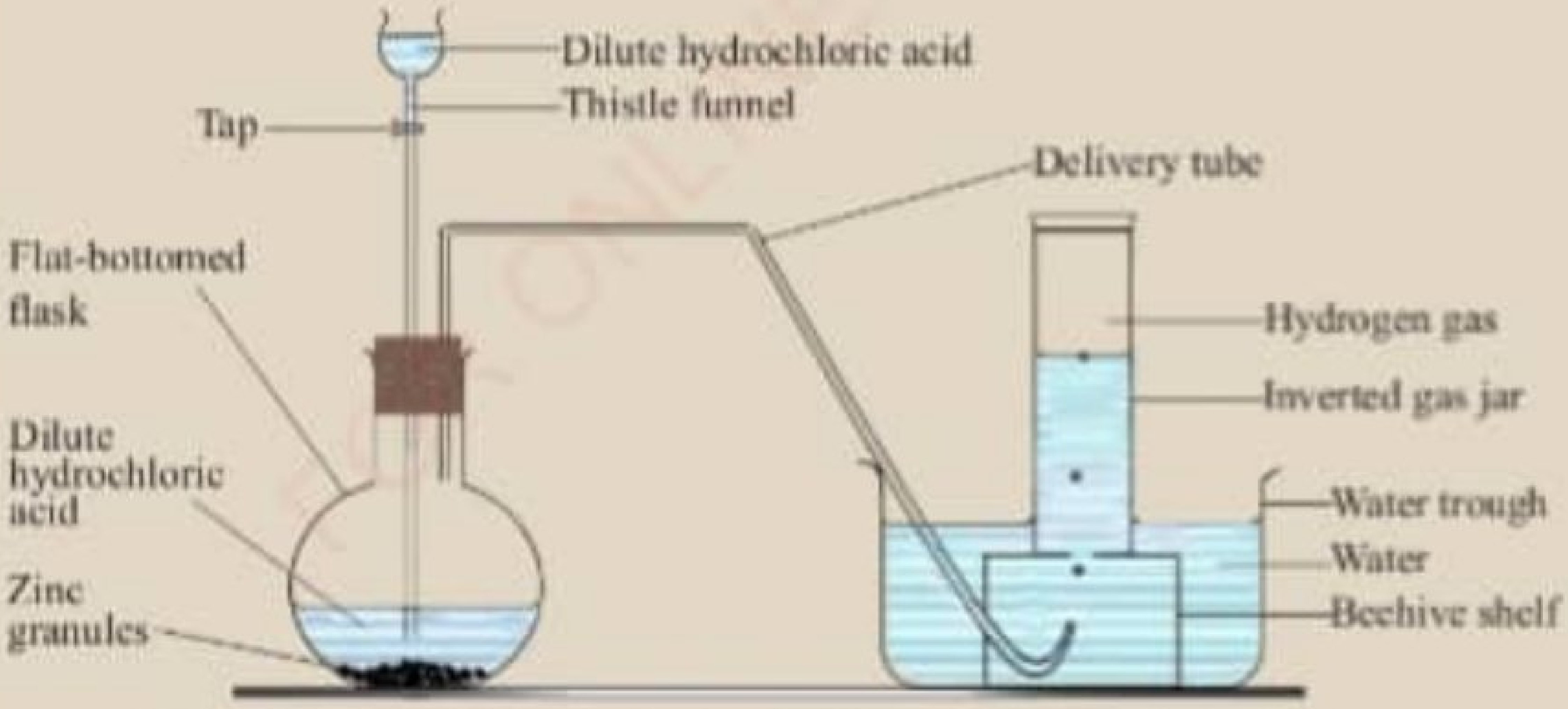


Figure 2.1: *Experimental set-up for the laboratory preparation of hydrogen gas*

Questions

1. What is the colour of the gas produced?
2. Why is it possible to collect hydrogen using this method?
3. What would happen if a gas jar containing hydrogen gas was not tightly closed?

Exercise 2.1

1. Why does hydrogen rise high in the atmosphere?
2. Hydrogen is not often found free on its own on the Earth's surface, instead it is found in combination with many other elements. Explain.
3. Why is hydrogen gas collected by downward displacement of water?
4. Why should a gas jar of hydrogen be tightly closed with a lid?
5. Among the substances formed by the combination of hydrogen with other elements are organic compounds. Give at least ten examples of such compounds.

Activity 2.2

Aim: To test some properties of hydrogen.

Requirements: Flat-bottomed flask, thistle funnel, gas jars and their lids, water trough, beehive shelf, two-holed rubber bungs, test tubes, blue and red litmus papers, and zinc granules

Procedure

1. Set up the apparatus as shown in Figure 2.1 of Activity 2.1.
2. Collect the hydrogen gas over water, ensure that you only remove the gas jar when it is full.
3. Collect some of the gas in the test tubes and stopper with rubber bungs as shown in Figure 2.2.

Figure 2.2: *Test tubes filled with hydrogen gas*

4. Place a burning splint at the mouth of one of the test tubes. Record your observation.

5. Take another sample and waft some of the gas to your nose. Record your observation.

6. Remove the stopper then invert a test tube containing air over the test tube containing hydrogen (Figure 2.3). After one minute, test the gases in the test tubes using a burning splint. Record your observations.

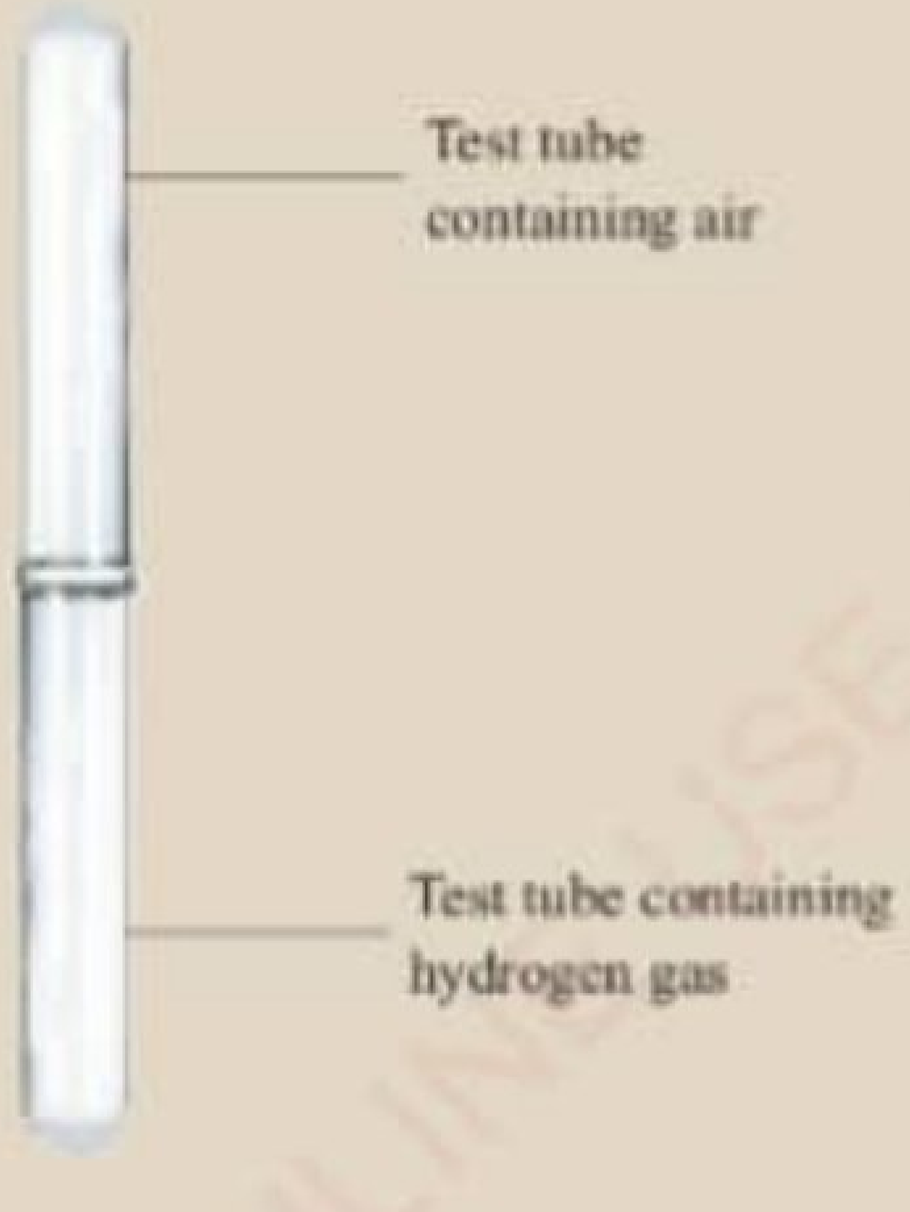


Figure 2.3: *Test tube filled with air inverted over the one filled with hydrogen gas*

7. Place moist blue and red litmus papers in another test tube containing hydrogen gas. Record your observations.

Note: For safety, do not taste the gas in the laboratory.

Questions

1. What happens when a burning splint is placed at the mouth of the test tube containing hydrogen gas?
2. Describe the smell of hydrogen gas.
3. What happens when a burning splint is put at the mouth of inverted test tube in step 6?
4. What happens to the moist blue and red litmus papers placed in a gas jar containing hydrogen gas?

Properties of hydrogen

The physical and chemical properties of hydrogen gas are as follows:

Physical properties

Hydrogen has the following physical properties:

1. It is colourless, odourless and tasteless.
2. It is lighter than air; it has a density of 0.0899 g/dm^3 compared to air which has a density of 1.225 g/dm^3 at standard temperature and pressure.
3. It is slightly soluble in water.

Chemical properties of hydrogen

Hydrogen has many chemical properties, which include:

1. At high temperatures, hydrogen combines easily with other chemical substances.
2. It does not usually react with other elements at room temperatures.
3. It is highly flammable and burns with a blue flame. However, under the presence of some impurities in air, the flame may appear yellow.
4. A mixture of hydrogen and oxygen explodes when lit.
5. It reacts slowly with oxygen to produce water. A catalyst can be used to speed up the reaction.
6. It is neither basic nor acidic. It is neutral to litmus papers.

7. It reacts with oxides and chlorides of many metals to produce free metals.
8. It does not support combustion but ignites with a "pop" sound explosion.
9. Dry hydrogen reduces metal oxides into their free metals.

The commonly used drying agent in the reduction reactions is anhydrous calcium chloride. During the reduction process, hydrogen is used as a reducing agent. This means that it removes oxygen from its substances. *Reduction* is the removal of oxygen from a substance or addition of hydrogen to a substance. In the reduction of metal oxides using hydrogen, water is produced in the form of steam. When the reaction is complete, hydrogen is allowed to flow until the new formed metal cools. This helps to prevent the metal from being oxidized in the air to its oxide. *Oxidation* is the addition of oxygen to a substance or removal of hydrogen from a substance.

Reduction reaction can be exemplified by the reactions of copper(II) oxide and lead(II) oxide with hydrogen. Copper(II) oxide is reduced to copper metal as shown in the following chemical equations:

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

$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(g)$$

On the other hand, lead(II) oxide is reduced to lead metal as described by the folowing equations:

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

$$PbO(s) + H_2(g) \longrightarrow Pb(s) + H_2O(g)$$

Activity 2.3

Aim: To investigate the products formed when dry hydrogen is passed over heated metal oxides.

Requirements: Combustion tube, U-tube, rubber bungs, porcelain bowl, bent glass tubes, dry copper(II) oxide, lead oxide, anhydrous copper(II) sulphate, source of hydrogen gas, and anhydrous calcium chloride

Procedure

1. Set up the apparatus as shown in Figure 2.4.
2. Neatly pack anhydrous calcium chloride in a U-tube.

3. Put dry copper(II) oxide in a porcelain bowl and place it in a combustion tube. Also, place anhydrous copper(II) sulphate in the combustion tube.
4. Ignite the hydrogen gas at the end of the combustion tube.
5. Heat the copper(II) oxide in the combustion tube until there is no more change.
6. Stop the heating, but let the hydrogen gas continue passing through the tube until it cools down. Record all your observations.
7. Repeat the experiment using lead(II) oxide instead of copper(II) oxide.

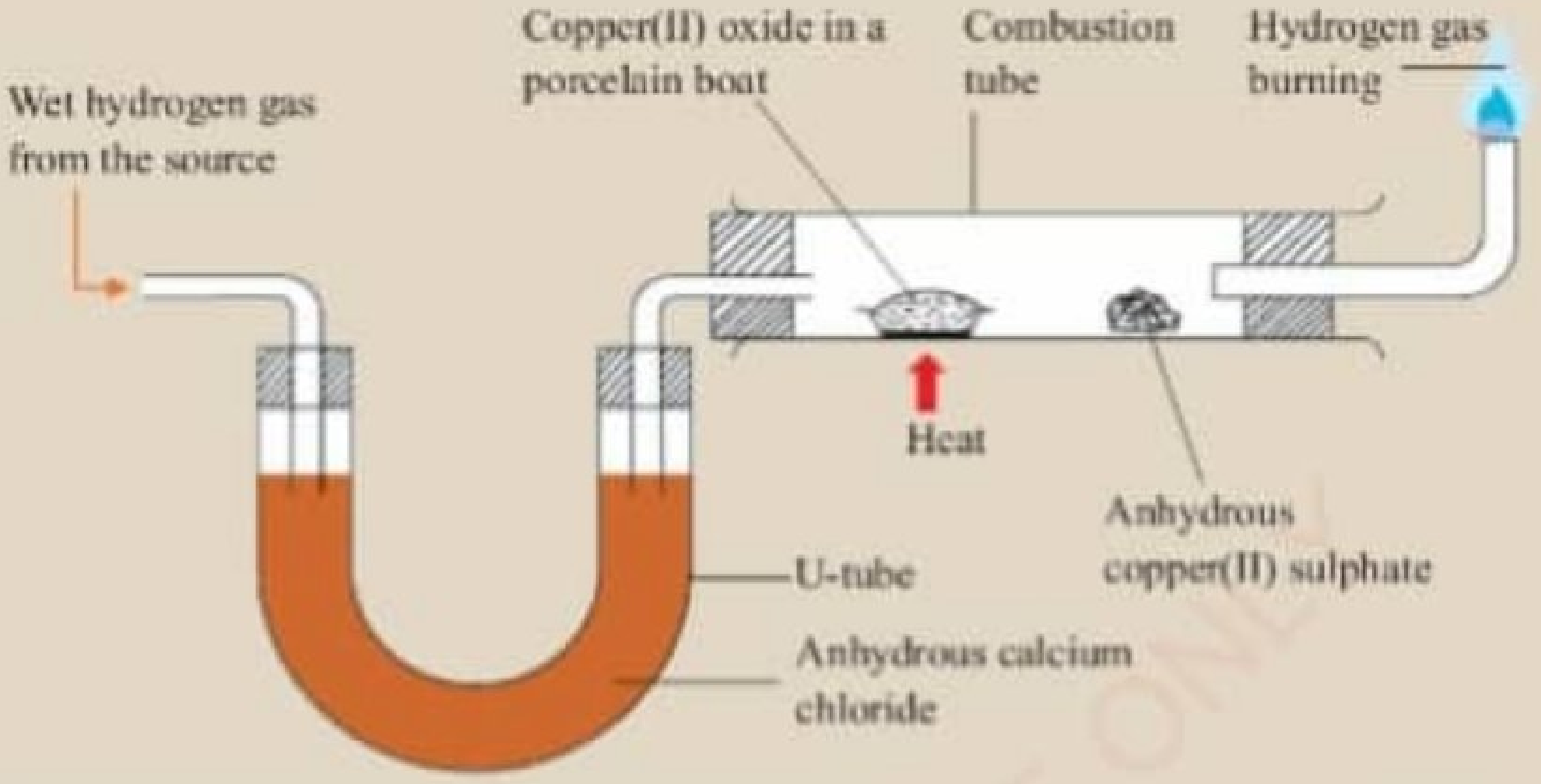


Figure 2.4: *Experimental set-up for the reaction between dry hydrogen gas and heated copper(II) oxide*

Note: In each case, water in the form of steam is produced as the other product. This can be evidenced by anhydrous copper(II) sulphate turning blue.

Questions

1. What was the colour of the flame of the lit hydrogen?
2. Why is the hydrogen allowed to continue flowing even when the heating is stopped in this experiment?
3. What are the products in each of the reactions in this experiment?
4. What would happen if anhydrous calcium chloride was not used in this experiment?
5. What is the colour change when copper(II) oxide is heated?
6. What is the use of anhydrous copper(II) sulphate in this experiment?

Exercise 2.2

1. Write **TRUE** for a correct statement and **FALSE** for an incorrect statement.
   (a) Hydrogen is less dense than air.
   (b) Hydrogen supports combustion.
   (c) A mixture of hydrogen and oxygen burns with a hot blue flame.
   (d) Hydrogen is found in the largest amount on the Earth's surface.
   (e) Hydrogen is used in making margarine because it has a good taste.
2. Why is hydrogen gas collected over water?
3. After collecting hydrogen gas in the gas jar, it is necessary to cover it with a lid. Explain.
4. Is hydrogen gas basic or acidic? Justify.

Industrial production of hydrogen

Pure hydrogen gas is manufactured industrially by the *electrolysis of water* or by the *steam reforming* of natural gas (methane).

Electrolysis of water

Electrolysis of water is a process that decomposes water into oxygen and hydrogen gas by means of an electric current. The electric current is passed through the water. The electrical power source is connected to two plates (called electrodes) that are placed in the water. Hydrogen is collected at the negative plate (the *cathode*), while oxygen collects at the positive plate (the *anode*). The set-up for the electrolysis of water is called the *Hofmann voltameter* (Figure 2.5).

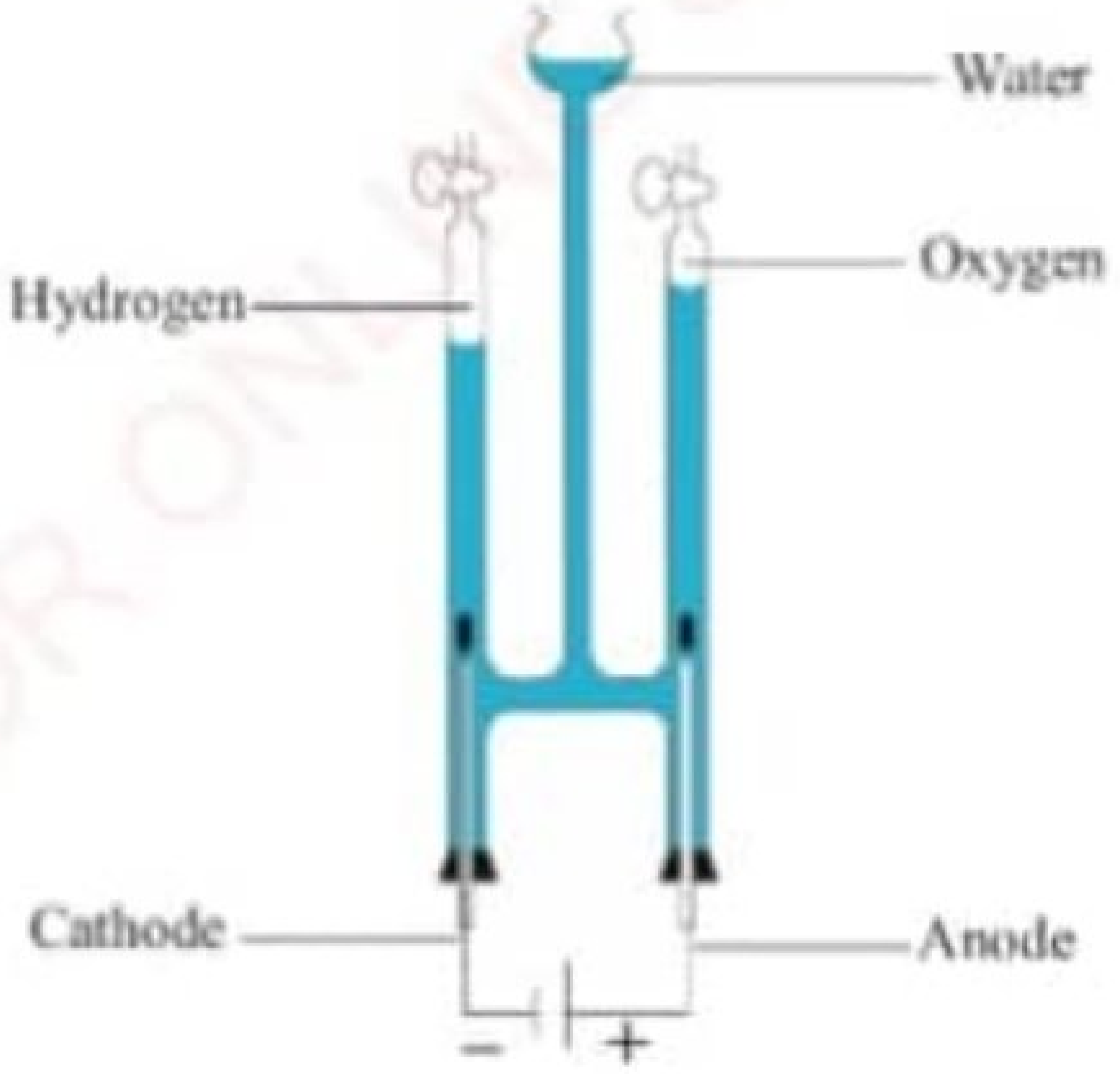


Figure 2.5: *Set-up of a Hofmann voltameter for the electrolysis of water*

Figure 2.7: *Products made of synthetic fibres*

Manufacturing of margarine

Hydrogen is used in the manufacturing of margarine (hardening of oil) by bubbling it through liquid oil in the presence of nickel as a catalyst. This process is called *hydrogenation*. Examples of margarine are shown in Figure 2.8.

Figure 2.8: *Examples of margarine*

Welding and metal cutting

Hydrogen combines with oxygen to produce the oxy-hydrogen flame. This flame is very hot and the temperature can rise up to 3000 °C. This flame can be used for welding and metal cutting as shown in Figure 2.9.

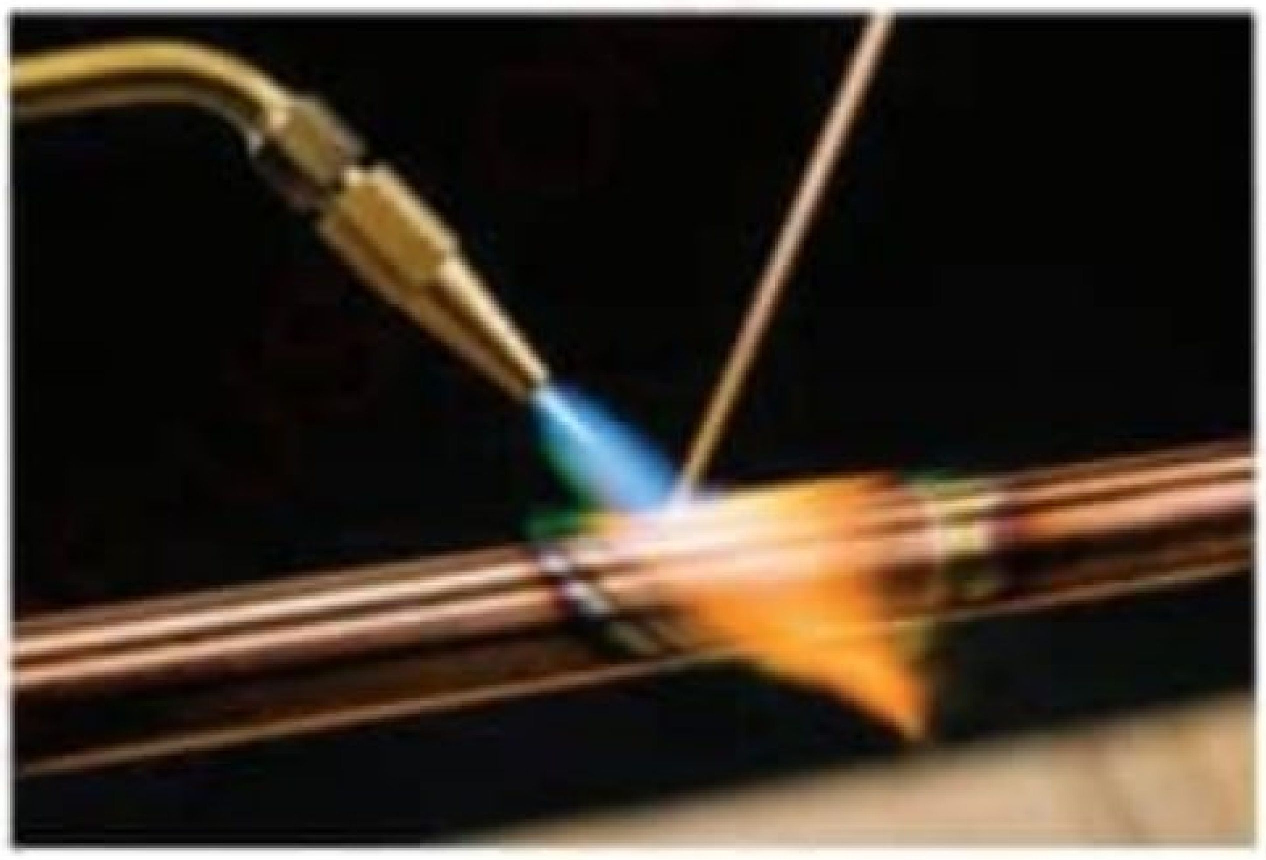


Figure 2.9: *Welding of metals using oxy-hydrogen flame*

Manufacturing of hydrochloric acid

Hydrogen is used in the manufacturing of hydrochloric acid. It reacts with chlorine to form hydrogen chloride gas, which is then dissolved in water to form hydrochloric acid as shown in the following equations:

Hydrogen gas + chlorine gas ⟶ Hydrogen chloride gas

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$$

Hydrogen chloride gas + Water ⟶ Hydrochloric acid

$$HCl(g) + H_2O(l) \longrightarrow HCl(aq)$$

A plant for the manufacturing of hydrochloric acid is shown in Figure 2.10, and the packaging bottles are shown in Figure 2.11.

Figure 2.10: *Hydrochloric acid manufacturing plant*

Figure 2.11: *Packaging bottles for hydrochloric acid*

Hydrogen as a fuel

Hydrogen is used to prepare water gas which is the mixture of carbon monoxide and hydrogen. Water gas can be used as a fuel. It can be burnt to propel rockets. Figure 2.12 shows a rocket that uses water gas being launched.

Figure 2.12: *Rocket that uses water gas being launched*

Filling weather balloons

Hydrogen is a light gas, therefore, it is used by meteorologists to fill weather balloons. The balloons carry instruments that record information on various elements of weather in the upper atmosphere. Figure 2.13 shows a weather balloon filled with hydrogen.

Figure 2.13: *Weather balloon filled with hydrogen*

The main uses of hydrogen can be linked to its properties. Table 2.1 summarises the relationships between some uses and the properties of hydrogen.

Table 2.1: *The relationships between some uses of hydrogen and its properties*

S/N	Use	Property
1.	Manufacturing of ammonia	Readily reacts with other substances, for example, nitrogen and chlorine
2.	Manufacturing of hydrochloric acid	
3.	Production of oxy-hydrogen flame	Highly flammable
4.	Preparation of water gas	
5.	Filling weather balloons	Lighter than air
6.	Manufacturing of margarine	Reducing agent

Exercise 2.3

1. What is the test for hydrogen?
2. Give examples of substances made by the reactions or combination of hydrogen with other substances.
3. Suppose there were no hydrogen in the universe, what would happen?
4. Outline any four uses of hydrogen.
5. Comment on the fact that most of the uses of hydrogen are related to its properties.
6. You have learnt that hydrogen is used in welding. What do you understand by this term?

Task

1. Carry out searches from books, newspapers and other sources on the uses of hydrogen. Note down your findings.
2. In groups, discuss your findings and compile your reports that include illustrations (pictures and diagrams).

Chapter summary

1. Hydrogen is a colourless, odourless, and highly flammable gas.
2. Hydrogen can be produced in the laboratory by the reaction of dilute hydrochloric acid with some metals such as zinc.
3. Reduction is the removal of oxygen from a substance or the addition of hydrogen to a substance.
4. Oxidation is the addition of oxygen to a substance or the removal of hydrogen from a substance.
5. Hydrogenation is the process of passing hydrogen through liquid oil to harden it.
6. Pure hydrogen is manufactured industrially by electrolysis of water or by steam reforming of natural gas (methane).

Exercise 2.3

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5. Hydrogenation is the process of passing hydrogen through liquid oil to harden it.
6. Pure hydrogen is manufactured industrially by electrolysis of water or by steam reforming of natural gas (methane).

7. Electrolysis of water is a process which decomposes water into oxygen and hydrogen gases, with the aid of an electric current.
8. Hydrogen has many uses. These include the manufacturing of ammonia, margarine, hydrochloric acid, and water gas. It is also used in weather balloons and in the production of the very hot oxy-hydrogen flame.

Revision exercise 2

1. Write **TRUE** for a correct statement and **FALSE** for an incorrect statement.
   (a) Hydrogen gas is slightly denser than air.
   (b) Hydrogen reacts with chlorine to give hydrochloric acid.
   (c) Hydrogen gas is found in compound forms on the Earth's surface, and in the lower atmosphere.
   (d) The Haber process is used in the industrial manufacturing of ammonia.
   (e) Reduction is the removal of hydrogen from a substance.
   (f) Hydrogen gas is the most abundant gas on the Earth's surface.
   (g) Hydrogen is used in the manufacturing of methanol, which is used to make plastics and fertilisers.
   (h) Hydrated copper(II) sulphate is used to test for the presence of water or moisture.
   (i) Addition of oxygen to a substance is reduction.
   (j) Hydrogen is used to make water vapour for powering rocket engines.
2. Choose the correct answer for each of the following items:
   (i) The name hydrogen originates from the Greek words "*hydro*" and "*genes*" meaning
      (a) water fearing gas.
      (b) water forming gas.
      (c) fire forming gas.
      (d) electric power generator.

(ii) In the laboratory preparation, hydrogen is collected through
  (a) upward delivery.
  (b) downward displacement of water.
  (c) the Bunsen burner.
  (d) the delivery tube.

(iii) The name given to a colourless, and highly flammable gas used in the production of ammonia is
  (a) hydrogen.
  (b) oxygen.
  (c) carbon dioxide.
  (d) helium.

(iv) The common method used in industrial production of hydrogen gas is called
  (a) electrode.
  (b) steam reforming.
  (c) reduction.
  (d) decomposition.

(v) Which of the following is not true about hydrogen?
  (a) It is odourless.
  (b) It is lighter than air.
  (c) It supports combustion.
  (d) It burns with a pale blue flame.

(vi) Which of the following gases if mixed with hydrogen, would produce a very hot flame of up to 3000 °C?
  (a) Oxygen
  (b) Neon
  (c) Chlorine
  (d) Argon

(vii) Which of the following cannot be used to prepare hydrogen gas?
  (a) Reaction of water with carbon at room temperature
  (b) Reaction of dilute acids with zinc

(c) Electrolysis of water
(d) Reaction of water with certain metals

(viii) Due to its lightness, hydrogen is used in
(a) rocket engines.
(b) making water.
(c) weather balloons.
(d) margarines.

(ix) Which of the following compounds is not likely to contain hydrogen?
(a) Ammonia
(b) Water
(c) Water gas
(d) Zinc granules

(x) Where is hydrogen likely to be found in its free state?
(a) In the upper atmosphere
(b) Near the Earth's surface
(c) In the lower atmosphere
(d) In the sun and the stars

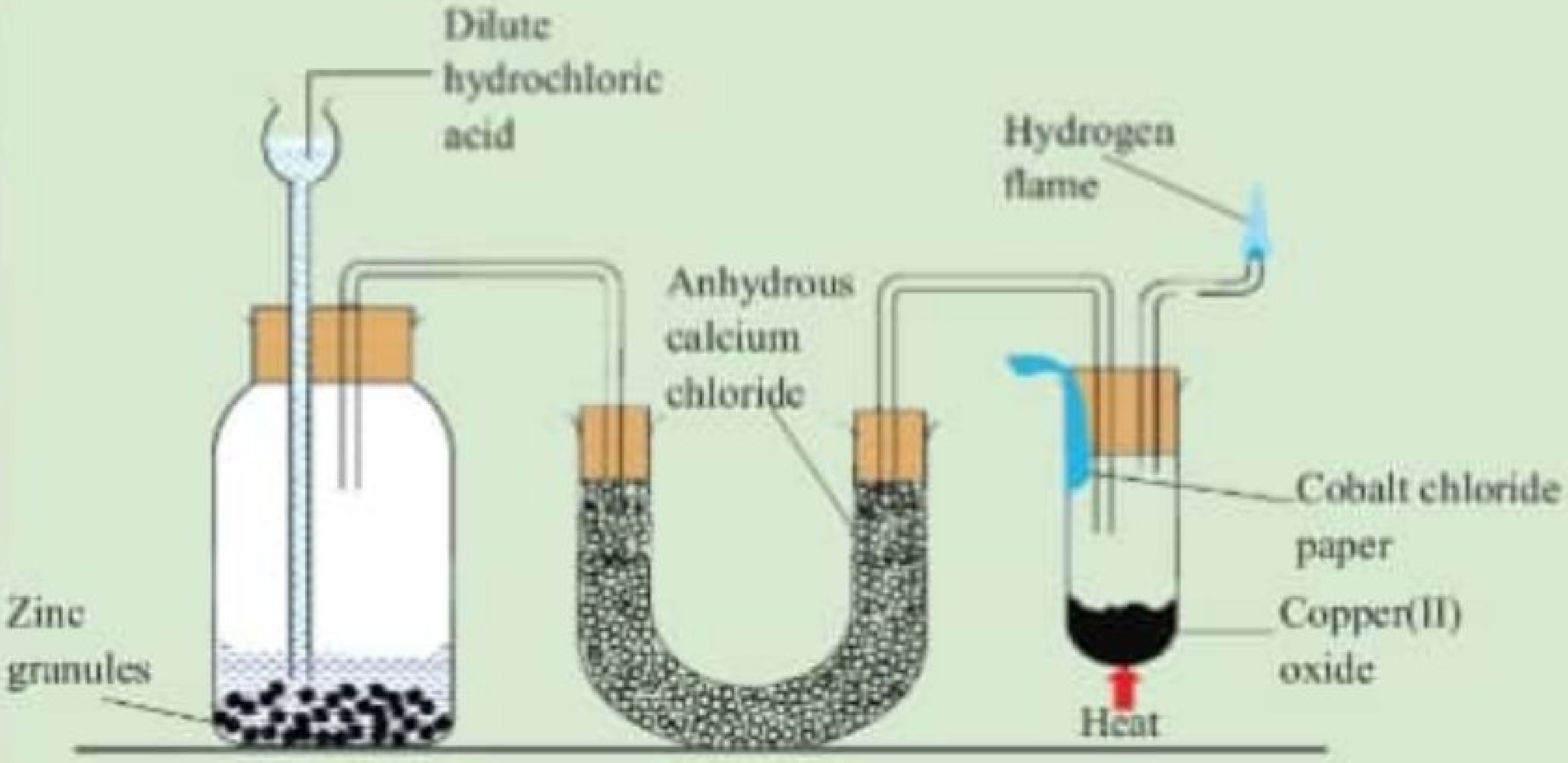
3. The main uses of hydrogen can be linked to its various properties. Match each use in list A against the related property from list B.

List A	List B
(a) Inflating weather balloons	(i) It readily combines with other elements.
(b) Manufacturing of ammonia	(ii) It is denser than air.
(c) Manufacturing of margarine	(iii) It is lighter than air.
(d) Production of oxy-hydrogen flame	(iv) It is an oxidizing agent.
	(v) It is highly flammable.
	(vi) It is a reducing agent.
	(vii) It burns with a blue flame.
	(viii) It relights a glowing splint.

4. State the physical properties and chemical properties of hydrogen.

5. Describe two methods for the industrial manufacturing of hydrogen.

6. Study the diagram below and answer the questions that follow.



(a) What is the colour of the solid product?

(b) Name the products formed.

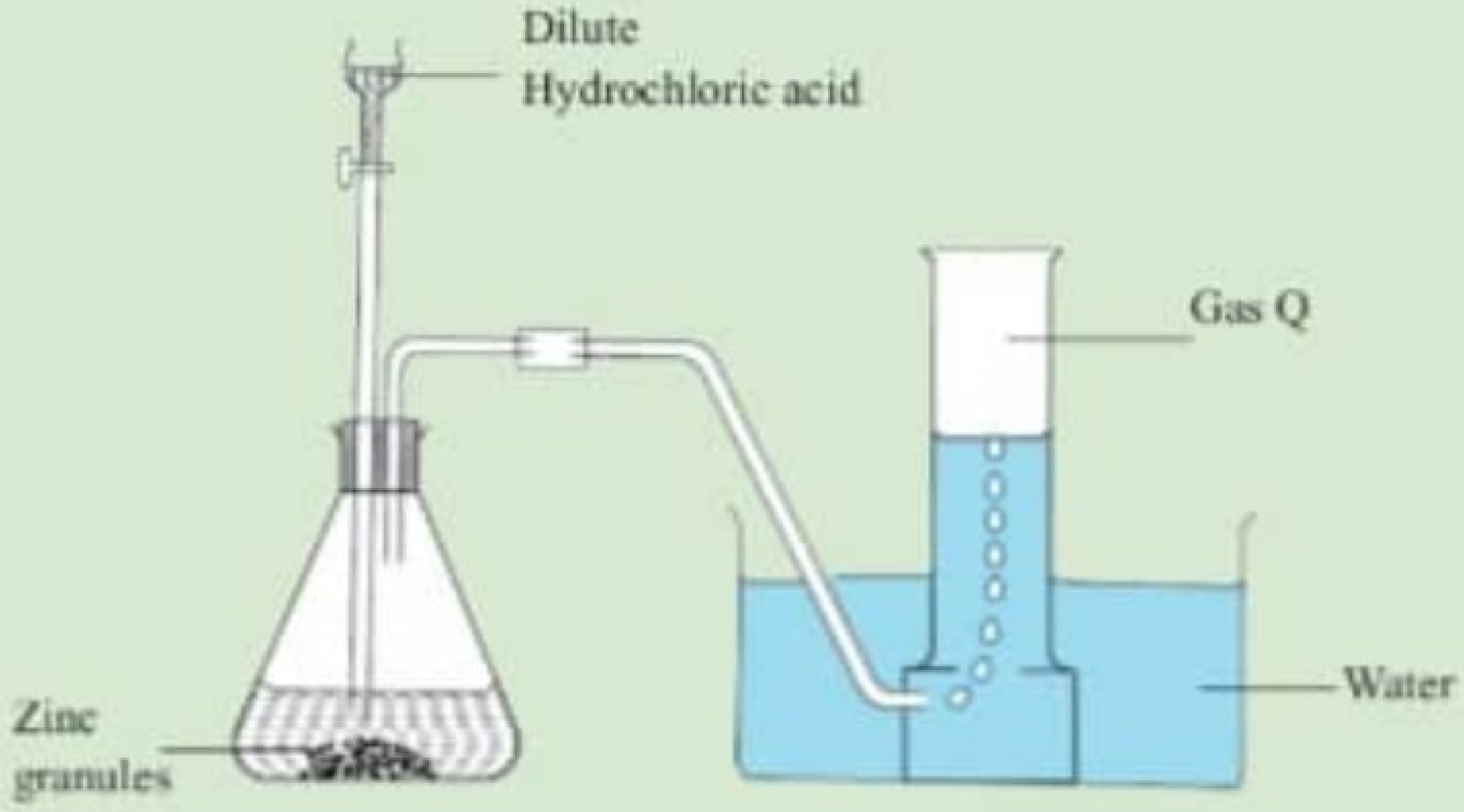
(c) What is the role of the following chemicals?
  (i) Hydrochloric acid and zinc granules
  (ii) Anhydrous calcium chloride
  (iii) Cobalt chloride paper

7. Hydrogen gas is a very promising energy source, yet its uses as a major source of energy are very limited. Explain this in terms of its storage, safety and production.

8. Briefly describe two methods of large-scale production of hydrogen gas.

9. Explain the origin of the term hydrogen.

10. The following figure shows a set-up for the preparation of gas Q in the laboratory:



(a) Identify gas Q.

(b) What properties of Q make it possible to be collected as shown in the figure?

(c) Describe the properties of gas Q which relate with its uses.

Chapter Three

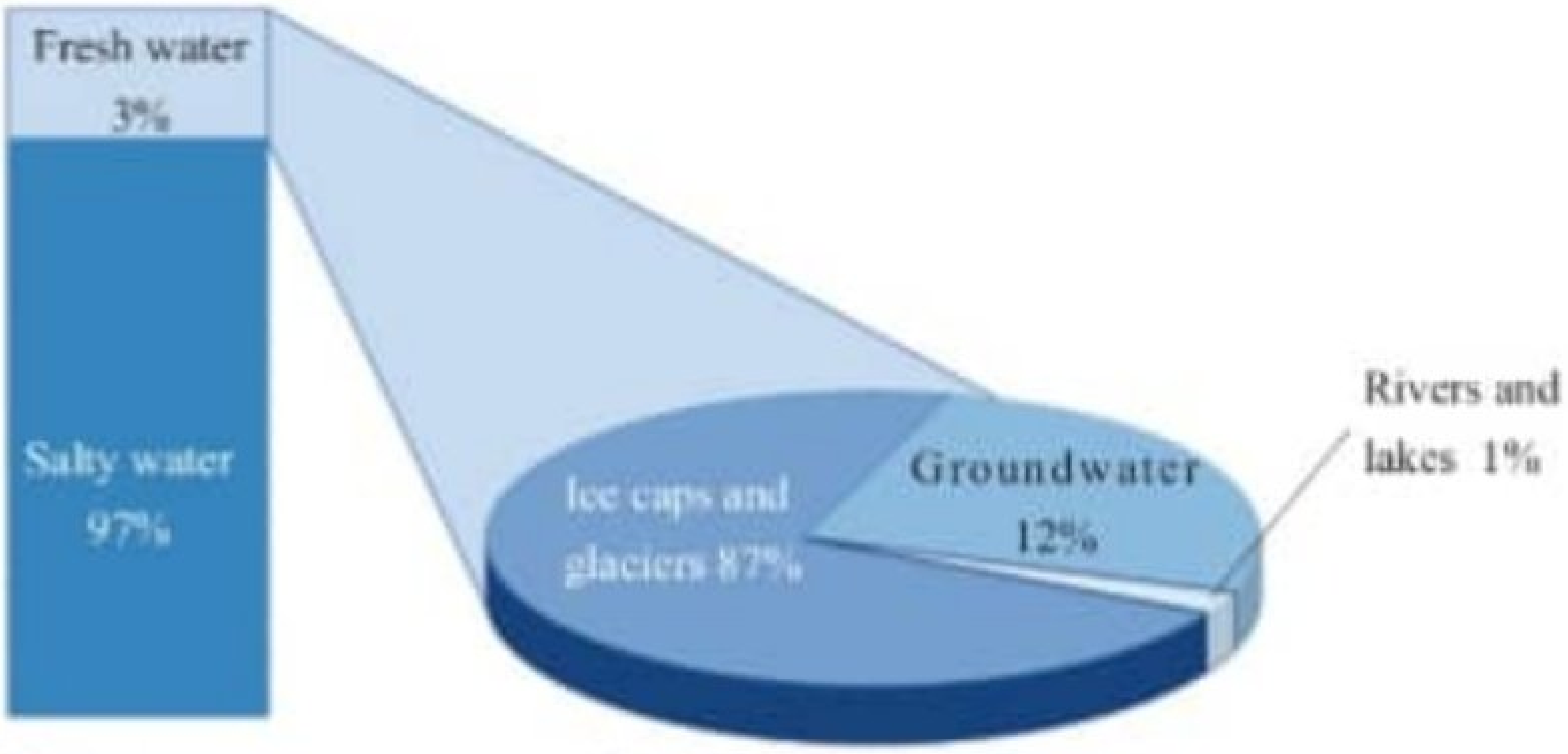
Water

Introduction

Water is an inorganic chemical substance composed of hydrogen and oxygen atoms. It can exist in three main states, which are gaseous, liquid, and solid. It is essential for sustainability of life for all living things. In this chapter, you will learn about the occurrence and nature of water, its properties, water cycle, and the relationship between water cycle and environmental conservation. Moreover, you will learn about the uses of water and the importance of water treatment and purification. The competencies developed will enable you to protect water sources and use them sustainably.

Occurrence and nature of water

Water is one of the most plentiful and essential compounds on Earth. It is essential for the sustenance of all living things. Apart from being a habitat for some animals and plants, it is also a major constituent of the bodies of living things. Water occurs in three main states: *solid*, for example ice, snow and hail; *liquid*, for example dew, mist, and rain; and *gaseous*, for example steam or vapour. About 97% of Earth's water is saline (salty), while only 3% is fresh water. However, out of the fresh water that is appropriate for most of our daily uses, some of it is not easily accessible. About 87% of the fresh water is ice, 12% is groundwater and only 1% is fresh water which originates from rivers and lakes. Groundwater, is not easily accessible for use, whereas the fresh water from various sources such as rivers, lakes and ponds is easily accessible but highly prone to contamination. Figure 3.1 shows the distribution of the Earth's water in terms of the salty water and the fresh water.



(a) *Earth's water* (b) *Fresh water*

Figure 3.1: *Distribution of Earth's water*

The water cycle

Water is continually moving above and below the Earth's surface, as water vapour, liquid water and ice. It is never lost, but is continually being recycled all-round the globe in different systems. This phenomenon is called the *water cycle (hydrological cycle)*. The water cycle goes repeatedly through four main stages, namely evaporation, condensation, precipitation, and collection. Figure 3.2 shows a schematic diagram of the water cycle.

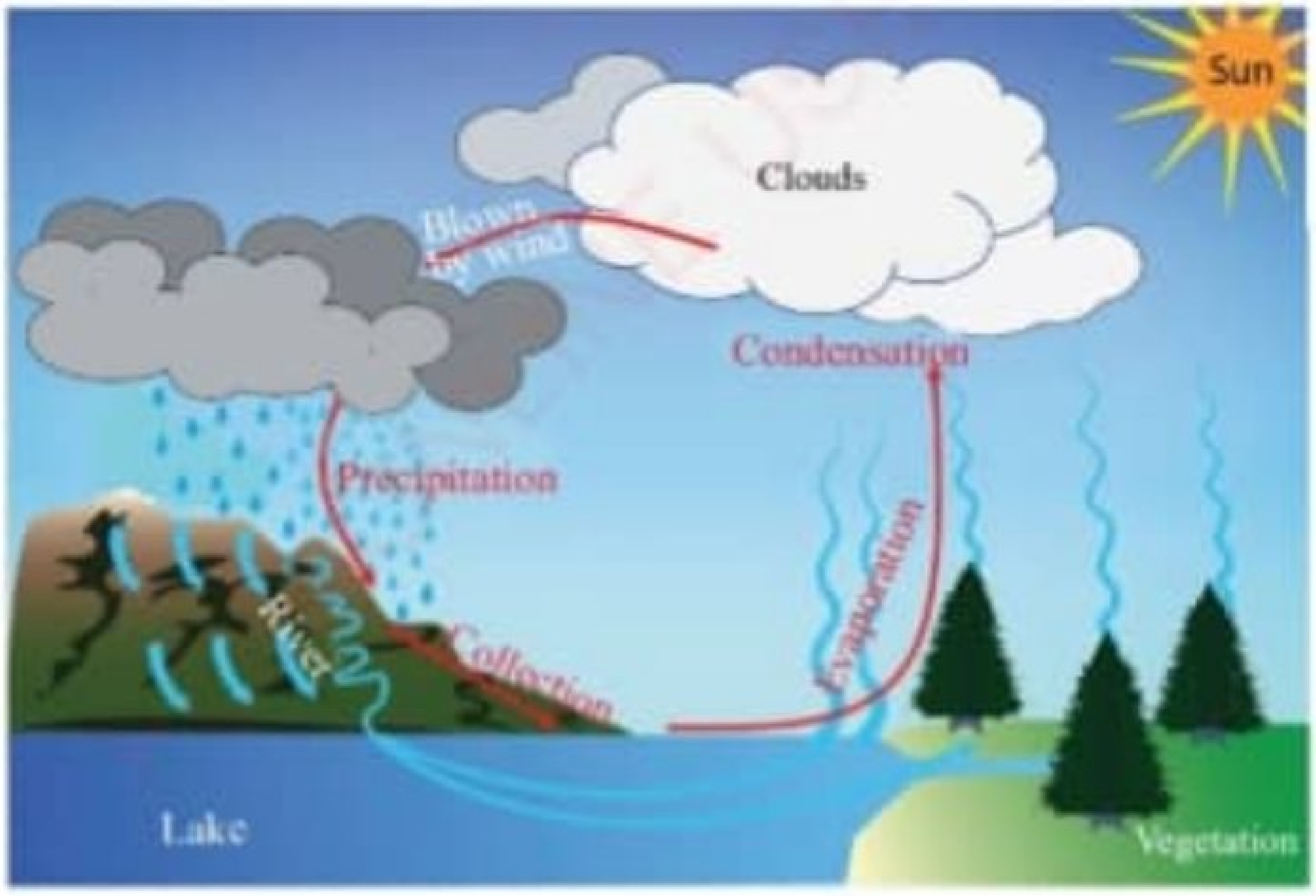


Figure 3.2: *Water cycle*

Evaporation

Evaporation is the process whereby liquid water changes into vapour or steam. This process can also occur in plants by transpiration. *Transpiration* is a process whereby water moves from the inner part of a plant through its leaves and then to the atmosphere through evaporation. Human beings and animals also lose water to the air through respiration and sweating. The sun provides the energy for heating water bodies and turning it into vapour or steam.

Condensation

Condensation occurs when vapour meets the cold condition of the atmosphere. After evaporation, the water vapour in the atmosphere cools into liquid, forming water drops and clouds. Figure 3.3 shows a photograph of clouds formed due to condensation process.

Figure 3.3: *Photograph of clouds formed due to condensation*

Precipitation

Precipitation occurs when the condensed atmospheric water falls under gravitational pull from clouds. It occurs in different forms, such as rain, hail, and snow as shown in Figure 3.4.

Rain

Hail

Snow at the top of a mountain

Snow on the ground

Figure 3.4: *Forms of precipitation*

Collection

After the water falls back to Earth through precipitation, it may end up in the oceans, lakes, rivers, ponds or on land. When it falls on land, it normally infiltrates into the soil and become part of the groundwater through which plants and some animals use it. Also, run-off may occur and water will be collected in the oceans, lakes or rivers, where evaporation takes place, and thus, the cycle starts all over again.

Water cycle and environmental conservation

Water bodies

Environmental degradation destroys the quality of water in the sources. Different practices lead to water pollution and water unavailability for different purposes. Examples of such practices include marine dumping, mining activities, burning of fossil fuels, urban development, and the uses of fertilisers and pesticides. These practices may hinder the suitability and usability of water. Water pollution may occur due to inputs of soluble and insoluble substances. Figure 3.5 shows part of the polluted water body with different types of pollutants.

Figure 3.5: *Part of a polluted water body*

Water vapour and pollutants

Water evaporates from various water bodies to form water vapour in the atmosphere. Gases such as sulphur dioxide, nitrogen dioxide and carbon dioxide combine with the water vapour to form fog which makes visibility difficult. Therefore, pollutants need to be controlled not to contaminate the atmosphere.

Acid rain

Acidic gases such as sulphur dioxide, carbon dioxide and nitrogen dioxide present in air dissolve in water vapour to form acid rain. Acid rain kills plants, animals and other living things in water bodies. Acid rain also accelerates the destruction of building materials like iron sheets and paints. Measures should be taken to control the release of acidic gases to the atmosphere.

Conservation measures

Water bodies should not be contaminated with pollutants. The wastes should be treated, recycled, and disposed-off accordingly. The industrial and domestic discharges should be limited. The gasous wastes should be treated or recycled instead of emitting them directly into the atmosphere.

Properties of water

Like most substances, water has its physical and chemical properties. The physical properties involve aspects such as colour, taste and smell. They also include melting, freezing and boiling points. The chemical properties involve the the behaviour of water when it is reacted with other substances.

Physical properties

The following are the physical properties of water:

1. It is colourless, odourless and tasteless.
2. It is the only substance that occurs naturally in all the three states of matter (solid, liquid and gas).
3. Pure water freezes at 0 °C and boils at 100 °C at standard pressure.
4. It expands (increases in volume) when it freezes. Ice is therefore less dense than liquid water.
5. Water dissolves more substances than any other liquid and is usually called the *universal solvent.*
6. It has a high surface tension. This means that water molecules have high cohesion forces which tend to clump together the water molecules in drops rather than spread out in a thin film.
7. It has a high specific heat capacity. This means it can absorb a lot of heat before it begins to get hot.
8. It is miscible with many liquids. Examples of liquids which are completely miscible with water include ethanol, acetone, acetonitrile, and methanol.

Activity 3.1

Aim: To measure the melting and boiling points of water.

Apparatus: Beaker, Bunsen burner, tripod stand, thermometer, retort stand and clamp, wire gauze, glass rod, stopwatch, and ice cubes

Procedure

1. Put some ice cubes in a beaker.
2. Clamp the thermometer in a vertical position but it should not touch the bottom of the beaker to cause direct heating of the thermometer.

3. Set the apparatus as illustrated in Figure 3.6. Adjust the thermometer so that it dips into the ice cubes, but does not touch the beaker. Record the temperature of the ice cubes.

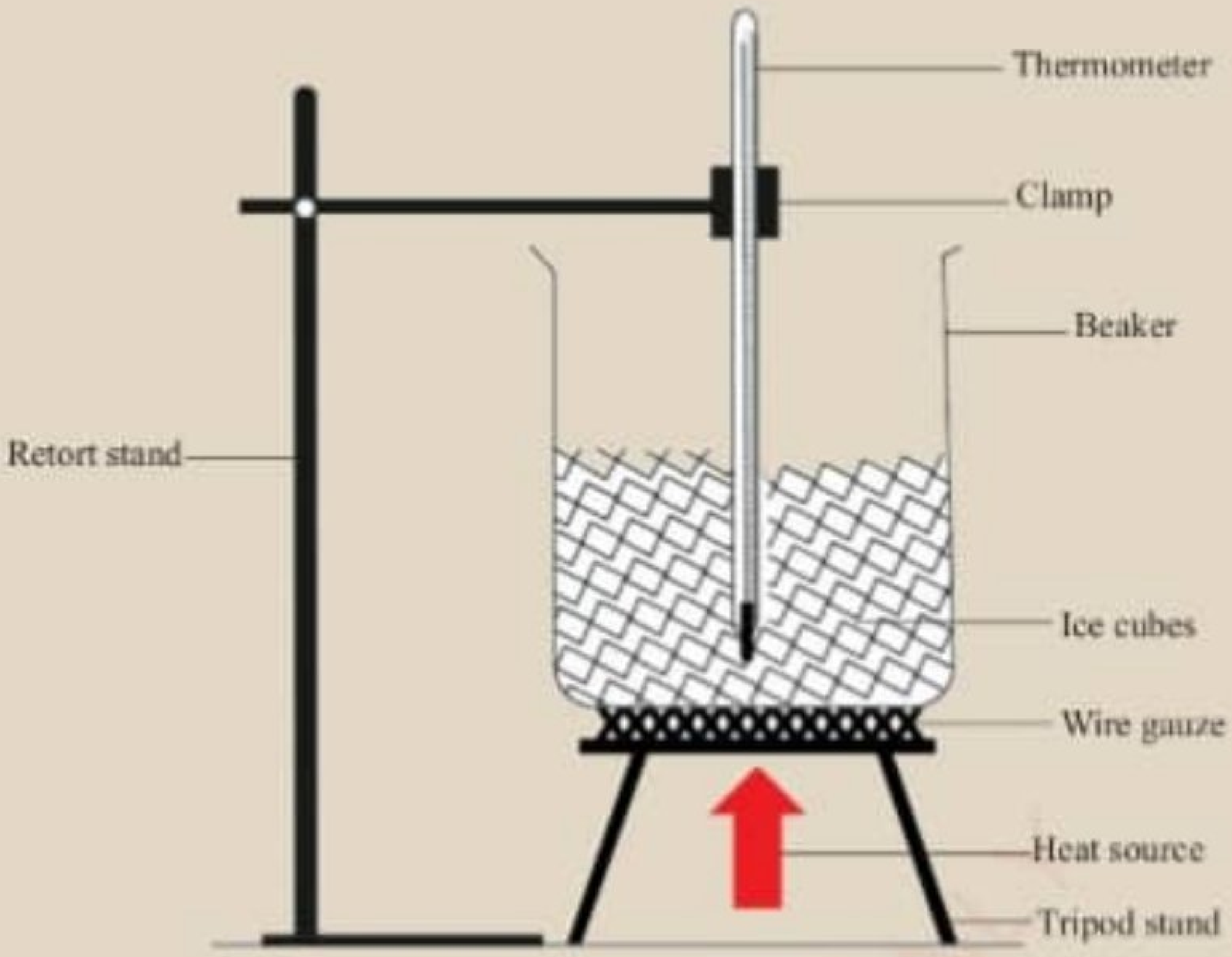


Figure 3.6: *Measuring temperature changes of water*

4. Heat the ice cubes while carefully stirring them with the glass rod. Record the temperature after every one minute, as shown in Table 3.1.

Table 3.1: *The temperature changes of water*

Time (minutes)	Temperature (°C)
0.0	T_0
1.0	T_1
2.0	T_2
3.0	T_3
4.0	T_4
5.0	T_5

5. Continue heating the ice and recording its temperature until all the ice cubes melt and the resultant liquid water starts to boil.

Questions

1. Why is the thermometer not allowed to touch the beaker?
2. What is the temperature of the ice cubes?
3. What is the temperature when the melting is just complete?
4. What is the temperature of the boiling water?

Chemical properties of water

The following are the chemical properties of water:

1. Pure water is neutral at room temperature; it is neither acidic nor basic.
2. Cold water reacts with some metals to form metal hydroxides and liberate hydrogen gas.
3. Steam can react with some metals to give the respective metal oxides and hydrogen gas.

Other chemicals which can react with water include blue cobalt(II) chloride and white anhydrous copper(II) sulphate. When cobalt(II) chloride paper is exposed to water, it changes from blue to pink. When anhydrous copper(II) sulphate is exposed to water, it dissolves to give a blue solution. The two reactions are also used to test the presence of water in a particular substance.

Activity 3.2

Aim: To demonstrate the chemical tests for water.

Requirements: Distilled water, blue and red litmus papers, cobalt(II) chloride paper, watch glasses, and anhydrous copper(II) sulphate

Procedure

1. Pour some distilled water into a watch glass.
2. Dip a strip of blue litmus paper into the water on the watch glass. Record the observation.
3. Repeat steps 1 and 2 using a red litmus paper and cobalt(II) chloride paper separately.

Blue litmus paper

Red litmus paper

Cobalt(II) chloride paper

4. Put a little anhydrous copper(II) sulphate on a watch glass and add some little distilled water. Record the observations.

Questions

1. What colour changes are observed on the blue litmus paper, red litmus paper, and the cobalt(II) chloride paper?
2. What change is observed on the anhydrous copper(II) sulphate when some distilled water is added?

Exercise 3.1

1. Choose the correct answer for each of the following items:
   (i) Water exists in three forms, which are solid, liquid, and vapour. Which among the following are examples of the liquid form of water?
      (a) Rain, snow and hail
      (b) Dew, rain and ice
      (c) Mist, steam and clouds
      (d) Mist, dew and rain
   (ii) Potable water is the one which is
      (a) good for transportation.
      (b) kept in pots for different uses.
      (c) clean and safe for drinking.
      (d) less contaminated with pollutants.
2. Explain the importance of the following in the water cycle:
   (a) Evaporation
   (b) Condensation

3. Mount Kilimanjaro is covered by a mass of ice that makes it important in different aspects. What could happen if the temperature at the mountain increased beyond its common environmental temperature?

Uses of water

Water is important for our daily uses such as in domestic purposes, transportation, recreation, and economic activities. It is also an important component in bodies of living things.

Water for daily uses

Water is used on day-to-day basis for various domestic purposes, such as drinking, cooking, and cleanliness. The water used for those purposes should be safe and clean. Some uses of water are shown in Figure 3.7.

Washing clothes *In the kitchen*

Figure 3.7: *Domestic uses of water*

Water for transportation

Water bodies like lakes, rivers, and oceans are suitable for transportation of people and goods using vessels such as boats and ships (Figure 3.8).

Figure 3.8: *Transportation in water*

Recreation

Water bodies are used for recreational purposes such as swimming, sport fishing, and in ocean sports such as scuba-diving (Figure 3.9).

Figure 3.9: *Water for recreation*

Water in economic activities

Water is used in various economic activities such as manufacturing industries, agriculture, mining, energy, construction, and fishing.

Manufacturing industries: Manufactured goods include chemicals, food, beverages, textile, and paper, among others. The manufacturing is made possible in the presence of water as a solvent, coolant, source of steam for steam engines, and as a medium for different mixtures.

Agriculture: In agriculture, water is used for irrigation, in animal dips, and for animal drinking.

Mining: Water is used as a solvent in the extraction of certain minerals, and separation of impurities.

Energy: Large water bodies, especially rivers and dams are used to generate electrical energy.

Construction: Water is used for the construction of different structures such as roads and bridges.

Fishing: Oceans, lakes, rivers, dams, artificial ponds, and other water bodies are used for fishing.

Water as a component of bodies of living things

Water is a major component of living cells. Therefore, living things need water for their survival, growth, and reproduction. Water makes about 75 percent of the human body.

Water as a solvent for different substances

Water is an important solvent for different substances including some foods and medicines. Many chemical processes are possible when the reacting substances are in aqueous forms. The food we take can be well assimilated when it is mixed with water. Many substances (solutes) dissolve in water. For this reason, water is called the *universal solvent*. However, there are some solutes that do not dissolve in water to an appreciable amount but dissolve in organic solvents.

Activity 3.3

Aim: To compare the solubility of different substances in water.

Requirements: Distilled water, common salt, sugar, diesel, cooking oil, kerosine, ethanol, liquid soap, chalk powder, egg shell powder, baking powder, ten test tubes, a spatula, and a measuring cylinder

Procedure

1. Put 30 cm^3 of distilled water into each test tube.
2. Add about half spatulaful of solid materials or about 5 cm^3 of the liquid materials into different test tubes, shake gently, one at a time.
3. Observe and note down what happens in each of the test tubes.

Questions

1. Which substances dissolve in water?
2. Which substances do not dissolve in water?
3. What name is given to the resulting solution after dissolving a substance in water?

Exercise 3.2

1. (a) Explain the importance of water.
   (b) Relate the different uses of water with its properties.
2. What would happen to living things if there were no water?
3. (a) Water is said to be a universal solvent. What does this mean?
   (b) Explain the importance of water as a solvent.

Water treatment and purification

Most of the Earth's water is not pure. It contains various impurities, and so requires treatment and purification before it can become clean and safe for use.

Water treatment is the process of making water usable for domestic, industrial, medical, and other purposes. The aim of the treatment process is to remove existing contaminants from water, thus improving it for safe uses. The treatment processes may be *physical* such as settling, *chemical* such as addition of chemicals for disinfection, or *biological* such as slow sand filtration.

Water purification is the removal of contaminants from treated water to produce drinking water that is suitable for human consumption. Substances that are removed include bacteria, algae, fungi, minerals such as iron and sulphur, and domestic and human-made chemical pollutants. It should be noted that most water treatment processes also include the purification process.

Domestic water purification

There are very few sources of safe drinking water, thus the treatment of water is necessary. A number of simple and diverse methods of treatment of water for consumption are available. These methods include boiling, use of artificial and natural purifiers, and the use of commercial filters.

Boiling

This is more or less the simplest way to treat water. Water is heated and let to boil before heating can be stopped. Boiling helps to kill disease-causing organisms such as bacteria. Boiling of water should be done in a clean and safe environment. For example, boiling of water in a clean pan (Figure 3.10). The boiled water is then allowed to cool before being filtered using a clean cloth.

Figure 3.10: *Boiling of water*

Use of purifiers

Chemical purifiers are usually in liquid or tablet forms. Examples of the chemical purifiers are sodium hypochlorite, ozone, chlorine, and chlorine dioxide. A recommended amount of the purifier is put in a specific amount of water in a container. The water is shaken or stirred well then left to settle for some time before it can be safe for drinking.

Use of commercial filters

Commercial filters work by allowing water to pass through materials such as activated charcoal or ceramic element that purify water. The entire filter unit is usually defined in terms of two components: the filter element (or media) through which the water passes and the filter system which houses the element. Examples of such water filters include ceramic, activated charcoal, and sand water filters (Figure 3.11). Simple filters can have layers of gravels, sand, activated charcoal, and clean cloth. The gravels trap all visible floating substances; the sand filters the smaller suspended particles; the charcoal kills some of the harmful bacteria; and the clean cloth filters the tiniest particles. Filtered water is therefore clearer, cleaner and safer to use than unfiltered water.

Figure 3.11: *Examples of ceramic water filters*

Activity 3.4

Aim: To assemble a small water filter.

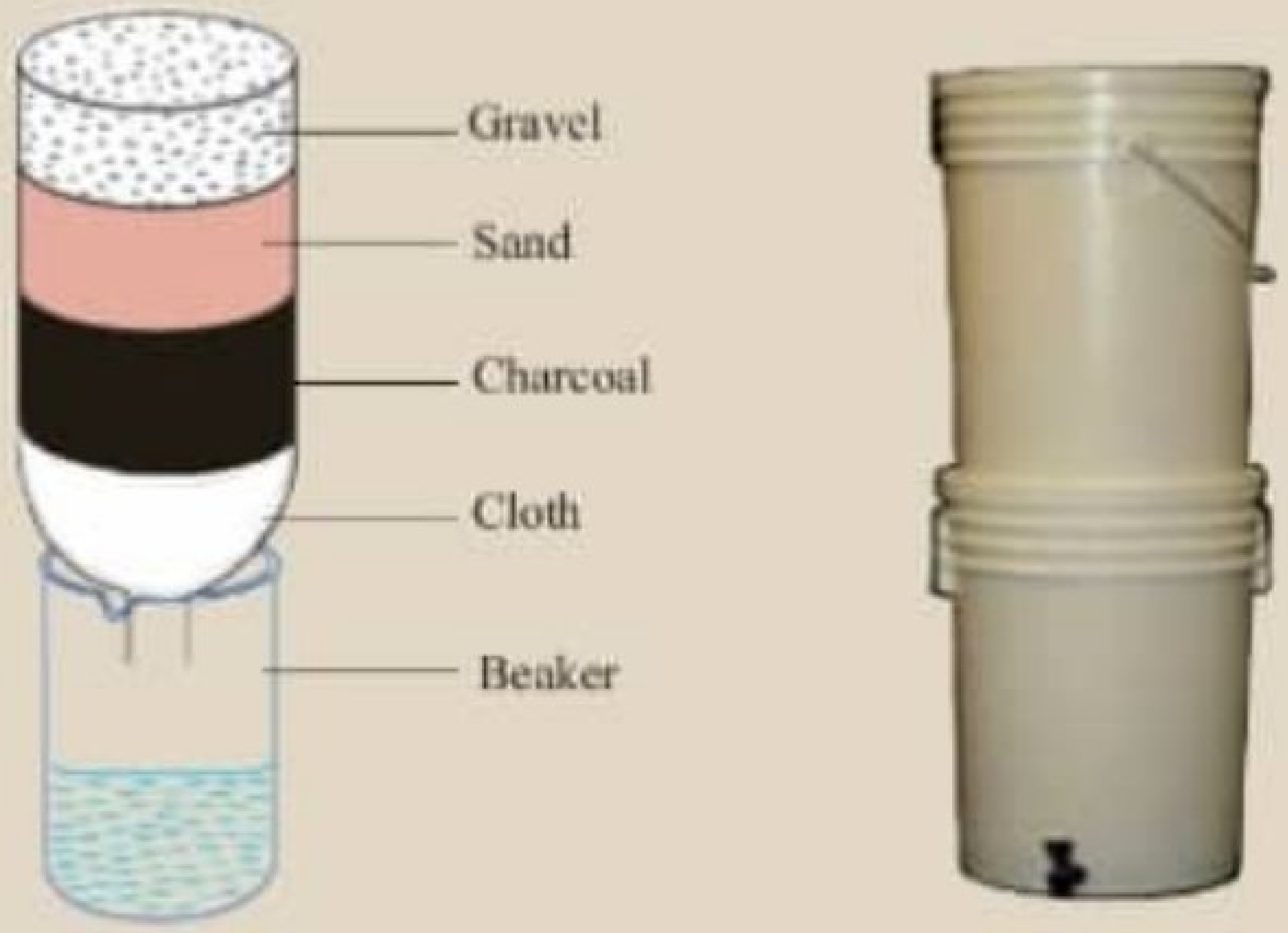
Materials: Piece of clean cloth, sand, activated charcoal, gravel, plastic bottle, beaker, and muddy water

Note: Do not use sooty charcoal.

Procedure

1. Cut off the bottom part of the plastic bottle.
2. Invert the cut bottle and insert the clean piece of cloth at the bottom end, as illustrated in Figure 3.12 (a).
3. Place some charcoal on top of the cloth, followed by some sand, then some gravel.

4. Place the inverted bottle into the beaker.
5. Pour some of the muddy water on top of the sand and leave it for some time. Record your observation.
6. Repeat steps 1 to 6 at home using buckets as shown in Figure 3.12 (b).



(a) A sketch of a filter *(b) A simple filter*

Figure 3.12: *Making a simple filter*

Questions

1. Is the filtered water different from unfiltered water? Give reasons.
2. What roles do the cloth, sand, gravel, and charcoal play?

Urban water treatment

Most of the water used in many urban areas is piped tap water. This water is usually obtained from sources such as rivers, streams, and lakes, but it goes through various processes before it can be safe for consumption. These processes are summarised in Figure 3.13.

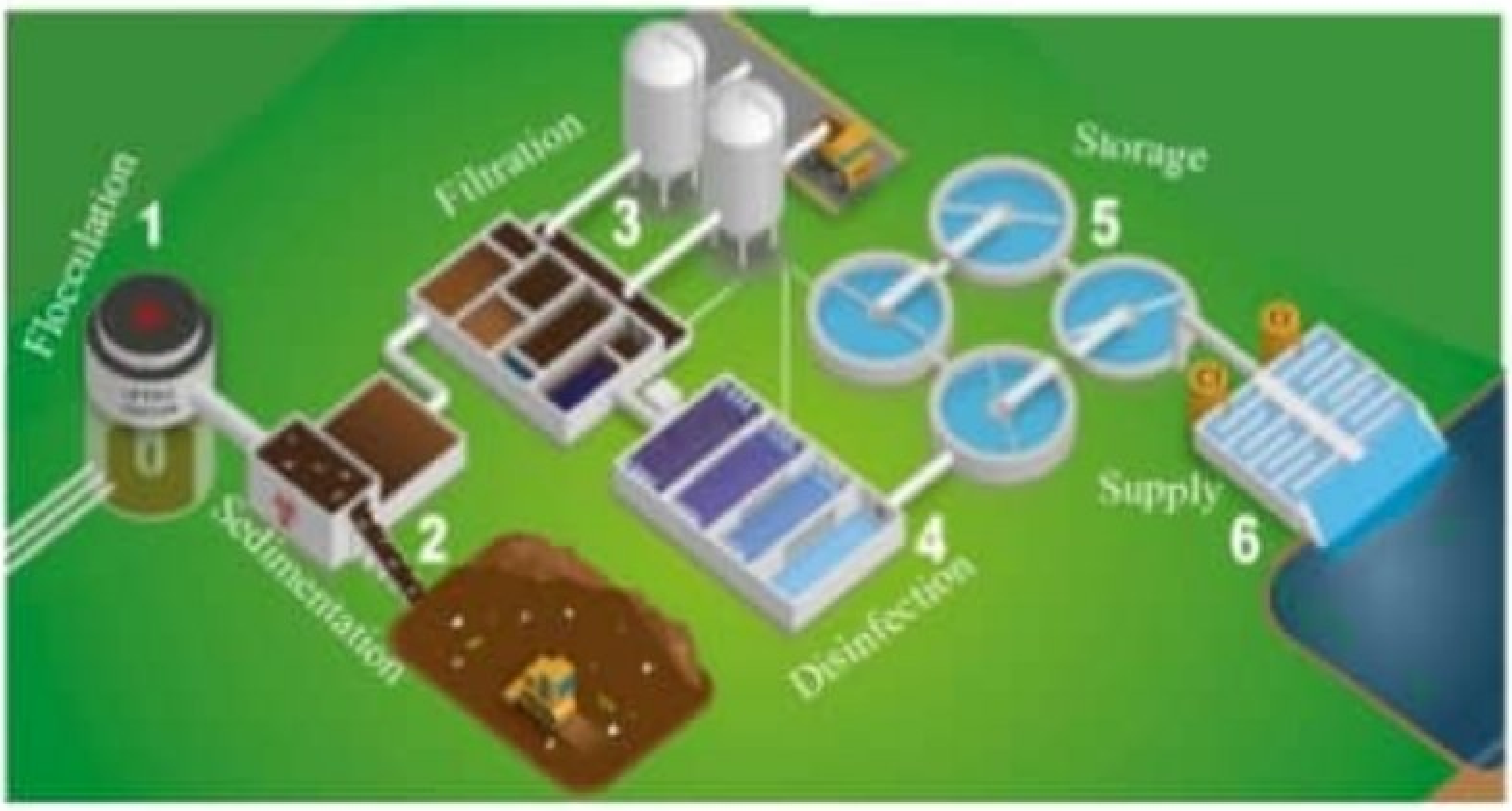


Figure 3.13: *Processes of water treatment*

The first stage of water treatment involves *coagulation* and *flocculation* (1). At this stage, chemicals such as aluminium sulphate, iron(III) sulphate, or sodium aluminate are added in water to bind together small particles present in water and form large particles called *flocs*. The second stage involves settling of flocs (large solid clusters) to the bottom of the reservoir by gravity due to its weight. This process is called *sedimentation* (2). The third stage is *filtration* (3) whereby the clear water on top of the flocs passes through the filter in order to remove the very small and dissolved particles such as bacteria, other organisms, dust, and some of the chemicals. The fourth stage is *disinfection* (4) which involves treatment of filtered water using disinfectants such as chlorine, sodium hypochlorite, or ozone. These chemicals kill any remaining microorganisms such as bacteria, hence protect the water from germs when is piped/pumped for consumption. After disinfection, water is *stored in tanks* (5) ready for *supplies* (6).

Activity 3.5

Site visitation

1. Visit a nearby water treatment plant.
2. Ask as many questions as possible to the plant specialist about what takes place in the water treatment process. Observe the processes that take place. Note the chemicals used.
3. Write a report on the visit.

4. Present the visit report before the class.
5. Discuss the findings of your visit.

Questions

1. Compare the treatment done at large scale with the small water treatment done at home.
2. What chemicals were used at different stages of the water treatment plant you visited?

Importance of water treatment and purification

Water treatment and purification are important due to the following reasons:

1. Water that has not been treated may contain harmful bacteria and other microorganisms that can cause different diseases such as diarrhoea, typhoid, cholera, and other illnesses. Untreated water will usually lead to usage of large amounts of detergents such as soaps for cleaning.
2. Treated water is the best for use in laboratories and medical facilities to ensure accurate results from experiments and effective medical treatments.
3. Treated water is suitable for use in factories to ensure the manufactured products are safe for consumption.
4. Treated water is more efficient to use for cleaning in industries and in domestic settings.
5. Treated water reduces corrosion of different containers and instruments.

Project

1. Carry out a small project to find out the various diseases and illnesses caused by using untreated water. In your environment, make use of a variety of scientific books and articles to establish some facts.
2. Make a chart to illustrate the information you obtained from your investigation. Ensure that the chart is informative, yet simple and clear for easy understanding.
3. Present your findings to the rest of the class.

Chapter summary

1. Water occurs on the Earth in three main states: solid, liquid and gas.
2. Water is tasteless, odourless and colourless.
3. Water cycle is the sequence that describes continuous movements of water on, above and below the surface of the Earth in different states.
4. The water cycle has four main stages: evaporation, condensation, precipitation, and collection.
5. Water treatment is the process of making water safe for use or disposal.
6. Water purification is the removal of contaminants from water to produce clean and safe water for drinking.
7. It is essential to treat water in order to reduce the occurrence of diseases or illnesses.

Revision exercise 3

1. Choose the correct answer for each of the following:

(i) Identify the process that involves conversion of water from vapour to liquid.
(a) Condensation
(b) Precipitation
(c) Evaporation
(d) Transpiration

(ii) Which among the following is the simplest way for water purification?
(a) Cooling
(b) Filtering
(c) Boiling
(d) Condensing

(iii) Fresh water constitutes about _______ percent of the total water on Earth.
(a) 87 (b) 97 (c) 3 (d). 12

(iv) What is the pH state of water at room temperature?

- (a) Basic
- (b) Acidic
- (c) Sour
- (d) Neutral

(v) Water can pass from the atmosphere to land and back into the atmosphere. What term represents such a sequence?

- (a) Water purification
- (b) Precipitation
- (c) Water cycle
- (d) Evaporation

(vi) Identify the process used to remove contaminants from water.

- (a) Water purification
- (b) Sedimentation
- (c) Electrolysis
- (d) Contamination

(vii) What is the form of water when precipitation occurs?

- (a) Rain and clouds
- (b) Rain, hail and snow
- (c) Water ice
- (d) Snow and hail

(viii) Which of the following statements about water is incorrect?

- (a) Water is used in industry as a coolant.
- (b) Water is used as a raw material to produce hydrogen.
- (c) Water is used as a solvent.
- (d) Water is used as a raw material to produce carbon.

(ix) What is the objective of water treatment?

- (a) To minimize water diseases
- (b) To remove unwanted materials

(c) To remove mud

(d) To remove dissolved substances

2. A form two student wanted to test the presence of water in an unknown compound using hydrated copper(II) sulphate. A small amount of hydrated copper(II) sulphate was placed on a watch glass followed by addition of few drops of the unknown compound. There was no change in the colour observed.

   (a) Why was there no change in the colour of the hydrated copper(II) sulphate?

   (b) Name two substances that could be used in place of the hydrated copper(II) sulphate to observe the required colour change.

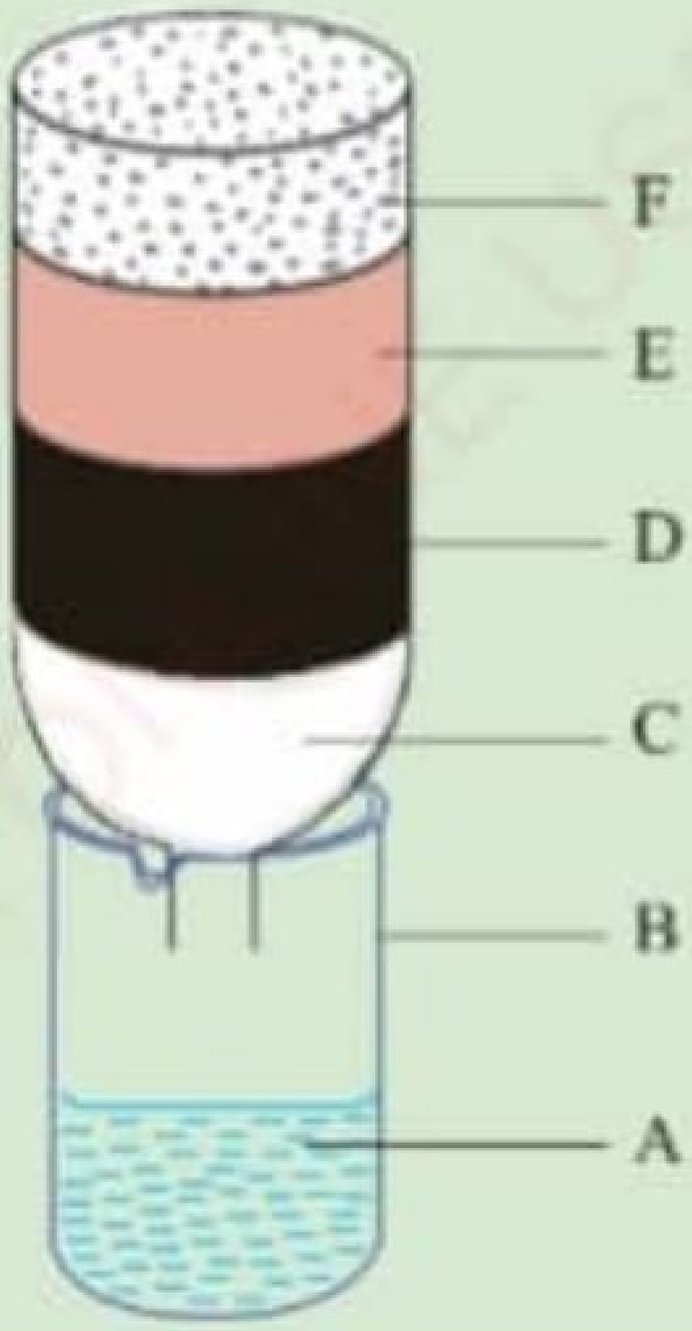
3. Describe the physical properties of water.

4. Why is water important in our daily life activities and in industries? Give at least five reasons.

5. Explain three ways by which water can be purified at home.

6. Explain two chemicals which are added in various stages during large-scale water treatment.

7. The following diagram represents a simple water filter:



(a) Name the parts labelled A to F.

(b) What is the importance of each part?

(c) What would be the disadvantages of using such a filter to obtain drinking water?

8. With the aid of a diagram, explain the processes that take place in the water cycle.

9. Water is important in different economic activities. Comment on this statement.

10. Differentiate water purification from water treatment.

11. Water is composed of hydrogen and oxygen atoms, thus, it could be used for the preparation of oxygen. It is also available in larger amount than potassium chlorate and hydrogen peroxide, which are the chemicals used to prepare oxygen in the laboratory. Why is it not used for the preparation of oxygen in the laboratory?

Chapter Four

Fuels and energy

Introduction

A fuel is any combustible substance which on burning in air gives a large amount of heat energy, that can be used economically for domestic, transportation and industrial purposes as well as other uses. Since combustion is a chemical process, fuels are also called chemical fuels. In this chapter, you will learn about categories of fuels and their characteristics, uses of fuels and the environmental effects of using charcoal and firewood. You will also learn about energy and the alternative sources of energy. The competencies developed will help you to use fuels properly and economically.

Categories of fuels

Fuels can be categorised on the basis of their occurrence and physical states.

Categories of fuels according to their occurrence

On the basis of their occurrence, fuels can be classified into *natural fuels* (or primary fuels) and *artificial fuels* (or secondary fuels).

Natural fuels

Natural fuels occur in nature, that is, they are not manufactured (not man-made). They include wood, coal, petroleum, and natural gas.

Artificial fuels

Artificial fuels are either manufactured in industries or derived from primary fuels through refinery. Artificial fuels include petrol, kerosene, diesel, alcohols, hydrogen, water gas, coal gas, and producer gas.

Categories of fuels according to their physical states

On the basis of their physical states, fuels can be classified as *solid fuels*, *liquid fuels*, and *gaseous fuels*. Table 4.1 shows different physical states of fuels with their examples.

Table 4.1: *Categorisation of fuels according to the physical states*

Physical state	Primary/ natural	Secondary/ artificial
Solid	Wood and coal	Charcoal/coke
Liquid	Crude petroleum	Kerosene, petrol, diesel, and biodiesel
Gaseous	Natural gas	Liquefied petroleum gas (LPG), coal gas, water gas, producer gas, hydrogen, and alcohols.

Characteristics of a good fuel

Fuels can be classified according to their effectiveness (usefulness) or productivity and convenience for use. The following are the characteristics that are considered when choosing a good fuel:

Energy value - A good fuel should have high energy value (calorific value). The energy value of a fuel is determined by the amount of energy produced per unit mass of the fuel. This is called the *heat value* or *calorific value* of the fuel.

Velocity of combustion - This refers to the rate at which a fuel burns. A good fuel should burn with moderate velocity for continuous supply of heat. It should not burn too fast or too slowly.

Ignition point - This is the temperature to which the fuel must be heated before it starts burning. A good fuel should have a proper (average) ignition point. A low ignition point is risky due to fire hazards, while high ignition point makes it difficult to start a fire with the fuel. Fuels with high ignition points are safe for transportation and storage.

Non-combustible material content - A good fuel should have no or low content of non-combustible materials. The non-combustible material is left in form of ash once the fuel burns. A high content of non-combustible materials lowers the heat value of the fuel. Figure 4.1 shows ashes from burnt substances. High contents of ashes per fuel burned indicate that the burnt substances are not good fuels.

Figure 4.1: *Ashes*

Non–hazardous products of combustion- A good fuel should give clean gases during combustion. The fuel should also give off very little or no smoke. In general, the combustion of a good fuel should not produce harmful substances like soot, and toxic substances. Figure 4.2 shows smoke from a chimney. This indicates that the burning fuel is not good or incomplete combustion takes place.

Figure 4.2: *Smoke of a burning fuel emitted from a chimney*

Pyrometric burning effect - This is the highest temperature that can be reached by the burning fuel. A good fuel should have high pyrometric effect. Burning gaseous fuels produce the highest pyrometric effect. Figure 4.3 shows a burning gaseous fuel.

Figure 4.3: *Burning gaseous fuel*

Availability - A good fuel should be readily available in large quantities.

Affordability - A good fuel should be cheap and affordable.

Ease of transportation and storage - A good fuel should be easy and safe to transport, handle, and store. Figure 4.4 shows transportation of a gaseous fuel using a truck.

Figure 4.4: *Transportation of liquefied petroleum gas (LPG)*

Effects in the environment - A good fuel should not pollute the environment during its production, storage, and use. Fossil fuels, which produce carbon dioxide on burning, are major contributors to environmental pollution (Figure 4.5). Solid fuels like wood and coal are not good due to the following reasons:

(i) They produce harmful gases when burnt.

(ii) They leave solid residues.

(iii) The resulting ashes can cause health problems.

Liquid fuels, like petrol, kerosene and diesel, burn more smoothly than solid ones. However, upon incomplete burning, they also produce poisonous gases and soot.

Figure 4.5: *Smoke emitted from burning fossil fuels*

Charcoal

Charcoal is made by the dry distillation of wood. The dry distillation of wood is done at a temperature between 400 °C and 450 °C in an earth-pit kiln or earth-mound kiln. In the earth-pit kiln, wood is heaped in a hemispherical pile in a central pit. It is then covered with soil or pieces of turf (sod) leaving only a few small air holes near the bottom. The wood is lit at the centre and allowed to burn until the whole pile is on fire. The air inlets are then closed. A smouldering combustion takes place, utilizing the oxygen and hydrogen components of the wood fibre. The products of this combustion are water, carbon dioxide, and volatile organic compounds which escape into the atmosphere. The pit is kept covered until the fire goes off and the charcoal cools. All the volatile matter is driven out in this process. The residue consists of carbon and the inorganic components of the wood. The yield of charcoal is only 20% by weight and 75% by volume of wood.

The earth-mound kiln works in the same way as the earth-pit kiln. However, instead of a pit, the wood is heaped in a pile above the ground surface (Figure 4.6). The earth-mound kiln is preferred where the soil is rocky or the water table is close to the surface.

Figure 4.6: *Earth-mound kiln*

Good charcoal (Figure 4.7) is porous, and brittle. It burns with a non-luminous flame and is easily ignited.

Figure 4.7: *Charcoal*

Coal

Coal is the most important solid fuel. It is a fossil fuel formed by the anaerobic (without oxygen) decay of plants that lived millions of years ago. The energy found in coal originates from the sun and it is stored in plants when photosynthesis takes place. This energy remains in the coal after the decay process of plants.

Composition of coal

Coal contains mostly carbon, but it also has hydrogen, oxygen, sulphur, nitrogen as well as some inorganic components (minerals), and water (moisture). The physical properties of coal include moisture, volatile matter, ash and fixed carbon (coke). These properties are expressed in percentages. Moisture is water or other liquids diffused in small quantities as vapour within a solid or condensed surface of coal. Volatile matter is the material that is driven off when coal is heated to about 950 °C in the absence of air. It consists of gases and low-boiling point organic compounds that condense into oils and tar when cooled. Ash is the non-combustible residue material left after coal is burnt. Coke is the material left after the volatile matters are driven off.

Types of coal

There are different types of coal that vary in composition and properties. The most important types of coal are lignite, sub-bituminous, bituminous and anthracite (hard) coal. Lignite is a soft brown coal, sub-bituminous is a hard lustrous dark brown coal, bituminous is a black and shiny coal, while anthracite is a hard coal with metallic lustre. The common types of coal are shown in Figure 4.8.


Lignite *Sub-bituminous coal* *Bituminous coal* *Anthracite*

Figure 4.8: *Major types of coal*

Destructive distillation

Destructive distillation is a process through which organic fuels such as wood, coal, and oil shale are decomposed by heating in the absence of air (oxygen) to obtain useful products such as coke, charcoal, oils and gases. Figure 4.9 shows a set-up on how destructive distillation can be conducted in the laboratory.

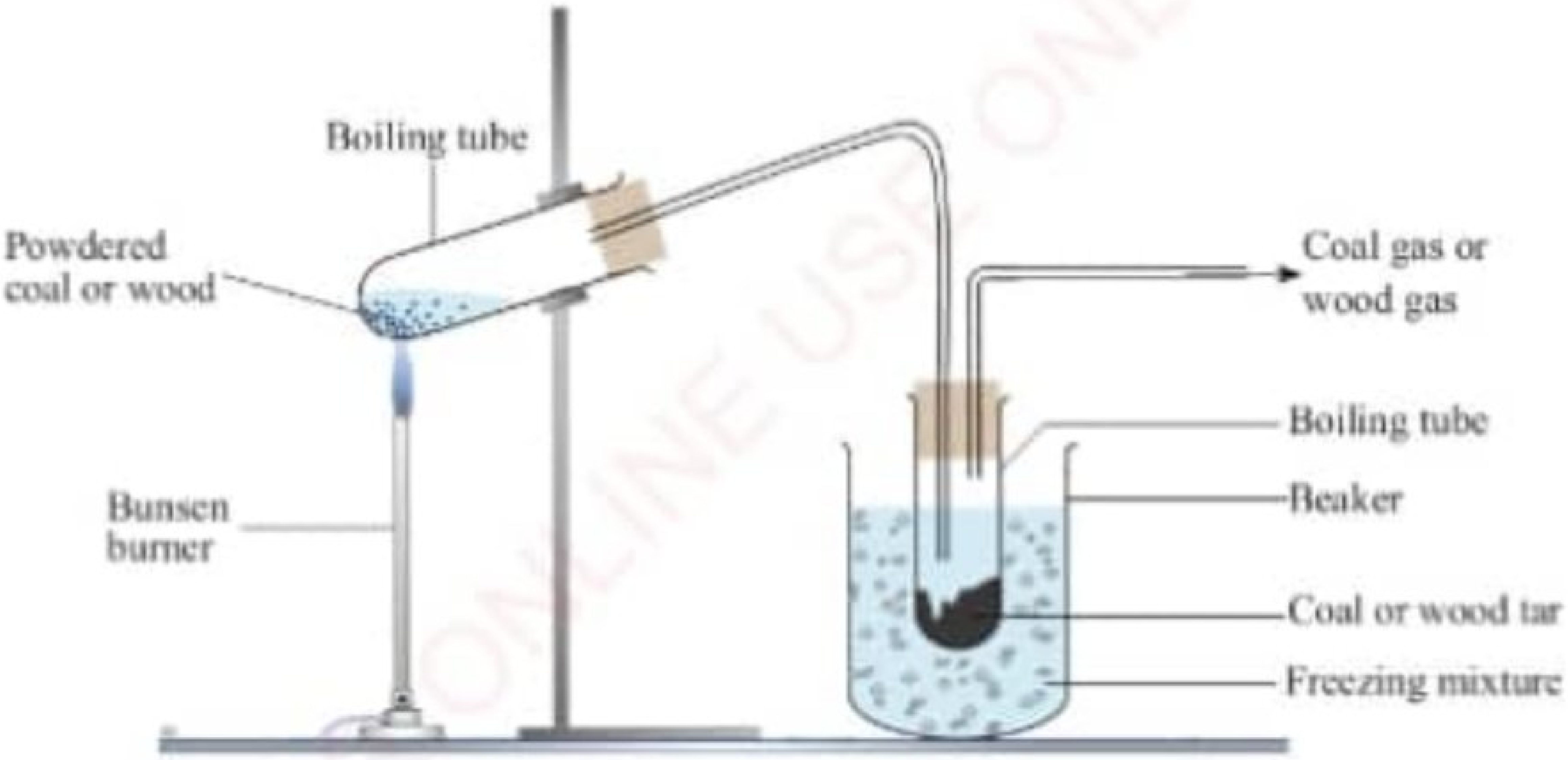


Figure 4.9: *Set-up on destructive distillation*

Destructive distillation of wood

In the destructive distillation of wood (Figure 4.10), wood gas is given off at the side tube. A mixture of water, methyl alcohol and acetic acid is collected in the boiling tube immersed in the freezing mixture. Wood tar collects at the bottom of the boiling tube.

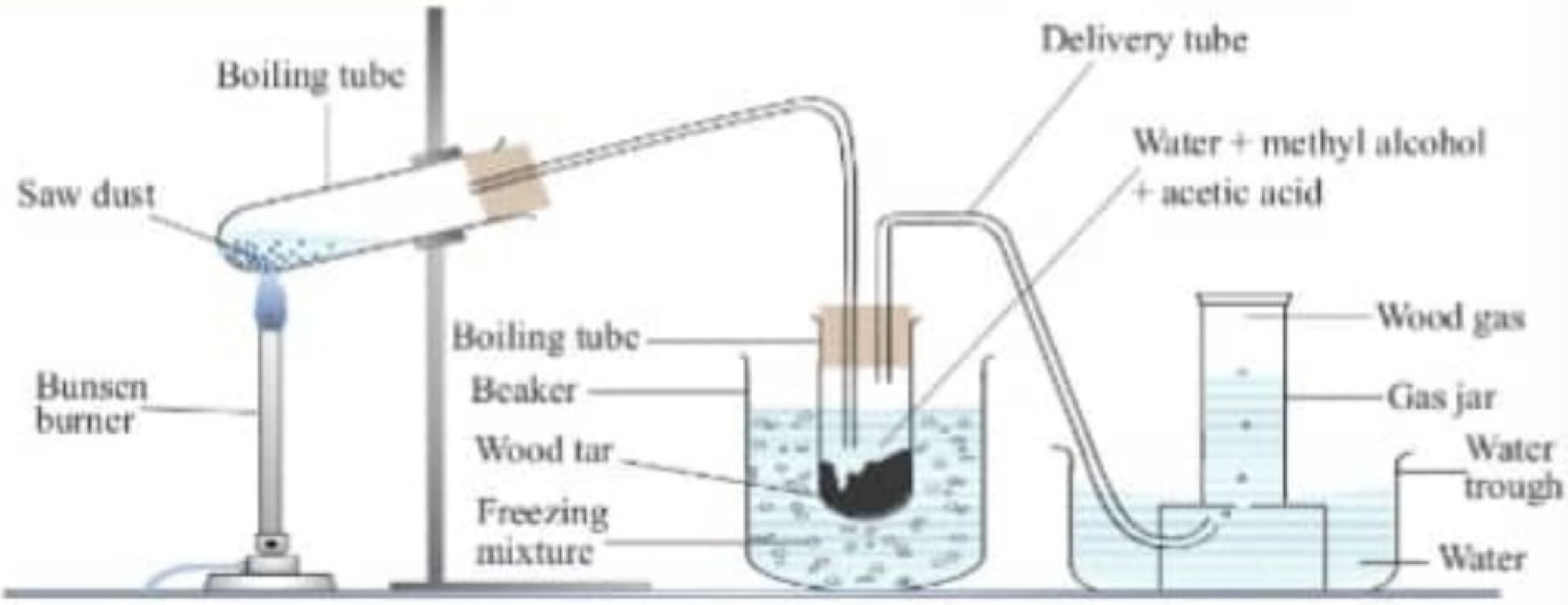


Figure 4.10: *Destructive distillation of wood*

Destructive distillation of coal

The main aim of destructive distillation of coal is to get rid of the volatile matter. Coal that contains a large amount of volatile matter burns with a smoky flame and has low energy value. When coal is heated strongly in the absence of air, destructive distillation occurs. The coal breaks down to give coal gas, ammonia, coal tar, and coke. The tar is collected in the boiling tube which contains water, while the coal gas escapes through the side tube. The residue left in the boiling tube is coke, which is nearly pure carbon. Coke is the most widely used coal product. It is mainly used in metal-extraction furnaces. Ammonia dissolves in the water forming a solution of ammonium hydroxide known as ammoniacal liquor. Figure 4.11 shows the destructive distillation of coal.

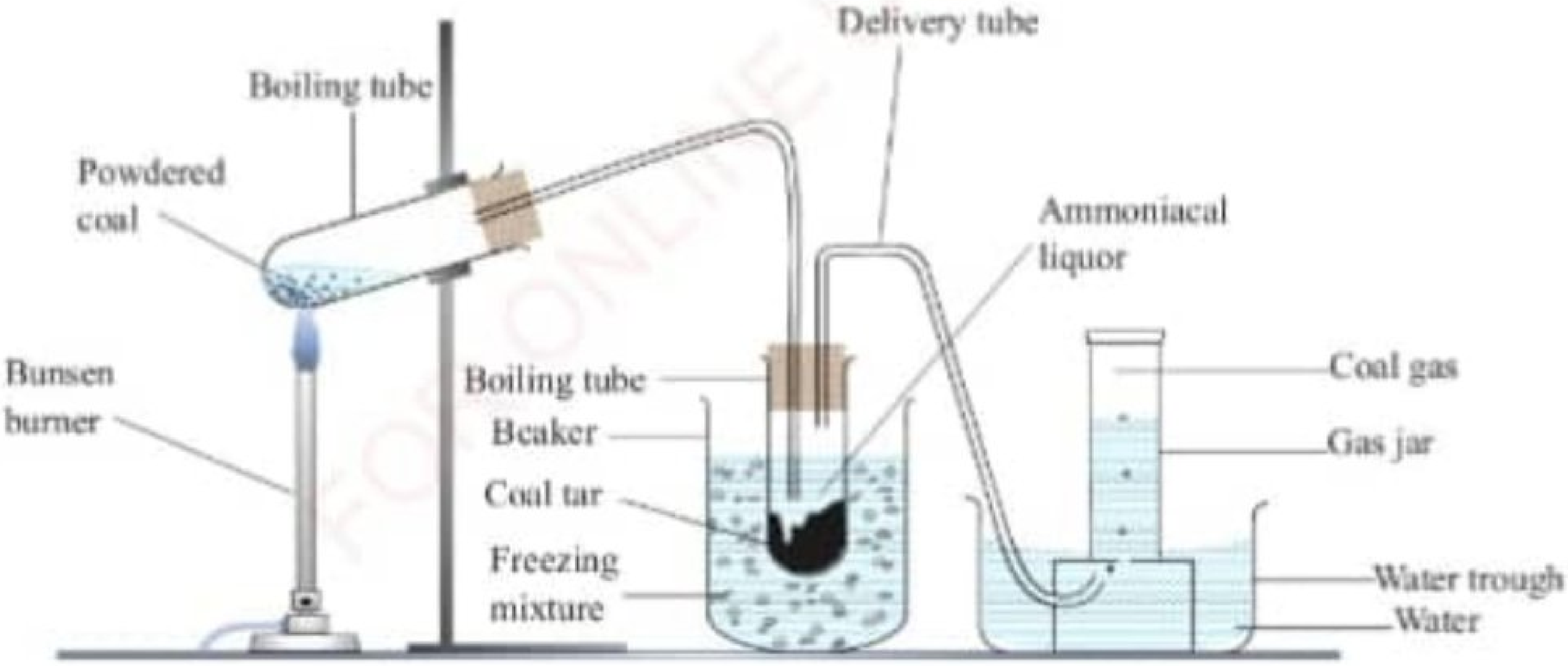


Figure 4.11: *Destructive distillation of coal*

Destructive distillation of coal is generally carried out in two types of kilns. These are the *beehive kiln* and the *Otto Hoffmann kiln*.

Beehive kiln

This is the earliest and the cheapest process of distilling coal. The kiln is a dome-shaped structure made up of bricks. It has two openings, one at the top for charging (adding) the coal, and the other on the side to discharge (remove) coke. A side door is also used for supplying air to ignite the coal. A uniform layer of coal is spread over the hearth (base) through the charging door. Air is supplied through the side door to ignite the coal. The volatile matter escapes and burns inside the partially closed side door. When the distillation is complete (which takes 3 to 4 days), the hot coke is quenched (cooled) with water and taken out through the side door. This process yields about 60% coke by mass. Figure 4.12 shows a beehive coke kiln.

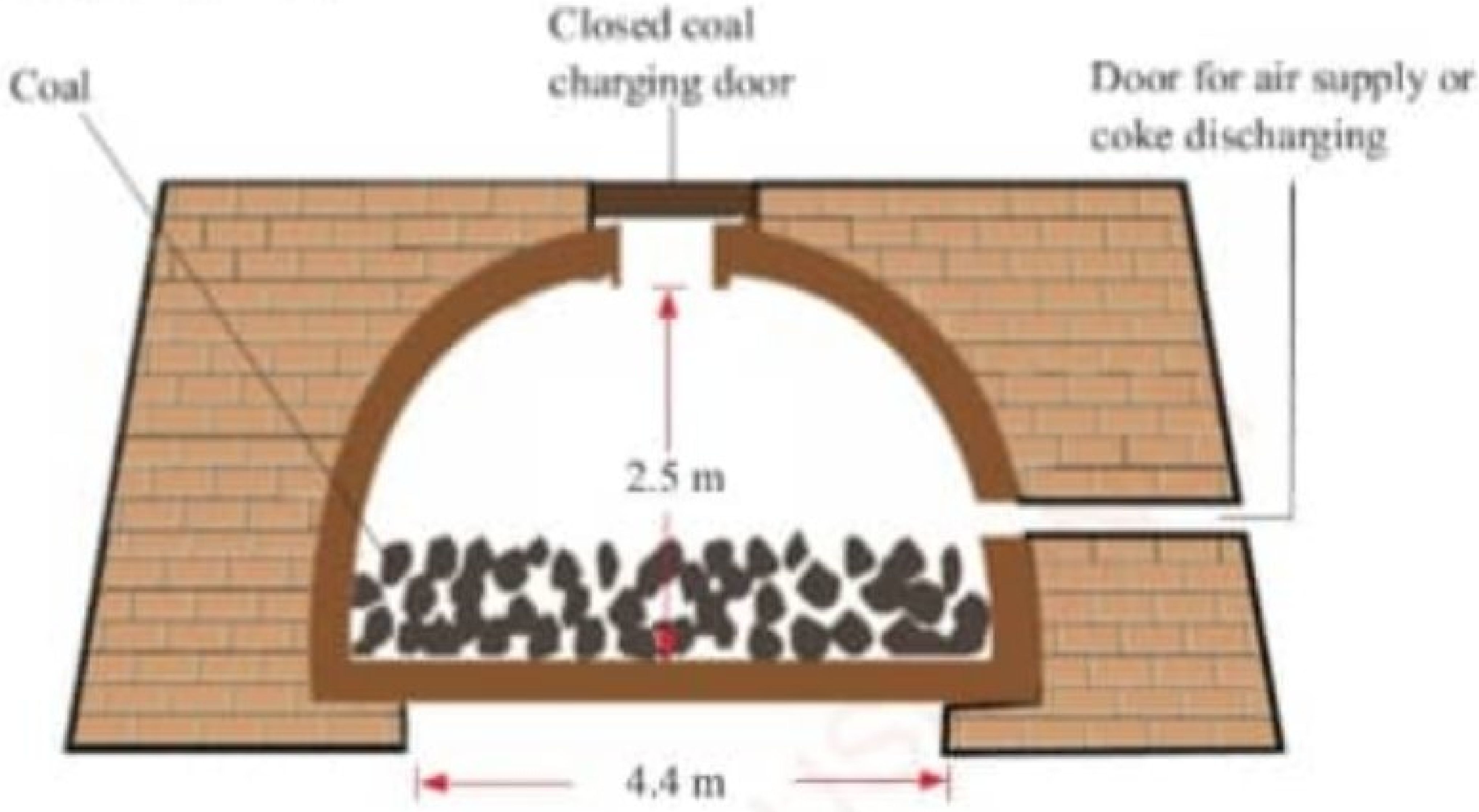


Figure 4.12: *Beehive coke kiln*

Otto Hoffmann kiln

The *Otto Hoffmann* kiln consists of a number of narrow silica chambers separated by spaces for burning gas. Each chamber has a charging hole at the top, a gas outlet and doors at each end for discharging coke as shown in Figure 4.13.

Coal is added into the chambers, then the chambers are closed. The coal is heated to drive out liquid or gaseous components in the materials. This is called *dry distillation*. Heating is done externally by a part of coal gas produced during the process, or by producer gas or by blast furnace gas. The heating is continued until the evolution of volatile matter stops, which may take about 24 hours. The coke that is formed is then pushed out and quenched using water spray. This is called *wet quenching*. In *dry quenching*, the red-hot coke is cooled using an inert gas like nitrogen. Dry quenching produces strong, dense, clean and non-reactive coke. The yield of coke from the Otto Hoffmann kiln is about 75% of coal by mass. Otto Hoffmann kiln has advantages over the beehive kiln. This is because the by-products of the distillation process, for example, ammonia, coal gas, benzol oil, and tar are also recovered. Figure 4.13 shows the Otto Hoffmann kiln.

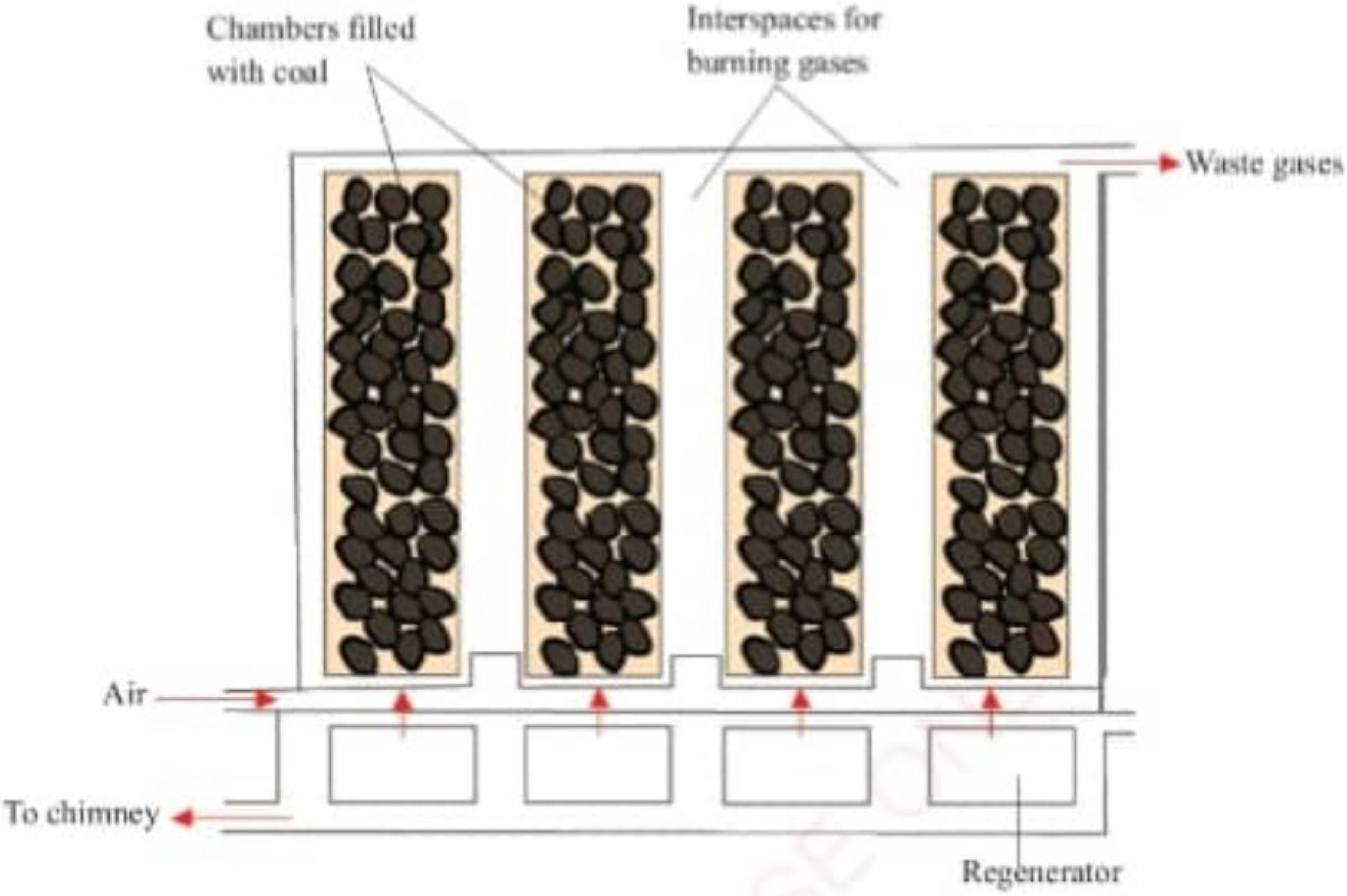


Figure 4.13: *Otto Hoffmann kiln*

Activity 4.1

Aim: To demonstrate the destructive distillation of wood and coal.

Requirements: Powdered coal, Bunsen burner, sawdust, crushed ice (freezing mixture), retort stand and clamp, boiling tubes, 2-hole stopper, test tube holder, delivery tubes, beaker, and wooden splints

Procedure

1. Put some powdered sawdust into the boiling tube.
2. Set the apparatus as shown in Figure 4.9.

3. Heat the sawdust slowly for about 20 minutes.
4. After about 10 minutes of heating, bring a burning wooden splint to the tip of the narrow tube. Record your observations.
5. Stop the heating and dismantle the set-up when there is no more observable change taking place in the boiling tube. Record all your observations.
6. Repeat the procedure (steps 1 to 5) using powdered coal in place of sawdust.

Questions

Referring to the two sets of the experiment;

1. What did you observe when you placed a burning wooden splint at the opening of the tube in step 4?
2. Describe the appearance of the mixture that:
   (a) remains in the boiling tube after heating.
   (b) forms in the boiling tube immersed in the freezing mixture.

Gaseous fuels

The most important gaseous fuels used in industries are natural gas, producer gas, and water gas.

Natural gas

Natural gas consists mainly of methane (about 95% of the total volume). Other components in natural gas are ethane, propane, pentane, nitrogen, carbon dioxide, and traces of other gases. Very small amounts of sulphur compounds are also present. Since methane is the largest component of natural gas, generally, properties of methane are used when comparing the properties of natural gas to other fuels. Natural gas has high calorific value. It mixes with air readily and does not produce smoke or soot. Purified natural gas contains no sulphur. It is lighter than air and disperses into air easily in case of leak.

Producer gas

Producer gas is a mixture of carbon monoxide and nitrogen. It is produced by burning a solid carbonaceous fuel, for example coke, in a limited supply of air. Carbonaceous fuels are fuels that contain a high proportion of carbon.

Production of producer gas

Producer gas is manufactured in a producer furnace (Figure 4.14). The furnace consists of a large air tight cylindrical vessel made of mild steel. The vessel is lined on the inside with fire bricks. At the bottom, there is a pipe for blowing in air and an opening for removing ash. Coal is added through a hopper at the top and the producer gas comes out through an exit near the top.

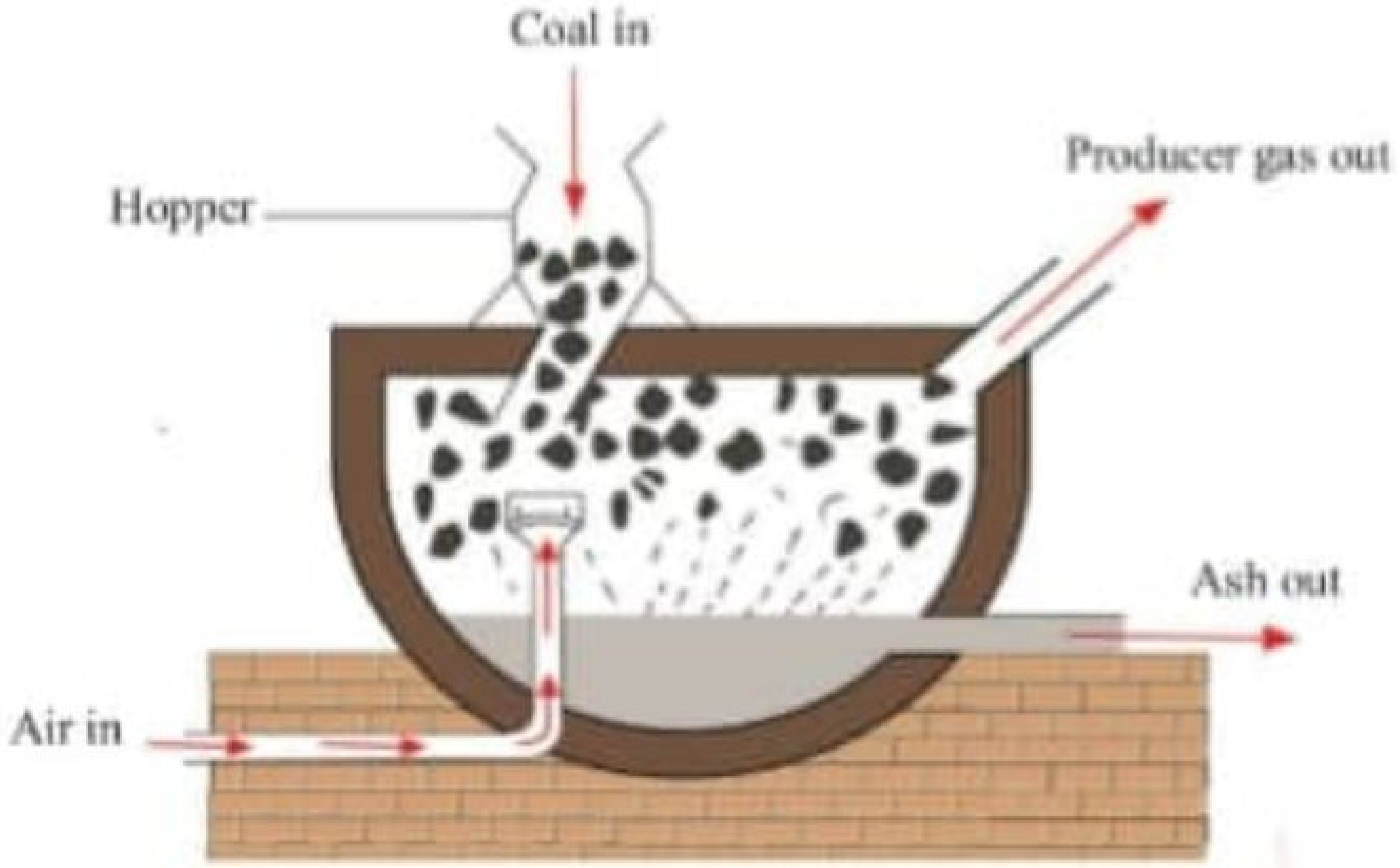


Figure 4.14: *Producer furnace*

When air, mixed with a little steam, is passed through the inlet, the carbon (from coal) combines with oxygen (from air) in the lower part of the furnace to form carbon dioxide. The carbon dioxide formed rises up through the red-hot coal and gets reduced to carbon monoxide. The nitrogen gas in the air is not affected at all during the process. Thus, a mixture of carbon monoxide and nitrogen, with traces of carbon dioxide and some organic compounds, comes out through the exit at the upper end of the furnace. Since more heat (406 kJ mol^{-1}) is produced in the lower part than is absorbed in the upper part (163 kJ mol^{-1}), some excess heat is obtained in the long run. This heat keeps the coal hot.

Properties of producer gas

The properties of producer gas include the following:

1. Producer gas is toxic.
2. It is insoluble in water.
3. It is heavier than air.

Composition of producer gas

The average composition of producer gas is shown in Table 4.2.

Table 4.2: *Composition of producer gas*

Gas	*Composition (%)*
Nitrogen	52–55
Carbon monoxide	22–30
Hydrogen	8–12
Carbon dioxide	3
Methane	Trace amounts

Uses of producer gas

Producer gas is used as a fuel for heating open-hearth furnaces (in steel and glass manufacturing), muffle furnaces and retorts (in the production of coke and coal gas). It also provides a reducing atmosphere in extraction of some metals.

Water gas

Water gas is a mixture of carbon monoxide and hydrogen, with small amounts of nitrogen, carbon dioxide, and methane.

Production of water gas

Water gas is produced in a water gas generator (Figure 4.15) by the action of steam on a bed of coke at 1000 °C. Since the reaction absorbs energy, the coke cools down a few minutes within the process, and the reaction proceeds in a different way to form carbon dioxide and hydrogen instead of water gas. In order to avoid such a reaction, the current of steam is alternated with a blast of air. Hence, carbon reacts with oxygen to yield carbon dioxide, a reaction which gives out energy (80 kJ mol^{-1}). Again, carbon reacts with oxygen to yield carbon monoxide, a reaction which gives out energy (247 kJ mol^{-1}). As a result of these reactions, the temperature of the carbon (coke) rises again. When the temperature reaches 1000 °C, the entry of air is stopped and steam is passed again. Thus, in modern water gas plants, steam and air are blown by alternation.

The period of steam blow (cold blow) is usually 4 minutes, while the period of air blow (hot blow) is very short (1–2 minutes). The duration of these periods is adjusted in such a way that maximum yield of water gas is obtained. During the steam blow, water gas is produced, which is led out through the water gas outlet. On the other hand, carbon dioxide, nitrogen and methane are produced in small amounts during the air blow. These are allowed to escape into the atmosphere.

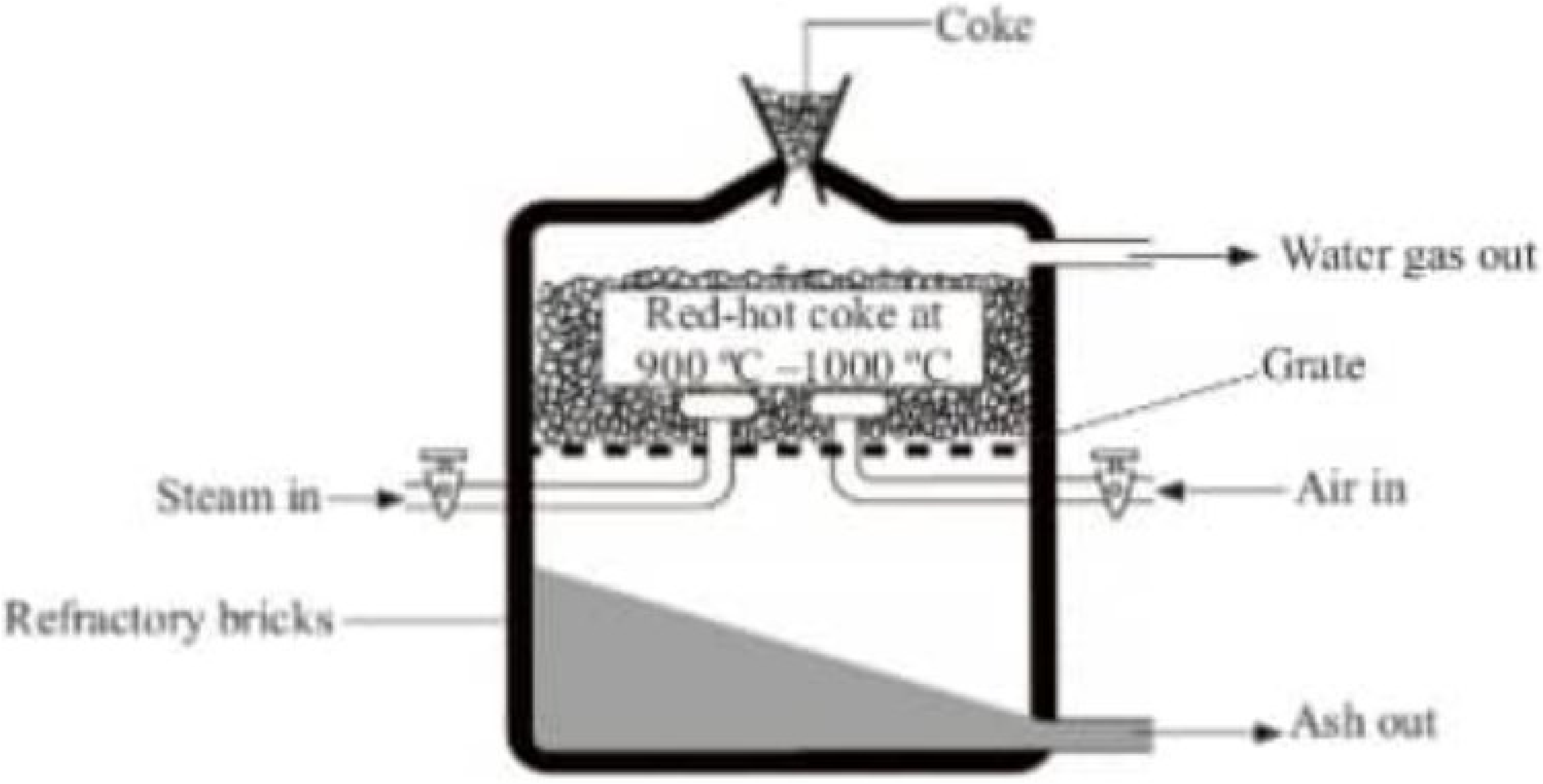


Figure 4.15: *Water gas generator*

Properties of water gas

1. Water gas burns with a non-luminous blue flame, hence is also called *blue water gas*.
2. It has an energy value of about 13628 kJ/mol.
3. It burns with high temperature flame of about 1200 °C.

Composition of water gas

The approximate composition of water gas is shown in Table 4.3.

Table 4.3: *Composition of water gas*

Gas	Composition (%)
Hydrogen	48
Carbon monoxide	44
Carbon dioxide	4.2
Nitrogen	3.0
Methane	0.8

Uses of water gas

Water gas has a high energy value and is therefore widely used as an industrial fuel, especially in the making of steel. It is also used in the preparation of hydrogen.

Uses of fuels

The uses of fuels depend on the types of fuels and their efficiency. Wood, and charcoal are used domestically as heat sources for cooking, boiling, and ironing. Petrol, diesel, and liquefied petroleum gases are used for running industrial plants, cars, planes, ships, and running trucks. Kerosene is used in kerosene stoves for cooking and as a source of light in the kerosene lamps. Coal is used in power plants to generate electricity, it is also used in industries to make dyes, insecticides and fertilisers.

Environmental effects of using charcoal and firewood

Burning of charcoal and firewood causes various effects in the environment. These include production of carbon dioxide which causes *global warming*. Carbon dioxide produced from various processes traps the heat of the sun in the lower atmosphere causing the Earth's average temperature to rise. This is referred to as the *greenhouse effect* that leads to *global warming*.

Making charcoal and firewood involves cutting down trees which leaves the land unprotected against wind blow and water flow that can cause soil erosion. Cutting of trees may also cause drought, which results into food insecurity. Figure 4.16 shows a flooded area and dryland caused by global warming.

Flooded area *Dryland*

Figure 4.16: *Some effects of global warming*

Energy

Energy is the capacity or ability of a body or system to do work. The SI unit for energy is the joule (J). Energy exists in two major forms, namely potential energy and kinetic energy. *Potential energy* is the energy in matter due to its position or state. Examples of potential energy include chemical energy, elastic energy, nuclear energy, and gravitational energy. *Kinetic energy* is the energy possessed by a body due to motion. The motion could be of waves, electrons, atoms, molecules or the object itself. Examples of kinetic energy include electric energy, radiant energy, thermal energy, and sound energy. *Mechanical energy* is the sum of kinetic energy and potential energy. Table 4.4 shows the different forms of kinetic energy and potential energy.

Table 4.4: *Examples of forms of kinetic energy and potential energy*

Kinetic energy	Potential energy
Electrical energy is energy possessed by electrical charges in motion. For example, electricity and lightning.	*Chemical energy* is energy possessed by matter due to its chemical make-up, that is, arrangement of atoms and molecules. For example, biomass, petroleum and natural gas.
Radiant energy is electromagnetic energy that travels in transverse waves. For example, visible light, X-rays, gamma rays, radio waves, and solar energy.	*Elastic potential energy* is energy stored in objects by the application of force. For example, compressed springs and stretched rubber bands.
Thermal or heat energy is the internal energy in substances caused by the vibration and movement of atoms and molecules within the substance. For example, geothermal energy.	*Nuclear energy* is energy possessed by an atom in its nucleus. Nuclear energy holds the nucleus together. The energy is released when nuclei are combined or split apart.
Sound energy is the movement of energy through substances in longitudinal waves. Sound is produced when force causes an object or substance to vibrate.	*Gravitational energy* is the energy possessed by a body due to its position or place. When an object is lifted or suspended in air, it possesses gravitational energy due to its position.

Conservation of energy

The *principle of conservation of energy* states that: *"Energy can neither be created nor destroyed, it can only be transformed from one form to another"*. In practice, appliances only convert energy to different forms but they do not create it. The efficiency of any appliance is always less than 100%. For example, an electric bulb converts electric energy to light and heat energy. The total energy input is always equal to the total energy output, irrespective of the form of energy.

Transformation of energy

The process of changing energy from one form to another is referred to as transformation of energy. Below are five examples of how energy is transformed from one form to another.

Changing mechanical energy to electrical energy

A hydroelectric power plant converts mechanical energy into electrical energy. Figure 4.17 shows the structure of a hydroelectric power plant.

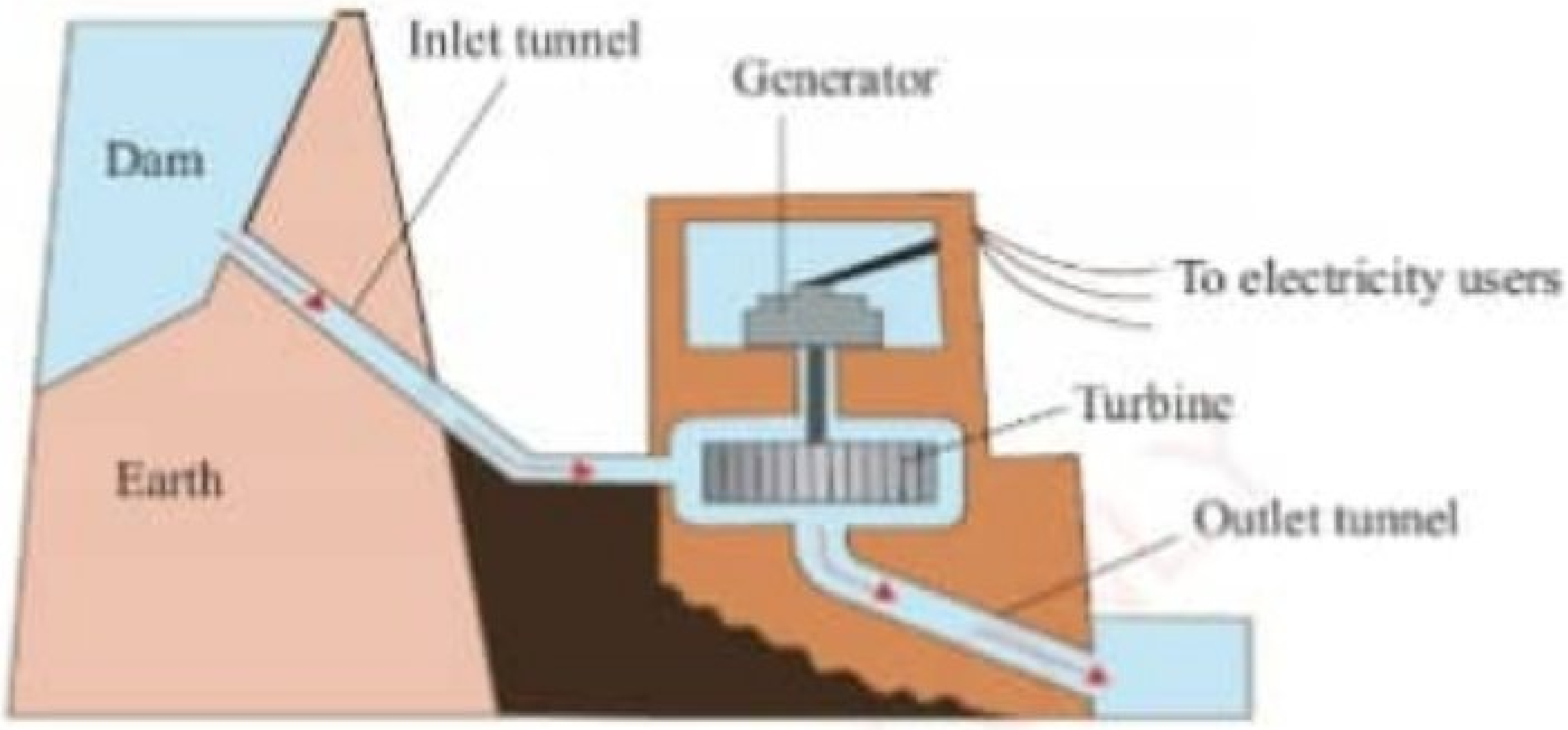


Figure 4.17: *Structure of hydroelectric power plant*

The still water in the dam possesses potential energy. When the water is allowed to flow out, the potential energy is converted to kinetic energy. As the water moves down to the turbine, its potential energy decreases while its kinetic energy increases. By the time the water reaches the turbine, most of the potential energy will have been converted to kinetic energy. When water is used to rotate the turbines to produce electricity, the kinetic energy possessed by the flowing water is converted to mechanical energy and then to electrical energy.

Electrical energy can also be generated from wind. This is done using a windmill. The wind possesses kinetic energy, which can rotate the blades of a windmill. If the windmill is used to rotate a dynamo, electrical energy is produced. The combination of a windmill and a dynamo converts mechanical energy from the wind to electrical energy.

Changing electrical energy to heat energy

Appliances that produce heat when connected to a source of electricity include the electric iron, electric kettle, electric cooker, electric heater and the tungsten bulb (Figure 4.18). These devices make use of high resistance wires. When an electric current passes through such wires, the electrical energy is converted to heat energy.

Insertion heater *Electric cooker* *Electric kettle* *Electric iron* *Electric bulb*

Figure 4.18: *Appliances that convert electrical energy to heat energy*

Change of electrical energy to mechanical energy

When a source of electric current is connected to an electric motor (Figure 4.19), a rotation of the motor occurs. In this way, electrical energy is converted to mechanical energy.

Figure 4.19: *Electric motor*

Conversion of solar energy to other forms of energy

The energy from the sun is called *solar energy*. Solar energy can be converted to other forms of energy. In photocells or solar panels, solar energy is converted to *electrical energy*. Figure 4.20 shows a solar panel fixed on a roof.

Figure 4.20: *A solar panel*

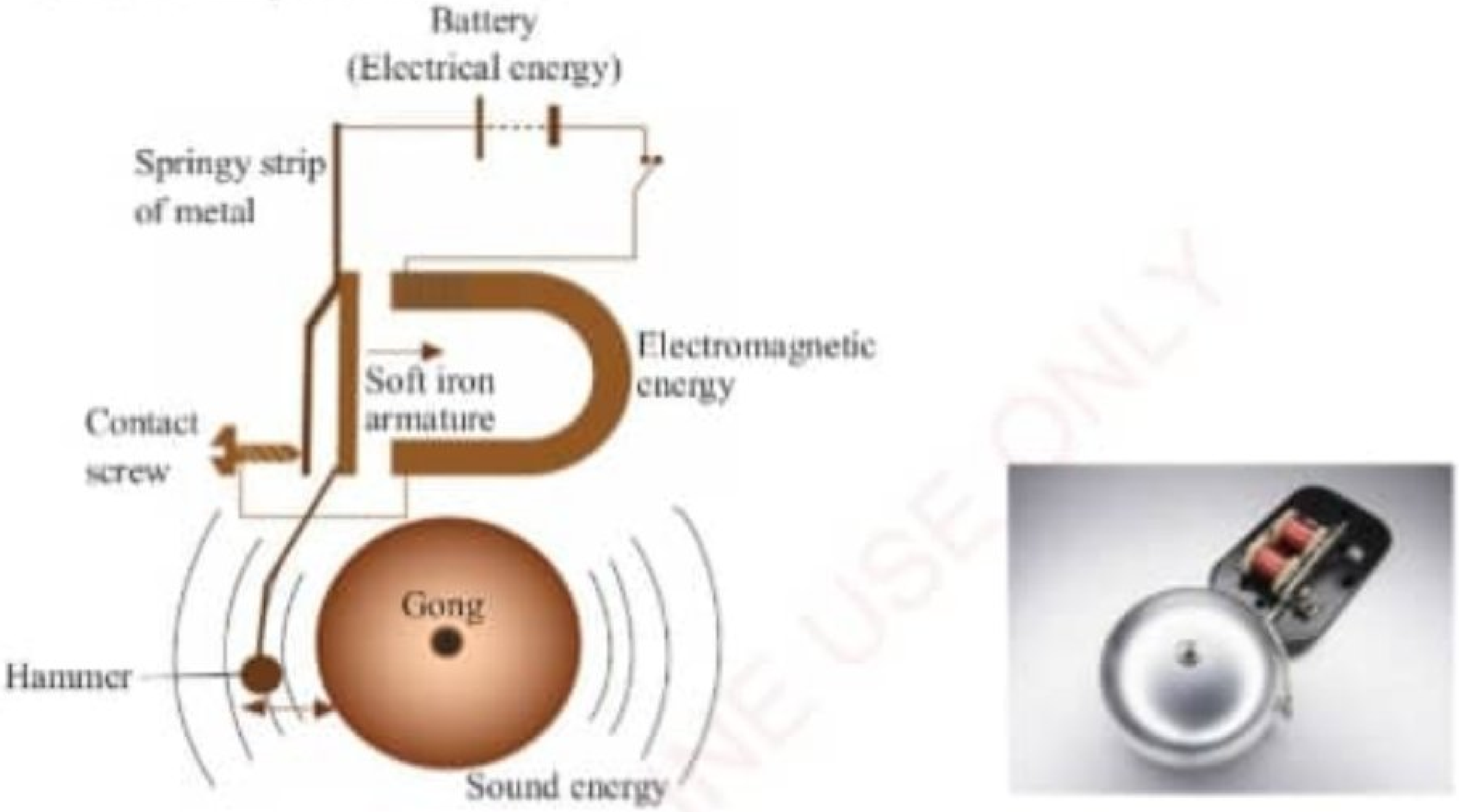
In solar cookers (Figure 4.21), a shiny metal surface is used to focus sun rays to produce heat. Solar energy is, therefore, converted to heat energy. A solar cooker can be in the form of a panel, box, parabola or tube.

Panel cooker *Box cooker* *Parabolic cooker* *Tube cooker*

Figure 4.21: *Forms of solar cookers*

Change of electrical energy to sound energy

In an electric bell, electrical energy is converted to sound energy. Figure 4.22 shows a simple electric bell.



(a) Sketch of an electric bell *(b) Photo of an electric bell*

Figure 4.22: *Electric bell*

The energy transformations in the electric bell shown in Figure 4.22 are summarised in Figure 4.`23.

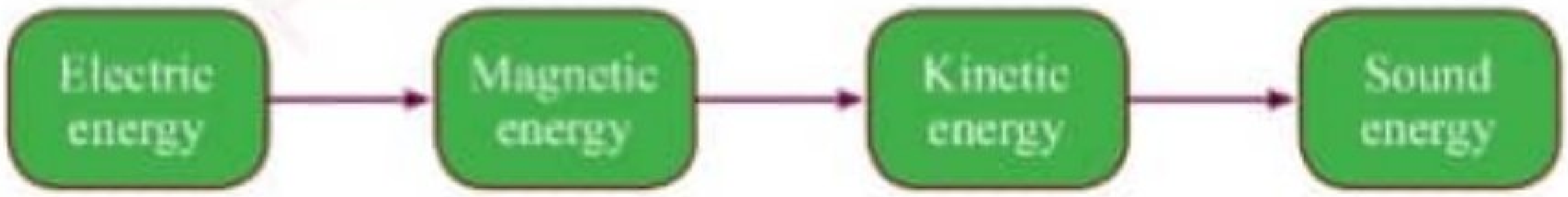


Figure 4.23: *Energy transformations in the electric bell*

Activity 4.2

Aim: To demonstrate the conversion of kinetic energy to electrical energy.

Requirements: Dynamos, connecting wires, stand with a clamp, a wheel (bicycle wheel), 6 V bulb, and bulb holder

Procedure

1. Set up the apparatus as shown in Figure 4.24.

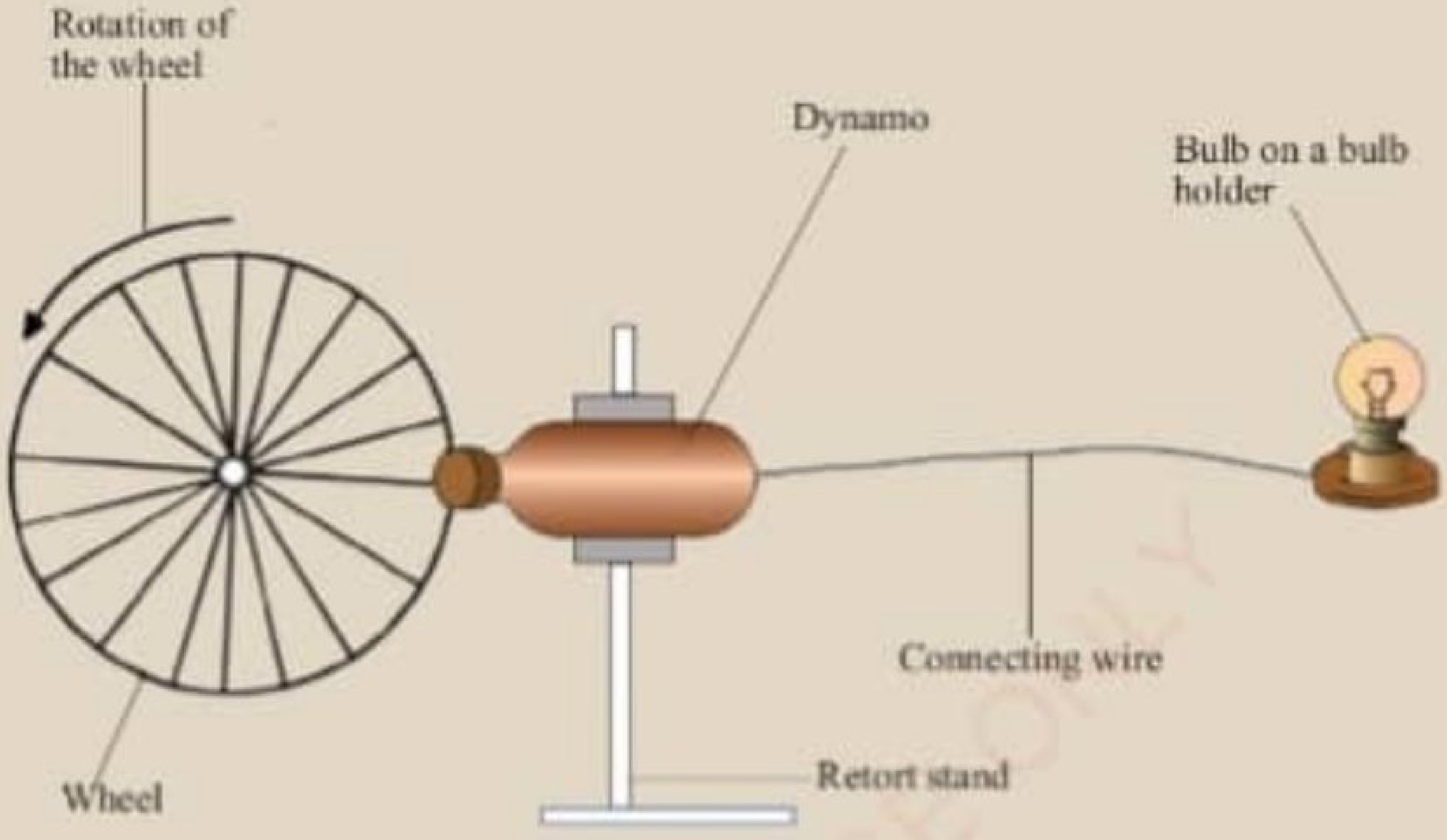


Figure 4.24: *Conversion of kinetic energy to electrical energy*

2. Fix the wheel from the centre so that it can rotate freely.
3. Connect the wires to the bulb holder and the dynamo.
4. Keep the wheel in contact with the dynamo before rotating.
5. Rotate the wheel at a constant speed.

Questions

1. What form of energy does the rotating wheel possess?
2. Describe the energy transformations that take place from the rotating wheel to the lit bulb.

Energy value of a fuel

The energy value of any fuel is the total amount of heat liberated by the complete combustion of a unit mass of the fuel in air (oxygen). Energy value is the most important characteristic of a fuel. Since energy is measured in joules or kilojoules and mass in grams or kilograms, energy value is given by the following equation:

$$\text{Energy value} = \frac{\text{Total energy liberated (J/kJ)}}{\text{Mass of the fuel used (g/kg)}}$$

The unit of energy value is joules per gram (J/g) or kilojoules per kilogram (kJ/kg).

Activity 4.3

Aim: To determine the energy value of methanol.

Requirements: Methanol burner with a lid, laboratory thermometer, metallic tin calorimeter, burette, weighing balance, retort stand and clamp, draught shield, gas lighter or matchbox, methanol, and distilled water

Procedure

1. Measure 200 cm^3 of distilled water using a burette and put it into an insulated metallic tin calorimeter.
2. Put some methanol, a volatile liquid in a burner. Weigh the burner with its contents. Note the mass.
3. Set the apparatus as shown in Figure 4.25.

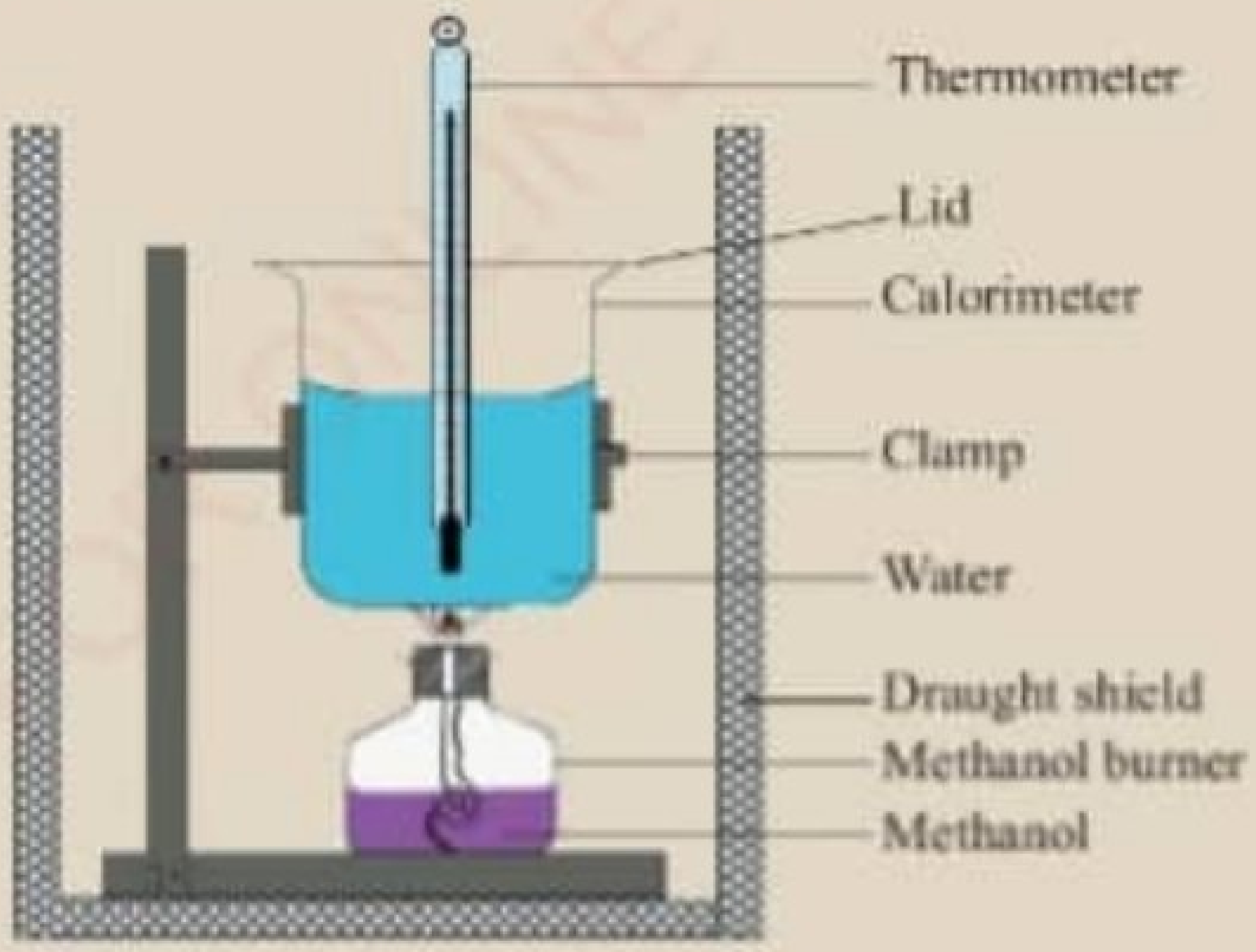


Figure 4.25: *Set-up of the determination of energy value of methanol*

4. Record the initial temperature of the water in Kelvin (K).
5. Light the methanol burner. Heat the water while stirring ensuring that the thermometer does not touch the bottom of the calorimeter, which is hotter than water.
6. Put off the burner when the temperature rises between 10 °C and 15 °C above the initial temperature.
7. Continue stirring and record the highest temperature reached in Kelvin.
8. Allow the burner to cool and weigh it again. Record your results as shown in Table 4.5.

Table 4.5: *Records for the determination of energy value of methanol*

Initial temperature of water (T_0)	
Final temperature of water (T_1)	
Temperature rise ($T_1 - T_0$)	
Initial mass of methanol + burner (M_0)	
Final mass of methanol + burner (M_1)	
Mass of methanol used ($M_0 - M_1$)	

Questions

1. What is the use of the calorimeter?
2. The methanol burner should have a tight lid. Why?
3. Why should the thermometer not touch the bottom of the calorimeter while stirring?
4. Determine the energy value of methanol. Given that the quantity of heat Q is calculated as:

   $Q = mc\Delta T$;

   where; m is the mass of water (kg), c is the specific heat capacity of water $= 4.18$ kJ kg^{-1} K^{-1} or 4.18 kJ kg^{-1} °C^{-1} and ΔT is the change in temperature.

   The volume of water = 200 cm^3 (0.0002 m^3); and density of water = 1000 kg m^{-3}

Example 4.1

The following results were obtained in an experiment to measure the heat value of biodiesel:

Initial temperature of water (T_0) = 24.7 °C = 297.7 K;
Final temperature of water (T_1) = 68.5 °C = 341.5 K;
Mass of biodiesel burnt = 56 g.

If the volume of water used in the experiment was 12 litres, determine the heat value of the biodiesel. (Specific heat capacity of water = 4.18 kJ kg^{-1} C^{-1}; density of water = 1000 kg m^{-3}).

Solution

Mass of water used = density × volume = 1000 kg m^{-3} × 0.012 m^3

= 12 kg

Change in temperature (ΔT) = $T_1 - T_0$ = (341.5 – 297.7) K = 43.8 K

$$Q = mc\Delta T$$

But, m = 12 kg, c = 4.18 kJ kg^{-1} K^{-1}, and ΔT = 43.8 K

Therefore,

$$Q = 12 \text{ kg} \times 4.18 \text{ kJ kg}^{-1} \text{K}^{-1} \times 43.8 \text{ K}$$
$$= 2197 \text{ kJ}$$

Thus, 0.056 kg of biodiesel gives 2197 kJ of heat

$$1 \text{ kg gives } \frac{2197}{0.056} = 39{,}232 \text{ kJ kg}^{-1} \text{ or } 39.23 \text{ kJ g}^{-1}$$

Therefore, the heat value of biodiesel is 39,232 kJ kg^{-1}

Exercise

1. A mass of 20.0 g of petrol was burnt in air. The heat produced was used to heat 2.5 litres of water. Given that, the heat value of petrol is 43640 kJ kg^{-1}, what was the temperature change of water?
2. Kerosene has a heat value of 43400 kJ kg^{-1}. Calculate the volume of kerosene required to raise the temperature of 20 litres of water from 24 °C to 100 °C.

   (Specific heat capacity of water = 4.18 kJ kg^{-1} K^{-1}; density of water = 1000 kg m^{-3}; density of kerosene = 810 kg m^{-3}).

Alternative sources of energy

Alternative sources of energy can be divided into *renewable* and *non-renewable sources*. Renewable sources of energy are those which are continually being replaced within short periods of time. They include solar energy, wind energy, and biomass. Non-renewable sources of energy are sources that cannot be replenished within short periods of time. They include fossil fuels such as petroleum, natural gas, coal, and nuclear energy.

Most of the energy being used in the world today comes from non-renewable sources of energy, mainly fossil fuels. These fuels are being used faster than they are being replaced. Since they take millions of years to form, they may get depleted in the near future. Therefore, there is a need to develop alternative sources of energy, especially the renewable ones.

Solar energy

Solar energy refers to energy that is obtained from sunlight. The origin of all the types of energy is the sun. Remember that energy cannot be created nor destroyed, but can just be changed from one form to another. Solar energy cannot be depleted. As long as the sun exists, there will always be solar energy reaching the Earth's surface. This type of energy is clean since it does not release harmful gases into the atmosphere. Solar energy can be tapped in various ways. These include the following:

(i) Using photovoltaic solar cells in generating electricity;

(ii) Using concentrated solar power in generating electricity;

(iii) Using parabolic mirrors (Figure 4.26) that focus the sun's rays to a central position for heating and cooking, and;

(iv) Using solar chimneys for heating and cooling.

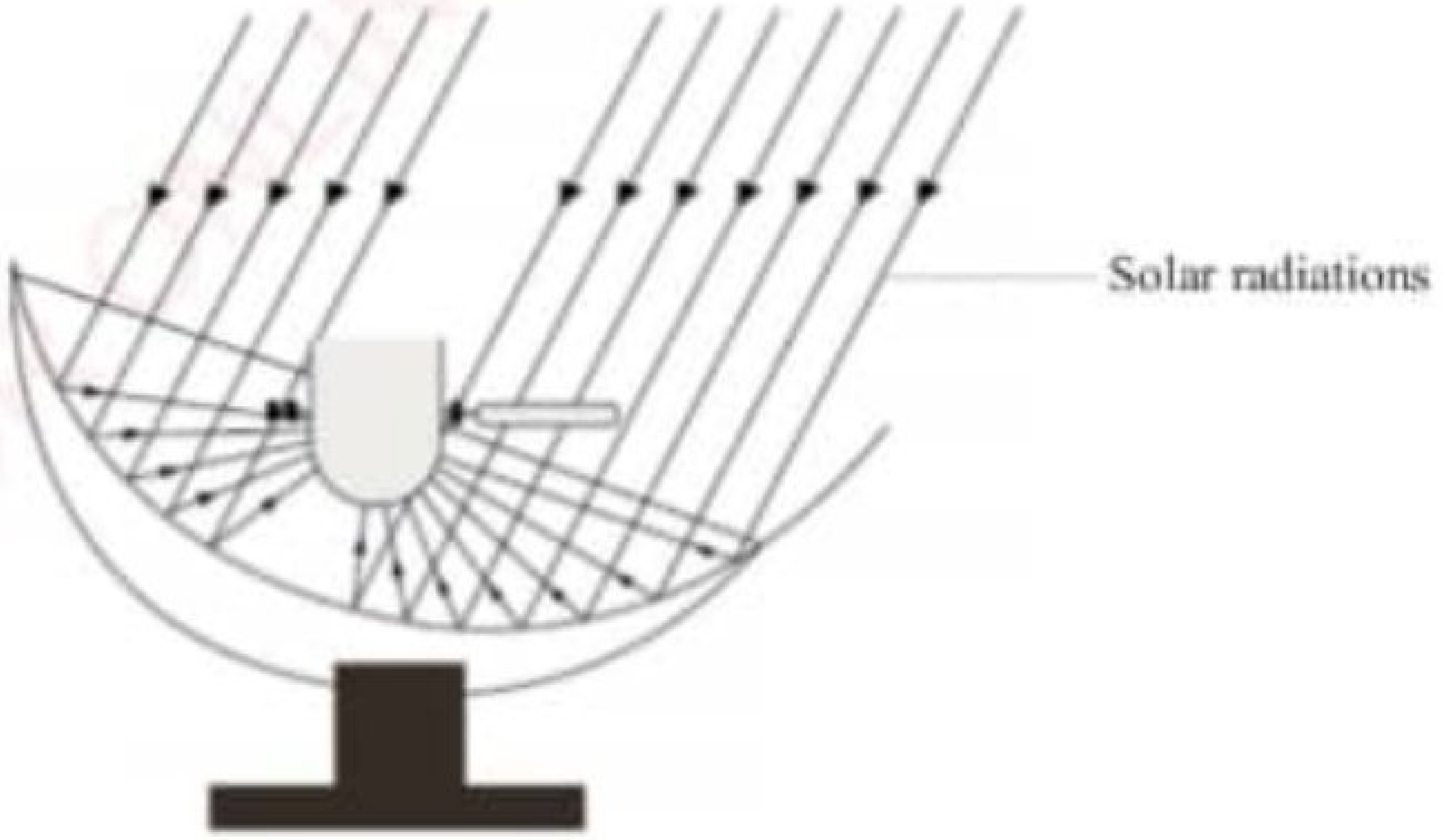


Figure 4.26: *Sketch of a solar cooker made using parabolic mirrors*

Biomass energy

Biomass refers to the organic matter in living organisms. Biomass energy is contained in organic compounds that are produced in growing plants and animals. Biomass energy is actually solar energy stored in organic matter. As plants grow, they use solar energy to make food in the form of carbohydrates through the process of photosynthesis. Carbohydrates are the organic compounds that make up biomass. When plants die, they decay and release the energy stored in the carbohydrates. Figure 4.27 shows a photo of some plants that are among the sources of biomass energy.

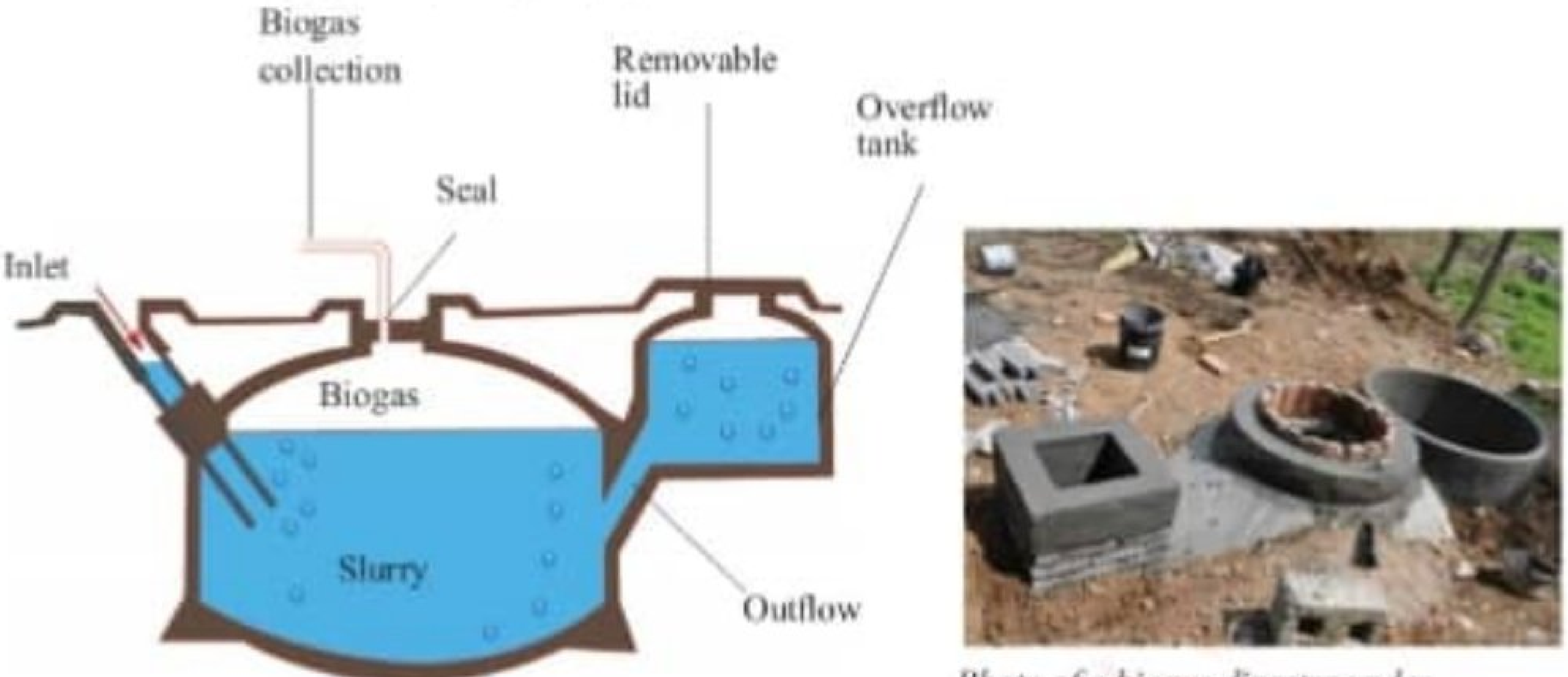
Figure 4.27: *Photo of maize plants that are among the sources of biomass energy*

Biomass is a renewable energy source, because the growth of new plants replenishes the supply. In addition, using biomass to produce energy is often a way of disposing waste materials. Biomass can be used directly as fuel or indirectly to produce liquid biofuel. Biomass fuels such as biodiesel, ethanol, and bagasse (a by-product of sugar cane processing) produced from agricultural products can be used in internal combustion engines and boilers.

The main advantages of biofuels are that, they contribute very little to global warming unlike fossil fuels. Since biofuels are produced from various sources such as straw, timber, manure, rice husks, sugarcane, flaxseed, and palm oil, their supplies are almost limitless. Other biodegradable outputs from industry, agriculture, households, and forestry can also be used as fuels. Biomass is definitely going to be the fuel for the future.

Biogas

Biogas is a gaseous fuel derived from decomposing biological waste. Biogas can easily be produced from both domestic and industrial wastes such as agricultural waste, sewage, and animal waste. The waste matter is put together and allowed to ferment naturally, thus producing biogas. This can be done by converting the existing waste disposal channels into biogas plants, sometimes called *biogas digesters* (Figure 4.28). When a biogas plant has extracted all the methane, the remains can be used as fertiliser.



Sketch of a digester

Photo of a biogas digester under construction

Figure 4.28: *Digester*

Wind energy

Wind is moving air. Wind energy is usually harnessed using windmills (Figure 4.29). The wind turns the blades of the windmills, which in turn run turbines and produce energy. Areas which have frequent strong wind currents such as offshore areas, and high-altitude areas, for example, Singida, Makambako and Mwenga in Tanzania are preferred locations for tapping this form of energy. Figure 4.29 shows the Mwenga windfarm in Mufindi District.

Figure 4.29: *Mwenga windfarm in Mufindi District-Tanzania*

(Source: *https://cxi.energy/project/installation-and-erection-of-the-2-4mw-mwenga-hydro-wind-farm-tanzania/*)

Wind power is a renewable source of energy which does not release harmful gases such as carbon dioxide and methane into the atmosphere. Wind strength near the Earth's surface varies, and thus, it cannot guarantee continuous power supply unless combined with other sources of energy.

Water power

Water possesses energy in the form of kinetic energy due to motion or thermal energy resulting from temperature differences. This energy can be harnessed and used. There are various forms of water energy. These include:

(a) *Hydroelectric energy*, which is the energy produced in a hydroelectric power plant;

(b) *Tidal stream energy*, which is the energy resulting from the flow of tides;

(c) *Wave energy*, which is the energy resulting from the movements of water waves; and

(d) *Ocean thermal energy*, which is the energy resulting from the temperature difference between the warmer surface of the ocean and colder deep parts of the ocean.

Geothermal energy

Geothermal energy is the heat that comes from the sub-surface of the earth. It is contained in the rocks and fluids beneath the earth's crust and can be found as far down the hot molten rock called magma. To produce power from geothermal energy, wells are dug deep into underground reservoirs to access the steam and hot water which can then be used to drive turbines connected to electric generators.

Chapter summary

1. Fuels are substances which undergo combustion reactions producing large amounts of heat that can be used in homes, transportation and industries as well as in other uses.
2. Fuels can be classified into natural (primary) fuels and artificial (secondary) fuels based on their occurrence.
3. Based on their physical states, fuels can also be classified into solid fuels, liquid fuels and gaseous fuels.
4. Fossil fuels contribute to environmental pollution and global warming.

5. Fossil fuels are derived from organic materials which died and decayed millions of years ago.
6. Fossil fuels might be exhausted in the near future since they are non-renewable.
7. Sources of energy are classified into renewable and non-renewable sources.
8. Energy is the capacity or ability of a system to do work.
9. The principle of conservation of energy states that *"energy can neither be created nor destroyed, it is only transformed from one form to another"*.
10. Alternative sources of energy include solar energy, biomass, biogas, wind, and water powers.

Revision exercise 4

1. Choose the correct answer for each of the following items:

   (i) Which of the following is not a fossil fuel?

   (a) Coal (b) Biodiesel
   (c) Natural gas (d) Petroleum

   (ii) Which of the following can be classified as a renewable source of energy?

   (a) Biomass (b) Diesel
   (c) Coal (d) Petroleum

   (iii) Wind is a promising future source of energy because
   (a) it does not produce harmful gases.
   (b) it does not involve chemical reactions.
   (c) it is renewable.
   (d) it cannot be seen.

   (iv) A good fuel is determined by its
   (a) energy value.
   (b) high content of non-combustible material.
   (c) scarcity.
   (d) high production of carbon dioxide.

(v) The main aim of destructive distillation of coal is
   (a) removal of oxygen in the atmosphere.
   (b) removal of volatile matter.
   (c) addition of volatile matter.
   (d) addition of oxygen in the furnace.

(vi) A good charcoal burns with
   (a) luminous flame.
   (b) non-luminous flame.
   (c) very low energy value.
   (d) high production of gases.

(vii) Gaseous fuels include
   (a) water gas and petrol.
   (b) water gas and kerosene.
   (c) water gas and producer gas.
   (d) coke and producer gas.

2. Describe the energy transformations that take place in each of the following cases:
   (a) Energy from the sun is used to generate electricity for lighting a house.
   (b) Mechanical energy from the waterfalls is used to generate electricity.
   (c) A bicycle wheel is used to turn a dynamo. The electric energy from the dynamo is used to power a bulb to produce light.

3. Give two examples of each of the following fuels:
   (a) Solid fuels
   (b) Liquid fuels
   (c) Gaseous fuels

4. (a) Explain why petroleum and coal are non-renewable sources of energy.
   (b) Give five alternative sources of renewable energy.

5. Explain the environmental effects of using charcoal as a source of fuel.

6. Explain the working mechanism of a biogas plant.

Chapter Five

Atomic Structure

Introduction

Substances are made up of very small particles called atoms. In this chapter, you will learn about the atomic theory, sub-atomic particles, arrangement of electrons, atomic number, mass number, and the isotopes. The competencies developed will help you to determine the composition, behaviour, and properties of different chemical substances.

The atom

You have already learnt that chemistry is the study of matter and its particulate nature. About the year 400 BC, a Greek philosopher known as Democritus was the first to consider the idea that matter is made up of particles. Such idea was not accepted because there was no experimental evidence to support it. About 2000 years later, an English man called John Dalton revived the discussion. He used experimental evidence to convince people that matter is made up of particles called atoms. It is through that experiment he deduced the Dalton's spherical model of the atom as shown in Figure 5.1.

Figure 5.1: *Dalton's model of the atom*

The atomic theory

In 1803, Dalton developed the theory about the atom. The four main points (assumptions) of Dalton's Atomic Theory are summarised as follows:

1. Matter is made up of tiny particles called atoms. (The word atom means 'unsplittable' in Greek).
2. Atoms can neither be created nor destroyed.
3. Atoms of the same element are identical, and have the same mass and properties. Atoms of a given element are different from those of any other element. The atoms of different elements can be distinguished from one another by their respective relative weights.

4. Compounds are formed by a combination of two or more different kinds of atoms. The atoms always combine in simple whole number ratios.

Dalton never imagined that anyone would ever be able to see an atom. However, modern technology has provided direct evidence that shows the positions and patterns of individual atoms. The use of modern technology has enabled scientists to carry out experiments on the atom that Dalton could not. This has led to slight modifications to the Dalton's Atomic Theory and thus formulated the so called modern concepts of Dalton's Atomic Theory.

These modifications include the following:

1. Atoms can be created or destroyed or split by means of nuclear reactions. For example, an atom of uranium-235 can be split into two separate atoms by a process called *nuclear fission*.
2. Some elements have atoms of more than one kind which differ slightly in mass. Such atoms are called *isotopes*. For example, carbon has three isotopes known as carbon-12, carbon-13, and carbon-14.
3. An atom is made up of smaller sub-atomic particles called *protons*, *neutrons*, and *electrons*.
4. Atoms of different elements may combine in many different ratios to form complex compounds.

Sub-atomic particles

In the nineteenth century, J. J. Thomson carried out experiments and described an atom as a sphere of positive charge, with negative particles called *electrons* spread throughout the sphere. This model of the atom was referred to as *'plum pudding'* model and is shown in Figure 5.2. Thomson therefore, managed to discover the *electron* among the three sub-atomic particles.

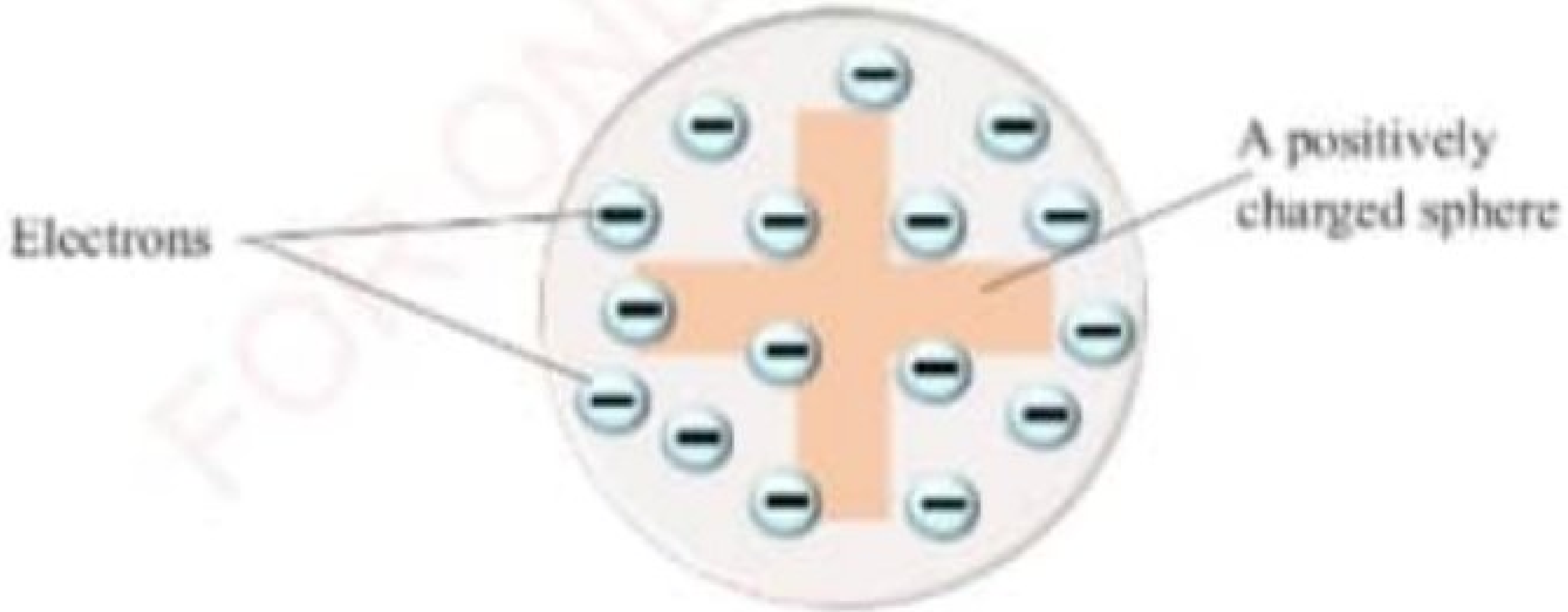


Figure 5.2: *Thomson's 'plum pudding' model of the atom*

However, another scientist called Ernest Rutherford reasoned that if Thomson's model was correct, then the mass of the atom was evenly spread throughout the atom. He carried out experiments and discovered that most of the mass of an atom is actually concentrated in the nucleus (central core) of the atom. Within the nucleus there are positively charged particles called *protons*. This was the second sub-atomic particle to be discovered.

Rutherford's findings are summarized as follows:

1. Protons, the positively charged particles of an atom are located in the nucleus.
2. Most of the mass of the atom is located in the nucleus.
3. The nucleus has a relatively smaller volume compared to the whole atom.
4. Electrons have very small masses compared to the protons.
5. Most of the space in an atom is empty.
6. Electrons are the negatively charged particles in an atom. They move around the nucleus in orbits.

Rutherford thus developed the *planetary model* of the atom as shown in Figure 5.3.

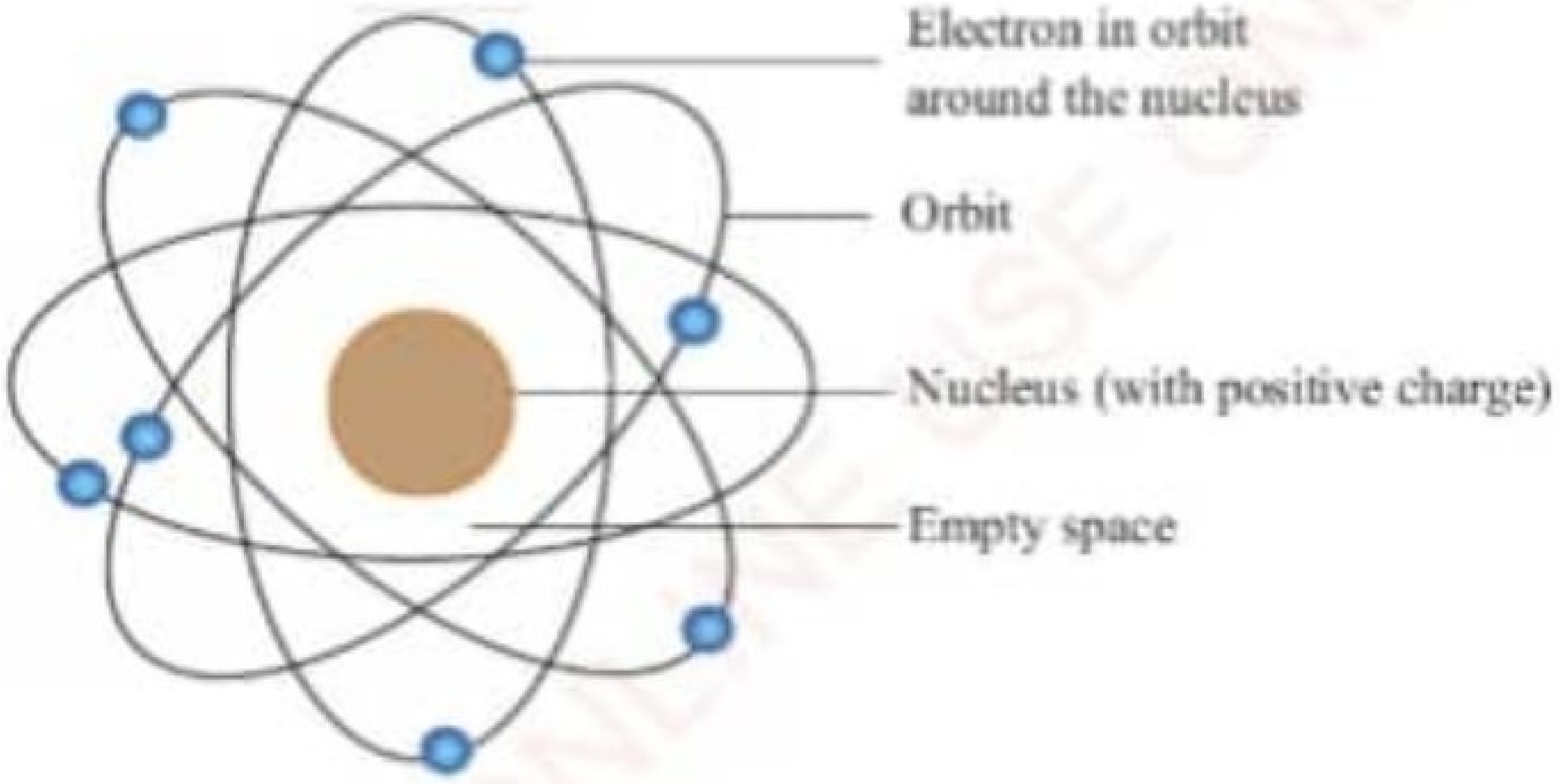


Figure 5.3: *Rutherford's planetary model of the atom*

In 1934, another scientist called Chadwick established that there were *neutrons* which also formed part of the nucleus. Figure 5.4 shows the location of neutrons in an atom. Neutrons have the same mass as the protons but no charge. They are located in the nucleus of an atom. They were the third sub-atomic particles to be discovered.

The properties of the neutrons are summarized as follows:

1. They have no charge (are neutral).
2. They have nearly the same mass as the corresponding protons.
3. They have a mass nearly 1840 times the mass of an electron.

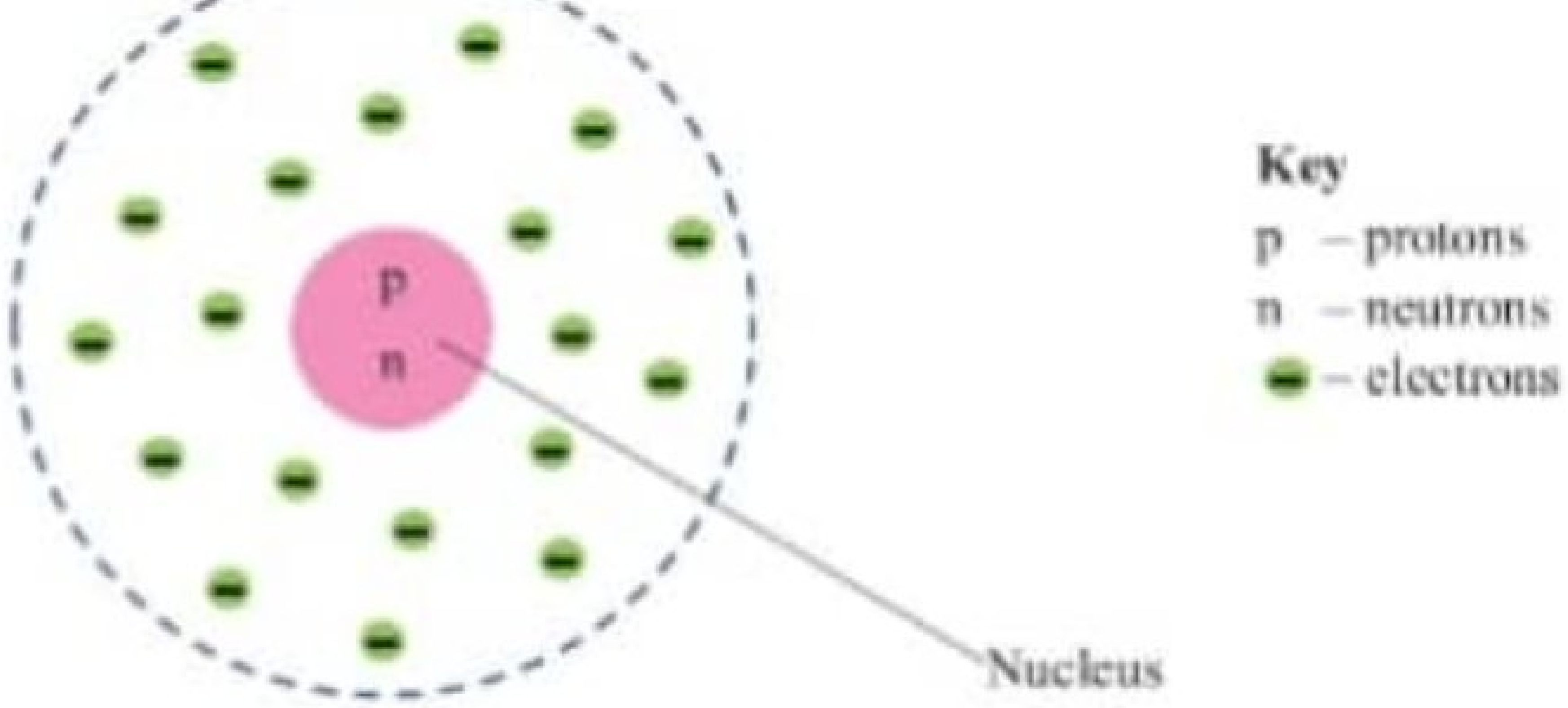


Figure 5.4: *Location of sub-atomic particles of the atom*

Table 5.1 gives a summary of the properties of sub-atomic particles of an atom.

Table 5.1: *Properties of sub-atomic particles*

Sub-atomic particle	Symbol	Location	Charge	Real mass (g)	Relative mass
Proton	p	In the nucleus	+1	1.6726×10^{-24}	1
Neutron	n	In the nucleus	0	1.6750×10^{-24}	1
Electron	e^-	Outside the nucleus	–1	9.109×10^{-28}	$\frac{1}{1840}$

Electron arrangement

In 1913, Neils Bohr suggested that electrons rotate around the nucleus in special regions called *shells* or *orbits*. These shells (also known as *energy levels*) are at fixed distances from the nucleus. Each shell can only hold a specific number of electrons. The maximum number of electrons held within each shell can be determined by the formula $2n^2$, where n is the position of the shell from the nucleus. According to this formula, the:

first shell can hold $(2 \times 1^2) = 2$ electrons;

second shell can hold $(2 \times 2^2) = 8$ electrons; and

third shell can hold $(2 \times 3^2) = 18$ electrons.

The first four shells are represented by the letters K, L, M, and N, respectively as shown in Figure 5.5. Each electron in an atom is in a particular shell and the electrons must first occupy the lowest available shell nearest to the nucleus.

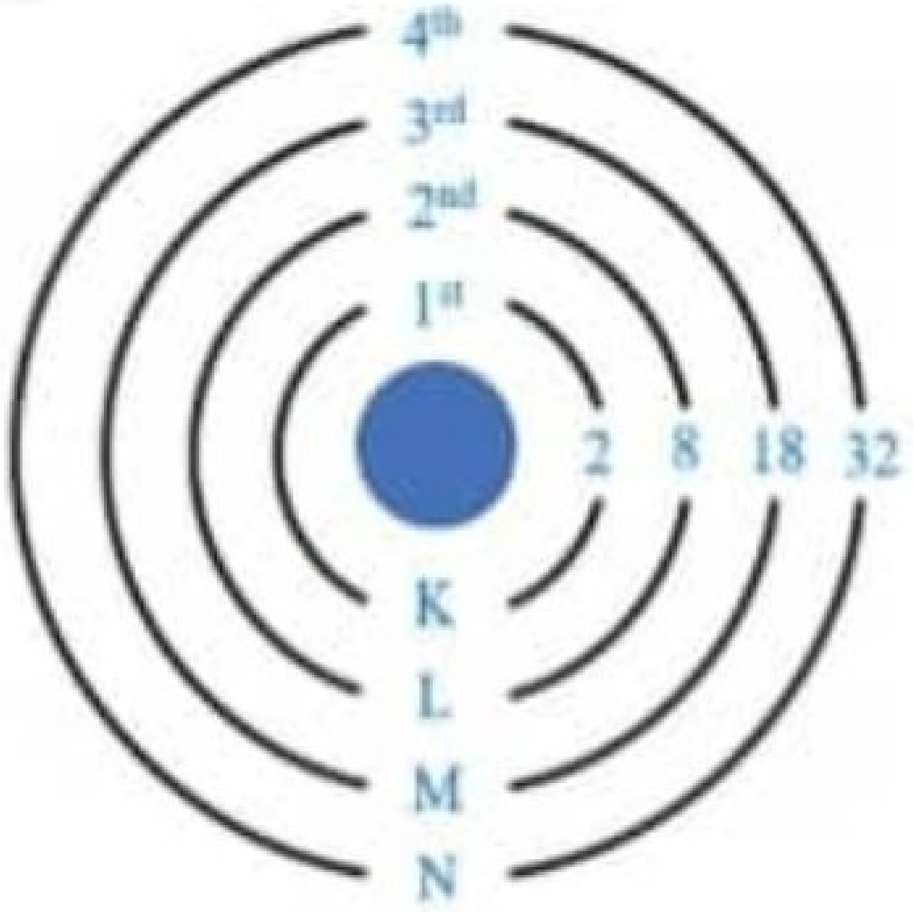


Figure 5.5: *First four shells of an atom*

For reasons beyond the scope of this book, the 3rd shell is more stable with 8 electrons. This is why even though the 3rd shell can hold up to 18 electrons, the potassium element which has 19 electrons has only 8 electrons in its 3rd shell, and the last electron moves to the 4th shell. This is the same for calcium which has 20 electrons, where the last two electrons move to the 4th shell after the 3rd shell is completely filled.

A shell which contains its maximum number of electrons is called a fully-filled shell. An atom with fully-filled outermost shell is said to be stable. Some atoms have 2 electrons (e.g. helium) or 8 electrons (e.g. neon). The elements with 2 electrons in their outermost shells are said to exhibit a *duplet state*, while those with 8 electrons are said to exhibit an *octate state*. Electrons are arranged so that the lowest shells are filled first. This arrangement of electrons in different shells in an atom is called *electronic configuration*. Figure 5.6 shows the diagrammatic electronic configurations of hydrogen, helium, neon, potassium, and sulphur atoms.

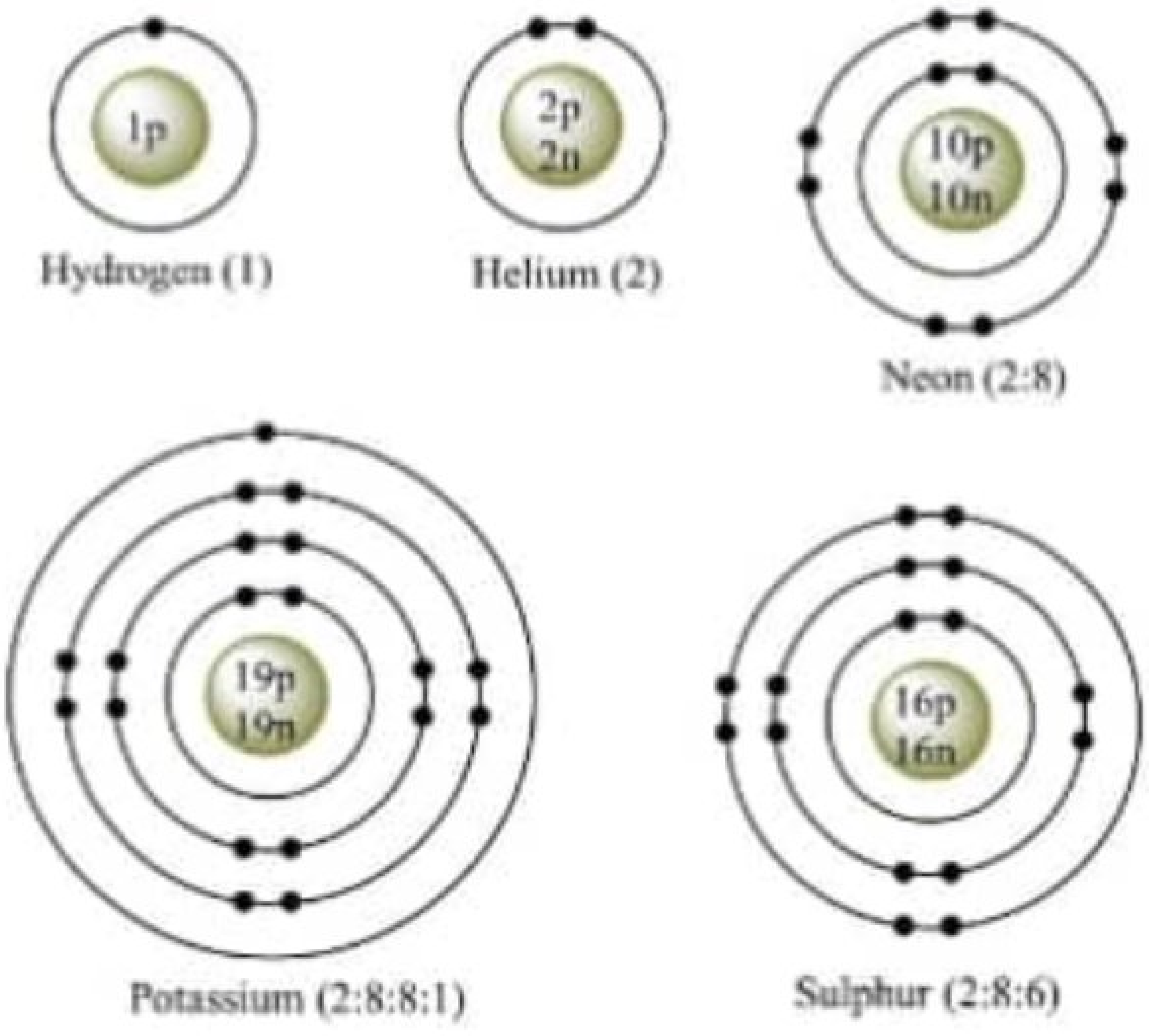


Figure 5.6: *Electronic configurations of some atoms*

Note that the hydrogen atom has no neutron in its nucleus. Helium and neon atoms have their outermost shells completely filled with electrons, and so they are stable atoms. The electrons are not fixed at particular positions within the sphere or shell, instead, they move extremely fast and can be at any point within the shell.

Bohr's findings provided more information about elements that are summarised in a table which shows the number of electrons in each shell. This table is called the *Periodic Table* (Appendix 1). Note that the number of electrons or protons in an atom determines the position of an element in the Periodic Table. At this level, you will learn the first twenty elements in the Periodic Table. The electronic arrangements of the twenty elements are shown in Table 5.2.

Table 5.2: *The electronic arrangements of the first twenty elements in the Periodic Table*

Element	Chemical symbol	Number of electrons in each shell					Electronic configuration (arrangement)
		Number of electrons	1st	2nd	3rd	4th	
Hydrogen	H	1	1				1
Helium	He	2	2				2
Lithium	Li	3	2	1			2:1
Beryllium	Be	4	2	2			2:2
Boron	B	5	2	3			2:3
Carbon	C	6	2	4			2:4
Nitrogen	N	7	2	5			2:5
Oxygen	O	8	2	6			2:6
Fluorine	F	9	2	7			2:7
Neon	Ne	10	2	8			2:8
Sodium	Na	11	2	8	1		2:8:1
Magnesium	Mg	12	2	8	2		2:8:2
Aluminium	Al	13	2	8	3		2:8:3
Silicon	Si	14	2	8	4		2:8:4
Phosphorus	P	15	2	8	5		2:8:5
Sulphur	S	16	2	8	6		2:8:6
Chlorine	Cl	17	2	8	7		2:8:7
Argon	Ar	18	2	8	8		2:8:8
Potassium	K	19	2	8	8	1	2:8:8:1
Calcium	Ca	20	2	8	8	2	2:8:8:2

Task 5.1

1. You are provided with twenty folded pieces of paper, each containing a name among the first twenty elements of the Periodic Table.
2. Form groups of three to five students and perform the following activities:
   (a) Pick randomly four pieces of folded papers.
   (b) List down the four elements you have picked.
   (c) Using coloured markers and manila papers, draw circles to represent the nucleus and the respective number of shells for each type of atom.

(d) Use dots of different colours to represent the three types of sub-atomic particles.

(e) Compile all the illustrations and let them be pinned on the laboratory notice-board. This should be systematically arranged in the order of the positions of the elements in the Periodic Table.

Atomic number and mass number

You have noted that atoms are made up of three main sub-atomic particles: protons, neutrons, and electrons. These particles have relations to the atomic number and mass number of the atom.

Atomic number

The *atomic number* is the number of protons in an atom. It is also known as the *proton number*. For example, the atomic number of hydrogen is 1 since it has only one proton. A sodium atom has 11 protons in the nucleus, therefore, its atomic number is 11. Since the number of protons is equal to the number of electrons in the atom, the atomic number also indicates the number of electrons in the atom. Thus, for a neutral atom, the atomic number is not only the number of protons in an atom, but it is also the number of electrons.

Therefore;

Atomic number = Number of protons = Number of electrons

Mass number

Protons and neutrons are found in the nucleus of an atom and are called *nucleons*. The sum of the protons and neutrons in one atom of an element is called the *mass number* or *nucleon number* or *atomic mass*. This number is actually taken as the mass of the atom since the mass of the electron is negligible.

Thus, Number of protons + Number of neutrons = Mass number

For example;

(i) Hydrogen has 1 proton and 0 neutrons. Therefore, its atomic number is 1, and mass number is 1 + 0 = 1.

(ii) Boron has 5 protons and 6 neutrons. Its atomic number is 5 and mass number is 5 + 6 = 11.

(iii) Nitrogen has 7 protons and 7 neutrons. Its atomic number is 7 and mass number is 7 + 7 = 14.

It is also possible to calculate the number of neutrons and number of electrons of an atom if its mass number and atomic number are given.

Example 5.1

Atom Q has a mass number of 49 and an atomic number of 24. What is its number neutrons? What is the number of electrons in atom Q?

Solution

Mass number = 49; atomic number = 24

(a) Neutron number = mass number – atomic number = 49 – 24 = 25

(b) Number of electrons = number of protons = atomic number = 24

Note: For the mass number with fractions, for example, chlorine (35.5), calculating the number of neutrons and electrons involves only a whole number. In this case, for chlorine, 35 is used.

Exercise 5.1

Complete the table below by filling in the number of protons, electrons, and neutrons of the atoms. The atomic numbers and mass numbers are given.

Atom	Atomic number	Mass number	Protons	Electrons	Neutrons
Sodium	11	23			
Oxygen	8	16			
Beryllium	4	9			
Fluorine	9	19			

Nuclide notation

Atoms of different elements can be represented by chemical symbols that indicate their respective atomic numbers and mass numbers. Using an arbitrary element X, the mass number (A) is placed on its upper left end, while its atomic number (Z) is placed on the lower left end. Thus, element X is shown as $^{A}_{Z}X$. This is known as the *nuclide notation*. The following are examples of nuclide representations of different atoms:

(i) Hydrogen ® $^{1}_{1}H$ (ii) Boron ® $^{11}_{5}B$

(iii) Nitrogen ® $^{14}_{7}N$ (iv) Oxygen ® $^{16}_{8}O$

With this information, it is possible to deduce the number of neutrons and electrons in the atom, and to write the electronic configuration. For example, in the oxygen atom, 16 is the mass number and 8 the atomic number. Therefore, the number of neutrons is 16 – 8 = 8. The nucleus of the oxygen atom can therefore be represented as shown in Figure 5.7.

Figure 5.7: *The nucleus of the oxygen atom*

Example 5.2

Potassium atom has 19 electrons and the mass number of 39.

(a) Workout the:

(i) atomic number, and

(ii) number of neutrons.

(b) The symbol for potassium is K. Give the nuclide notation.

(c) Show the representation of the nucleus of the potassium atom.

(d) Draw the electronic configuration of potassium.

Solution

(a) (i) Atomic number = number of protons = number of electrons = 19

(ii) Mass number = number of protons + number of neutrons

Number of neutrons = mass number – number of protons

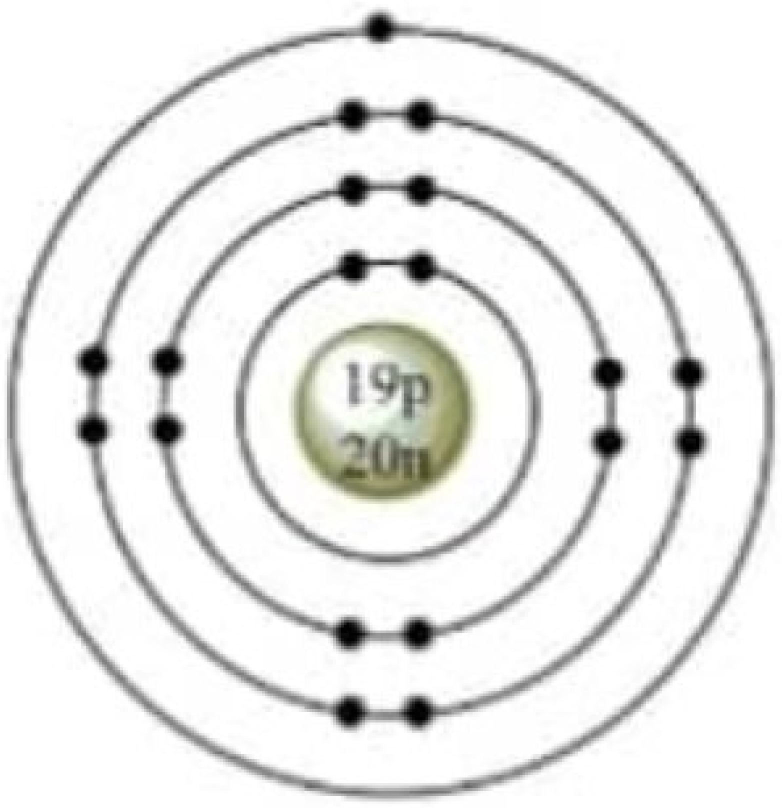
= 39 – 19

= 20

(b) $^{39}_{19}K$; where 39 is the mass number and 19 the atomic number.

(c)

(d)



Isotopes

Atoms of the same element have the same number of protons. However, the number of neutrons in the atoms of the same element may vary. This means that the atomic number of an element does not vary but the mass number can vary. Such atoms of an element are called *isotopes*. Isotopes are atoms of the same element with the same number of protons but different number of neutrons. Such existence of the element is called *isotopy*. Isotopy is the existence of atoms of the same element having the same atomic number but different mass numbers. It is also possible to get the number of sub-atomic particles in a given isotope.

Example 5.3

State the number of protons, neutrons, and electrons in the following isotopes:

(a) $^{12}_{6}C$ and $^{14}_{6}C$

(b) $^{1}_{1}H$, $^{2}_{1}H$ and $^{3}_{1}H$

Solution

(a) $^{12}_{6}C$, Mass number = 12

Number of protons = atomic number = 6

Number of electrons = number of protons = 6

Number of neutrons = 12 – 6 = 6

$^{14}_{6}C$, Mass number = 14
Number of protons = 6
Number of electrons = 6
Number of neutrons = 14 – 6 = 8

(b) $^{1}_{1}H$, Mass number = 1
Number of proton = 1
Number of electron = 1
Number of neutron = 1 – 1 = 0

$^{2}_{1}H$, Mass number = 2
Number of proton = 1
Number of electron = 1
Number of neutrons = 2 – 1 = 1

$^{3}_{1}H$, Mass number = 3
Number of proton = 1
Number of electron = 1
Number of neutrons = 3 – 1 = 2

Example 5.4

An isotope of carbon has a mass number of 13 and an atomic number of 6.

(a) Write its nuclide notation.

(b) How many neutrons does it have?

(c) How many electrons does it have?

Solution

(a) $^{13}_{6}C$

(b) Number of neutrons = 13 – 6 = 7

(c) Number of electrons = atomic number = 6

Many elements that occur naturally usually display isotopy. The most abundant (plentiful) isotope of an element is taken to be the representative of that element. This abundance is usually given in percentage. Examples of common elements that display isotopy are hydrogen, oxygen, carbon, chlorine, nitrogen, and neon (Table 5.3).

Table 5.3: *Examples of isotopes and their abundances*

Element	Chemical symbol	Atomic number	Isotopes	Abundance
Hydrogen	H	1	${}^{1}_{1}H$ (protonium or hydrogen)	99.99%
			${}^{2}_{1}H$ (deuterium)	0.01%
			${}^{3}_{1}H$ (tritium)	Very rare
Carbon	C	6	${}^{12}_{6}C$	98.9%
			${}^{13}_{6}C$	1.1%
			${}^{14}_{6}C$	Trace
Chlorine	Cl	17	${}^{35}_{17}Cl$	75%
			${}^{37}_{17}Cl$	25%
Oxygen	O	8	${}^{16}_{8}O$	99.8%
			${}^{17}_{8}O$	0.037%
			${}^{18}_{8}O$	0.20%
Neon	Ne	10	${}^{20}_{10}Ne$	90.5%
			${}^{21}_{10}Ne$	0.3%
			${}^{22}_{10}Ne$	9.2%
Nitrogen	N	7	${}^{14}_{7}N$	99.6%
			${}^{15}_{7}N$	0.4%
			${}^{13}_{7}N$	Very rare

Relative atomic mass

An atom is very small and it would be difficult to measure its actual mass. To overcome this difficulty, chemists developed a simpler way to express the mass of an atom. This involved expressing the mass of an atom in relation to a chosen standard atomic mass. The carbon atom was chosen as the standard atom (reference atom) and its mass was arbitrarily chosen as 12 units (not actual value). Then, using an instrument called a mass spectrometer, all the other atoms were compared to this standard atom. This reference is called the *carbon-12 scale*. For example, it was found that:

(i) the magnesium atom was twice as heavy as the reference atom; so its mass was put at 24.

(ii) the hydrogen atom was $\frac{1}{12}$ as heavy as the reference atom; so its mass was put at 1.

(iii) the helium atom was $\frac{1}{3}$ as heavy as the reference atom; so its mass was put at 4.

The mass of an atom obtained by comparing it with the arbitrary mass of a carbon-12 atom is called its *relative atomic mass* (R.A.M. or A_r). The relative atomic mass of an element is the average mass of one atom of the element relative to $\frac{1}{12^{th}}$ the mass of one atom of carbon-12. Therefore, R.A.M. may not necessarily be a whole number.

That is, $A_r = \dfrac{\text{Average mass of atom of an element}}{\frac{1}{12^{th}}\text{the mass of carbon -12 atom}}$

Table 5.4 gives the atomic numbers and relative atomic masses of the first 20 elements in the Periodic Table. From Table 5.4 you can see that not all relative atomic masses are whole numbers. This is because most elements display a degree of isotopy. The relative atomic mass of such elements is obtained by calculating the average mass of all the isotopes of each element.

For isotopic elements, the relative atomic mass (R.A.M.) can be calculated using the following formula:

Relative atomic mass (R.A.M.) = The sum of isotopic mass × Percentage abundance

Table 5.4: *Atomic numbers and relative atomic masses of some elements*

Element	Atomic number	Relative atomic mass
Hydrogen	1	1
Helium	2	4
Lithium	3	6.9
Beryllium	4	9
Boron	5	10
Carbon	6	12
Nitrogen	7	14
Oxygen	8	16
Fluorine	9	19
Neon	10	20.2
Sodium	11	23
Magnesium	12	24.3
Aluminium	13	27
Silicon	14	28.1
Phosphorus	15	31
Sulphur	16	32
Chlorine	17	35.5
Argon	18	39.9
Potassium	19	39.1
Calcium	20	40.1

Example 5.5

(a) Chlorine has two isotopes:

$^{35}_{17}\text{Cl}\ (75\%)$ and $^{37}_{17}\text{Cl}\ (25\%)$

The relative atomic mass of chlorine is:

$$\left(35 \times \frac{75}{100}\right) + \left(37 \times \frac{25}{100}\right)$$

$$= \frac{2625 + 925}{100}$$

$$= \frac{3550}{100} = 35.5$$

(b) Neon has three isotopes:

$^{20}_{10}Ne(90.5\%)$, $^{21}_{10}Ne(0.3\%)$, $^{22}_{10}Ne(9.2\%$

The relative atomic mass of neon is:

$$\left(20 \times \frac{90.5}{100}\right) + \left(21 \times \frac{0.3}{100}\right) + \left(22 \times \frac{9.2}{100}\right)$$

$$= \frac{1810 + 6.3 + 202.4}{100}$$

$$= \frac{2018.7}{100}$$

$$= 20.187$$

It can be noted that for both chlorine and neon, the R.A.M. is very close to the mass number of the isotope with the highest abundance, namely $^{35}_{17}Cl$ and $^{20}_{10}Ne$, respectively.

Task 5.2

Design a model of an atom using locally available materials. For example, you can use painted circles of different sizes to construct the shells of an atom. Use different colours to indicate the nucleus containing protons and neutrons. You can also use small coloured balls to indicate electrons in the respective shells.

Chapter summary

1. An atom is the smallest particle of an element. It can only be split or destroyed by nuclear reaction.
2. There are three major sub-atomic particles, namely:

   (a) protons (positively charged),
   (b) neutrons (neutral), and
   (c) electrons (negatively charged).

3. Protons and neutrons are located in the nucleus of an atom while electrons are found in the shells or energy levels around the nucleus.

4. The arrangement of electrons in different shells of an atom is known as electronic arrangement or electronic configuration.

5. Each shell can contain only a certain number of electrons, with the maximum being $2n^2$, where n is the position of the shell from the nucleus.

6. For any element:
   Number of protons = atomic number
   Number of electrons = number of protons = atomic number
   Number of neutrons = mass number – atomic number

7. Isotopes are atoms of the same element with the same number of protons but different number of neutrons.

8. The relative atomic mass of an element is the average mass of one atom of the element relative to $\frac{1}{12^{th}}$ the mass of one carbon-12 atom.

Revision exercise 5

1. Choose the correct answer for each of the following items:
   (i) Which description corresponds to a proton?
      (a) Relative mass = 1, charge = + 1
      (b) Relative mass = 1, charge = 0
      (c) Relative mass = $\frac{1}{1840}$, charge = –1
      (d) Relative mass = 4, charge = +2.

(ii) What are nucleons?
   (a) Neutrons and electrons
   (b) Neutrons and protons
   (c) Electrons and protons
   (d) Protons, neutrons and electrons

(iii) What is the mass of an electron compared to that of a proton?
   (a) $\frac{1}{18}$ (b) $\frac{1}{184}$ (c) $\frac{1}{1840}$ (d) $\frac{1}{18400}$

(iv) Most atoms are neutral because
   (a) the nucleus is only made up of neutrons.
   (b) they have equal numbers of electrons and protons in the shells.
   (c) the neutrons normally have zero charge.
   (d) the number of electrons balances out the number of protons in the atom.

(v) Which of these statements is true about isotopes of an element?
   (a) The number of protons is the same, but the number of neutrons is different.
   (b) The number of neutrons is the same, but the number of protons is different.
   (c) The number of protons and neutrons is the same, but the number of electrons is different.
   (d) The number of protons is the same, but electrons are added to the nucleus.

(vi) An isotope of cadmium has an atomic number of 48 and a mass number of 112. This means that cadmium atom has
   (a) 48 protons, 64 neutrons, and 48 electrons.
   (b) 64 protons, 48 neutrons, and 64 electrons.
   (c) 48 protons, 112 neutrons, and 48 electrons.
   (d) 112 protons, 48 neutrons, and 112 electrons.

(vii) Which sub-atomic particles are in equal number with protons in a neutral atom?
   (a) Electrons
   (b) Neutrons

(c) Electrons of its ion

(d) Neutrons of its ion

(viii) What is the maximum number of electrons in the innermost shell of an atom?

(a) 3 (b) 4

(c) 1 (d) 2

(ix) The atomic number for an element T is 9. What is its electronic configuration?

(a) 2 : 4 : 3 (b) 2 : 5 : 2

(c) 2 : 7 (d) 2 : 6 : 1

(x) Which value represents the mass number of an atom that has 12 neutrons and 11 electrons?

(a) 22 (b) 24

(c) 23 (d) 11

2. Copy the following table and fill the missing details of the arbitrary elements given.

Element	Atomic number	Mass number	Number of protons	Number of electrons	Electronic configuration
P	17	35			
Q		40	20		
R		12			2:4
S		9		4	
T	3	7			
U		23			2:8:1
V		31	15		

3. State the number of protons, neutrons and electrons in the following atoms:

(a) $^{27}_{13}Al$ (b) $^{137}_{56}Ba$ (c) $^{1}_{1}H$ (d) $^{90}_{38}Sr$ (e) $^{235}_{92}U$

4. Explain briefly the following terms:

(a) Proton

(b) Neutron

(c) Electron

(d) Atomic number
(e) Isotope
(f) Mass number
(g) Nuclide notation
(h) Relative atomic mass

5. Draw the structure of an atom according to the

(a) Dalton's Atomic Theory.

(b) Rutherford's Atomic Model.

6. An isotope of neon has a mass number of 21 and an atomic number of 10.

(a) Write its nuclide notation.

(b) How many neutrons does it have?

(c) How many electrons does it have?

7. Oxygen exists naturally in three different isotopes, which are $^{16}_{8}O$ (99.76%), $^{17}_{8}O$ (0.04%) and $^{18}_{8}O$ (0.20%). Calculate the relative atomic mass of oxygen.

8. Write down the four assumptions of the Dalton's Atomic Theory.

Chapter

Six

Periodic classification

Introduction

There are many elements which have been discovered, and therefore, it is very dificult to study each element separately. The elements are better studied and dealt with when they are classified. The best way to classify them is by using a table known as the Periodic Table. In this chapter, you will learn about the meaning and development of the Periodic Table, periodicity, and the general trends. The competencies developed will build the basics of how to study elements, which will also be a foundation to the understanding of Chemistry in general.

Development of the Periodic Table

For a long time, chemists have used various ways of grouping elements with similar properties. The simplest of these has been classifying elements as either metals or non-metals. In 1866, a British chemist, John Newlands, thought of the idea of arranging elements in order of their increasing atomic masses. Newlands arranged the elements according to Table 6.1. At that time, the noble gases had not been discovered.

Table 6.1: *Newlands' first arrangement of elements*

H	Li	Be	B	C	N	O	F	Na	Mg	Al	Si	P	S	Cl	K	Ca

Newlands noticed that an element tends to display characteristics similar to the 8th element in front of it. He arranged the elements in columns according to a law he called the *Law of octaves* (Table 6.2). However, his classification was unfortunate since he grouped together certain elements which had very different characteristics. For example, oxygen (O) was placed in the same group as iron (Fe) and sulphur (S). Newlands' ideas were therefore rejected by many scientists.

Table 6.2: *Newlands' octaves of elements*

H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe

A Russian chemist, Dimitri Mendeleev, later improved the Newlands' ideas and convinced other chemists to use them. He intended to illustrate recurring trends (periodic trends) in the properties of elements. In 1869, Mendeleev summarised his *Periodic Law* which states that: *the properties of elements are a periodic function of their relative atomic masses.* He arranged elements in order of their increasing atomic masses and by similarity of properties. This resulted in an early version of the Periodic Table of elements (Table 6.3). A vertical column of elements is called a *group* and a horizontal row is called a *period.* He, however, left gaps in the table predicting that there were existing elements yet to be discovered. His table did not include the noble gases, which had not been discovered at the time.

Table 6.3: *Part of Mendeleev's Periodic Table of 1871*

Group / Period	1	2	3	4	5	6	7	8
1	H							
2	Li	Be	B	C	N	O	F	
3	Na	Mg	Al	Si	P	S	Cl	
4	K	Ca	–	–	–	–	–	Ti, V, Cr, Mn, Fe, Co, Ni
5	Cu	Zn	–	–	As	Se	Br	

Modern Periodic Table

The Modern Periodic Table of elements is a table of elements arranged systematically according to their increasing atomic numbers. It is a result of several modifications to Mendeleev's Periodic Table. The modifications were made as new elements were discovered and new theories developed to explain the chemical behaviour of elements. Note that, unlike the Mendeleev's Periodic Law, the elements in the Modern Periodic Table are listed in order of increasing atomic numbers.

Periodicity and general trends

In this section, you will learn about the periodicity and the general trends in the Periodic Table.

Periodicity

The Modern Periodic Law states that *'the properties of elements are a periodic function of their atomic numbers'*. The layout of the Periodic Table demonstrates recurring (periodic) chemical properties. The periodic recurrence of similar properties when elements are arranged according to their atomic numbers is called *periodicity*. For example, lithium is described as a very reactive metal, with one electron in the outermost shell. The eighth element after lithium is sodium. It is also a very reactive metal with one electron in its outermost shell. The eighth element after sodium is potassium, which is a very reactive metal as well, with one electron in its outermost shell.

Groups and periods

In the Periodic Table, the elements are arranged in groups and periods. The *groups* are the columns and the *periods* are the rows ofthe Periodic Table. Table 6.4 shows the first 20 elements in the Periodic Table. The full Periodic Table is given in Appendix 1.

Table 6.4: *Position and electronic arrangements of the first twenty elements in the Periodic Table*

Groups	I	II	III	IV	V	VI	VII	VIII/0
Period 1	$_{1}\text{H}$ 1							$_{2}\text{He}$ 2
Period 2	$_{3}\text{Li}$ 2:1	$_{4}\text{Be}$ 2:2	$_{5}\text{B}$ 2:3	$_{6}\text{C}$ 2:4	$_{7}\text{N}$ 2:5	$_{8}\text{O}$ 2:6	$_{9}\text{F}$ 2:7	$_{10}\text{Ne}$ 2:8
Period 3	$_{11}\text{Na}$ 2:8:1	$_{12}\text{Mg}$ 2:8:2	$_{13}\text{Al}$ 2:8:3	$_{14}\text{Si}$ 2:8:4	$_{15}\text{P}$ 2:8:5	$_{16}\text{S}$ 2:8:6	$_{17}\text{Cl}$ 2:8:7	$_{18}\text{Ar}$ 2:8:8
Period 4	$_{19}\text{K}$ 2:8:8:1	$_{20}\text{Ca}$ 2:8:8:2						

Groups

From Table 6.4, you can observe that there are elements with different numbers of electrons in the outermost shells. For example:

(i) one electron in their outermost shells for hydrogen, lithium, sodium and potassium.

(ii) two electrons in their outermost shells for beryllium, magnesium and calcium.

Elements with the same number of electrons in their outermost shells belong to the same group. For example, lithium, sodium and potassium have one electron in their outermost shells, therefore, they belong to Group I. The groups are usually indicated using Roman numerals; I, II, III, IV, V, VI, VII, and VIII (Figure 6.1).

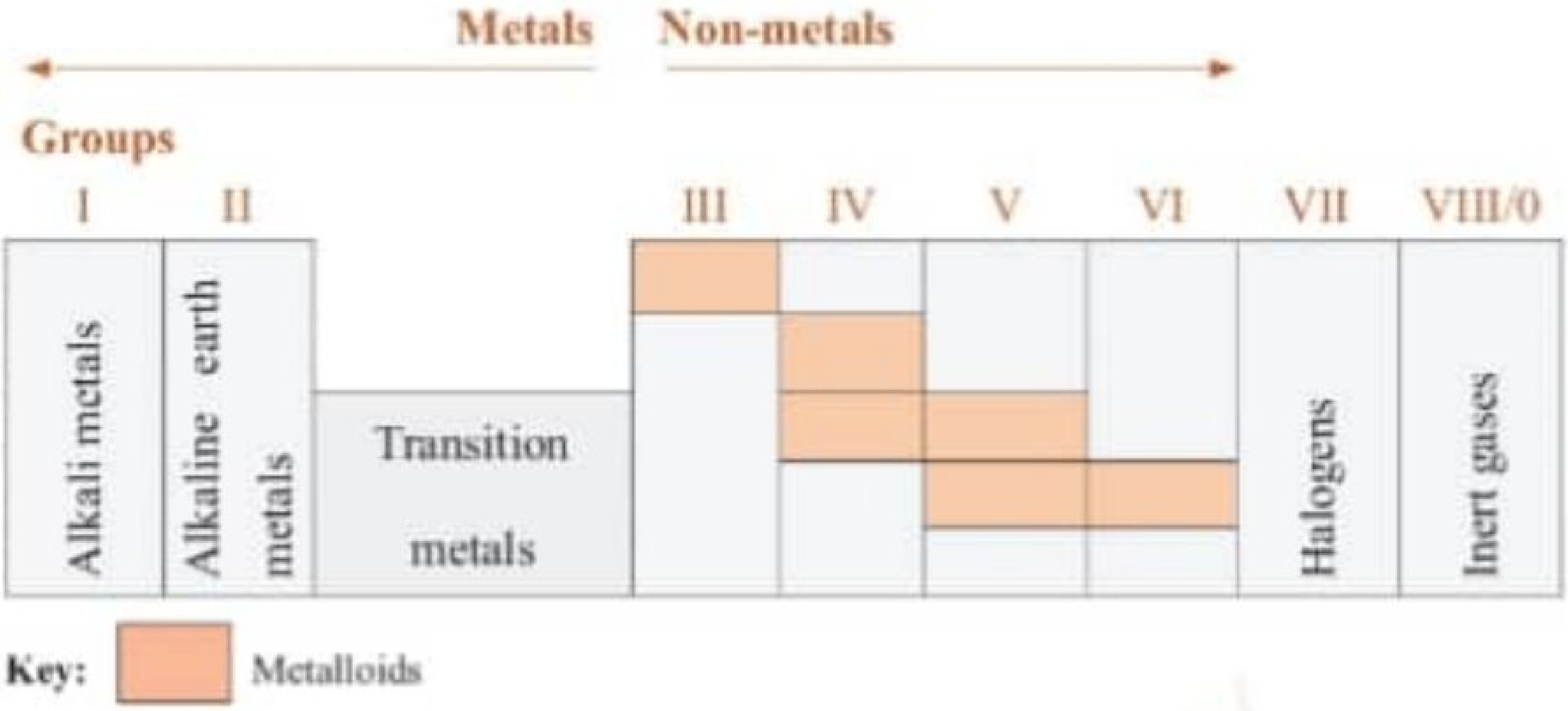


Figure 6.1: *Outline of the Periodic Table showing groups of elements*

It is important to note that the group number signifies the number of electrons in the outermost shell. For easy observation of such outermost number of electrons, you should write the electronic configuration of an element. The groups are numbered from left to right. The metals in Group I are called the *alkali metals* because they react with water to form alkaline solutions. Note that, even though hydrogen is placed in Group I, it is not an alkali metal. It carries some properties which are similar to those of group I elements and some which are similar to those of group VII elements. Therefore, some Periodic Tables place it above Group VII.

Group II metals are called *alkaline earth metals*. They have properties similar to those of Group I metals but they are less reactive compared to Group I metals. Group VII elements are called *halogens* ('salt formers' in Greek) because they react with metals to form compounds called salts. Group VIII elements are usually referred to as Group 0 elements. They are called *noble gases* and all their shells are completely filled with electrons. They were formerly called *inert gases* because they do not readily react to form compounds. The elements in the block between Group II and Group III are *transition elements*. These metals have high densities and melting points, form coloured compounds, and often act as catalysts.

On the basis of their general physical properties and chemical properties, nearly every element in the Periodic Table may be classified as either a metal or a non-metal. However, few elements tend to display both metallic and non-metallic characteristics. These are referred to as metalloids, and they include boron (B), silicon (Si), germanium (Ge), arsenic (As), antimony (Sb) and tellurium (Te). In some publications, germanium and antimony are usually classified as poor metals and the rest as non-metals.

Note: As solids, metals are more ductile than non-metals, whereas non-metals are more brittle than metals.

From Table 6.4, you can observe that elements with the same number of shells belong to the same period. The period number signifies the number of shells. Periods are numbered from 1 to 7. The electronic configuration entails the number of shells for each element. Periods and number of shells of the first twenty elements are shown in Table 6.5.

Table 6.5: *Periods and number of shells of the first twenty elements*

Period	Elements	Number of shells
Period 1	Hydrogen, helium	1
Period 2	Lithium, beryllium, boron, carbon, nitrogen, oxygen, fluorine, neon	2
Period 3	Sodium, magnesium, aluminium, silicon, phosphorus, sulphur, chlorine, argon	3
Period 4	Potassium, calcium	4

General periodic trends

Since elements in the Periodic Table have been placed in a systematic way, it is expected that there are trends within the periods and groups. The trends observed include variations in melting points, boiling points, densities, electronegativities, ionization energy, and atomic radii.

Melting point is the temperature at which a solid changes to form a liquid. Boiling point is the temperature at which a liquid changes to form a gas. Density is the degree of compactness of a substance, which means the mass per unit volume of a substance. *Electronegativity* is the ability or tendency of an atom to attract shared electrons towards itself. *Ionization energy* is the energy required to remove an electron from an atom or ion. *Atomic radius* is the distance between the nucleus of an atom and the outermost shell. *Reactivity* refers to how likely (or vigorously) an atom of a given element reacts with other substances.

Trends across periods

(i) The atomic radii of elements in a period decrease from left to right.

(ii) Elements on the left of the Periodic Table show metallic properties, while elements on the right show non-metallic properties.

(iii) Electronegativity increases from left to right.

(iv) The number of electrons and protons increase from left to right.

(v) The physical states of elements at room temperature (25 °C) vary from solid to gas.

General group trends

(i) Atomic radii increase down the group as successive shells are filled with electrons.

(ii) Densities increase down the group.

(iii) Melting points decrease down the group as the elements become more metallic in nature.

(iv) Electronegativity and ionization energy decrease down the group.

Trends in the groups

Group I: Alkali metals

Group I consists of five metals, namely lithium (Li), sodium (Na), potassium (K), rubidium (Rb), and caesium (Cs) as shown in the Table 6.6. Each of these elements has one electron in its outermost shell. Lithium, sodium and potassium react very readily with water or air, and are stored in oil.

Table 6.6: *Trends in Group I*

Name (Symbol)	Atomic number (z)	Electronic configuration	Atomic radius (picometres)	1st ionization energy (kJ /mol)	Melting point (°C)	Density (g/cm^3)	Electro-negativity
Lithium (Li)	3	2:1	152	526	180	0.54	1.0
Sodium (Na)	11	2:8:1	186	504	98	0.97	0.9
Potassium (K)	19	2:8:8:1	231	425	64	0.86	0.8
Rubidium (Rb)	37	2:8:18:8:1	244	410	39	1.5	0.8
Caesium (Cs)	55	2:8:18:18:8:1	262	380	29	1.9	0.7

Note: 1. Francium (Fr) is also an alkali metal, but is rarely included in the group. It is among the rarest naturally occurring elements.

2. The way electronic configurations of rubidium and caesium are written, is above the scope of this book. However, for the sake of understanding this section, the electronic configurations are written to show the period number, group number, and their relations to the trends.

The Group I elements have the following properties:

Physical properties

(i) They are good conductors of heat and electricity.

(ii) They are soft metals.

(iii) They have low density.

(iv) They have shiny surfaces when freshly cut.

Chemical properties

(i) They burn in oxygen or air with a characteristic flame colour to form white solid oxides. These oxides dissolve in water to form alkaline solutions of the metal hydroxides.

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

(ii) They react vigorously with water to give alkaline solutions and hydrogen gas.

$$\text{Metal} + \text{Water} \longrightarrow \text{Metal hydroxide} + \text{Hydrogen gas}$$

Group II: Alkaline earth metals

Group II consists of the following metals: beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra) as shown in Table 6.7. These elements have two electrons in their outermost shells. Magnesium and calcium are usually available in school laboratories.

Table 6.7: *Trends in Group II elements*

Name (Symbol)	Atomic number (z)	Electronic configuration	Atomic radius (picometres)	1st ionization energy (kJ/mol)	Melting point (°C)	Density (g/cm³)	Electro-negativity
Beryllium (Be)	4	2:2	112	899	14849	1280	1.5
Magnesium (Mg)	12	2:8:2	160	738	7730	651	1.2
Calcium (Ca)	20	2:8:8:2	197	590	4741	851	1.0
Strontium (Sr)	38	2:8:18:8:2	215	549	4207	800	1.0
Barium (Ba)	56	2:8:18:18:8:2	217	503	3420	850	0.9

Note: The way electronic configurations of strontium and barium (Table 6.7) are written, is above the scope of this book.

Group II elements have the following properties:

Physical properties

(i) They are harder metals than those in Group I.

(ii) They are silvery grey in colour when pure and clean. However, they tarnish quickly when left in air due to the formation of the respective metal oxides.

(iii) They are good conductors of heat and electricity.

Chemical properties

(i) They burn in oxygen or air with a characteristic flame colour to form a solid white oxide.

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

(ii) They react with water but much less vigorously than the elements in Group I.

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

(iii) The reactivity of metals increases down the group. For example, the reaction of calcium with water is vigorous, while that of magnesium with water is very slow.

Task

In groups, draw the Modern Periodic Table. You can make use of manila paper, coloured marker pens, ruler, and your Chemistry textbook. Using a different marker for each group of the Periodic Table, fill in the block with the respective chemical symbols for the elements. For each element, include the atomic number, relative atomic mass, and electronic configuration. The neatest and most accurate chart should be pinned on the classroom noticeboard.

Activity

Aim: To demonstrate the differences in reactivity of calcium and magnesium with water.

Requirements: Two test tubes, magnesium, calcium, distilled water, and measuring cylinder

Procedure

1. Transfer about 3 cm^3 of distilled water in a test tube.
2. Add a small amount of calcium (spatulaful) straight from the container in which it is stored. Record your observation.
3. Repeat steps 1 and 2 using a clean piece of magnesium ribbon. Record your observations.

Questions

1. What happens when a piece of calcium is dropped in water?
2. What happens when the magnesium ribbon is dropped in water?
3. Comment on the reactivities of calcium and magnesium.

Chapter summary

1. The Periodic Table of elements is a method of displaying chemical elements in a table format. It was developed after several modifications to the Mendeleev's Periodic Table.
2. Mendeleev's Periodic Law states that *"the properties of elements are periodic functions of their relative atomic masses"*.
3. The Modern Periodic Law states that *"the properties of elements change systematically according to their atomic numbers"*.
4. Periodicity refers to the regular periodic changes of properties of elements due to changes in atomic numbers.
5. Elements with the same number of electrons in their outermost shells belong to the same group.
6. The group number signifies the number of electrons in the outermost shell of an element.
7. Elements with the same number of shells belong to the same period.
8. The period number signifies the number of shells.
9. Electronic configuration of an element entails the electronic arrangements in the shell(s), number of shells, and the group to which the element belongs.

Revision exercise 6

1. Write the electronic configuration and indicate the *group* and *period* for each of the following elements in the table:

Element	Electronic configuration	Group	Period
Carbon			
Phosphorus			
Beryllium			
Oxygen			
Lithium			
Fluorine			
Silicon			
Magnesium			
Aluminium			
Potassium			
Nitrogen			

2. Choose the correct answer in each of the following items:

   (i) Non-metals are ________ than metals.
      (a) better conductors of electricity
      (b) more brittle
      (c) more ductile
      (d) better conductors of heat

   (ii) The electronic arrangement of an element is 2:3. This element is in _______ of the Periodic Table.
      (a) Group 2
      (b) Group 8
      (c) Period 3
      (d) Period 2

(iii) Which of the following statements does not describe the alkaline earth metals?

- (a) They burn in oxygen to form solid white oxides.
- (b) They become less reactive down the group.
- (c) They are good conductors of heat and electricity.
- (d) They react with water.

(iv) Mendeleev classified elements on the basis of

- (a) mass number.
- (b) atomic number.
- (c) proton number.
- (d) neutron number.

(v) An element T with electronic configuration 2:8:3 belongs to group ___ and period __ of the Periodic Table.

- (a) III and 3
- (b) III and 2
- (c) II and 3
- (d) II and 2

(vi) Identify the electronic configuration of an element with 16 electrons.

- (a) 2:8:3
- (b) 2:8:4
- (c) 2:8:6
- (d) 2:10:4

(vii) Hydrogen is placed in group I elements because

- (a) it is a metal.
- (b) it loses an electron.
- (c) it is an inert gas.
- (d) its atomic number is 1.

(viii) The list of elements belonging to alkali metals includes;

- (a) sodium, potassium, and lithium.
- (b) magnesium, potassium, and sodium.

(c) magnesium, lithium, and sodium.
(d) magnesium, calcium, and beryllium.

(ix) Group I elements burn in oxygen to form
(a) metal oxides.
(b) non-metal oxides.
(c) hydroxides.
(d) carbon dioxides.

(x) The electronegativities of elements __________ across the periods.
(a) decrease
(b) increase
(c) remain the same
(d) fluctuate

3. Write **TRUE** for a correct statement and **FALSE** for an incorrect statement.
(a) Lithium and sodium belong to the same period in the Periodic Table.
(b) Mendeleev classified elements according to the Law of Octaves.
(c) Helium and argon belong to the same group.
(d) Beryllium, magnesium and aluminium are all alkaline earth metals.
(e) Helium, neon and argon have their shells completely filled with electrons.
(f) The number of protons increase from left to right across the periods of the Periodic Table.
(g) The Group I metals of the Periodic Table are harder than those of Group II.
(h) For any group in the Periodic Table, the densities of the elements increase down the group.

4. Briefly explain the following terms:
(a) Periodicity
(b) Halogens
(c) Periodic Law
(d) Alkali metals
(e) Electronegativity
(f) Ionization energy

(g) Transition elements

(h) Alkaline earth metals

5. A particular metal reacts slowly with water to give a strong alkaline solution. In which group of the Periodic Table would you place it?

6. (a) What are noble gases?

   (b) In which group of the Periodic Table do noble gases belong?

   (c) What is common about the noble gases regarding the following properties?

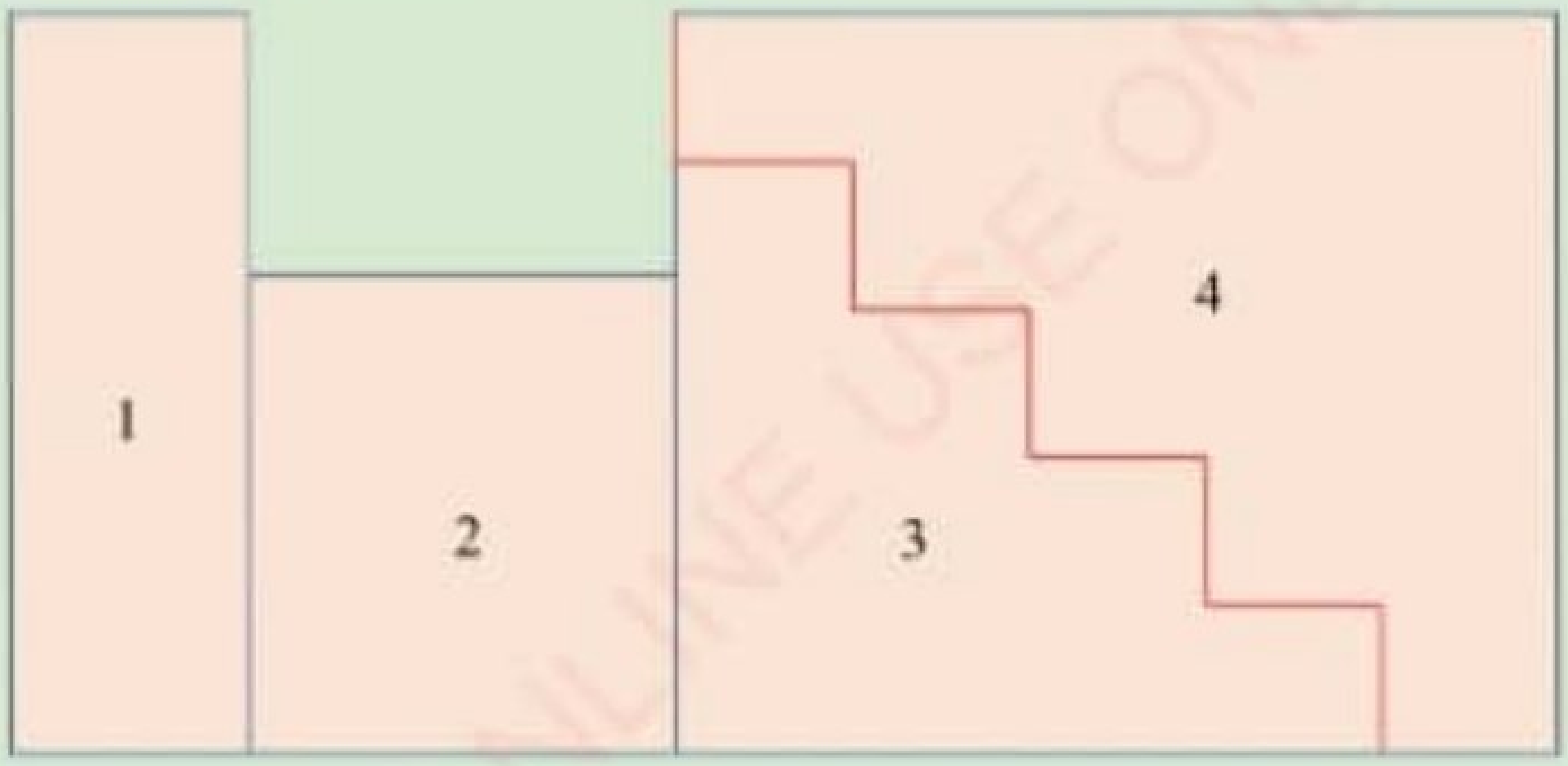
      (i) Electronic arrangements

      (ii) Chemical reactions

7. (a) Why are certain elements in the Periodic Table referred to as metalloids?

   (b) Give three examples of metalloids.

8. The following diagram represents the Periodic Table with four areas denoted by 1, 2, 3 and 4:



   (a) Which area is most likely to contain non-metals?

   (b) Which area is most likely to contain elements whose oxides dissolve in water?

   (c) Which area contains transition elements?

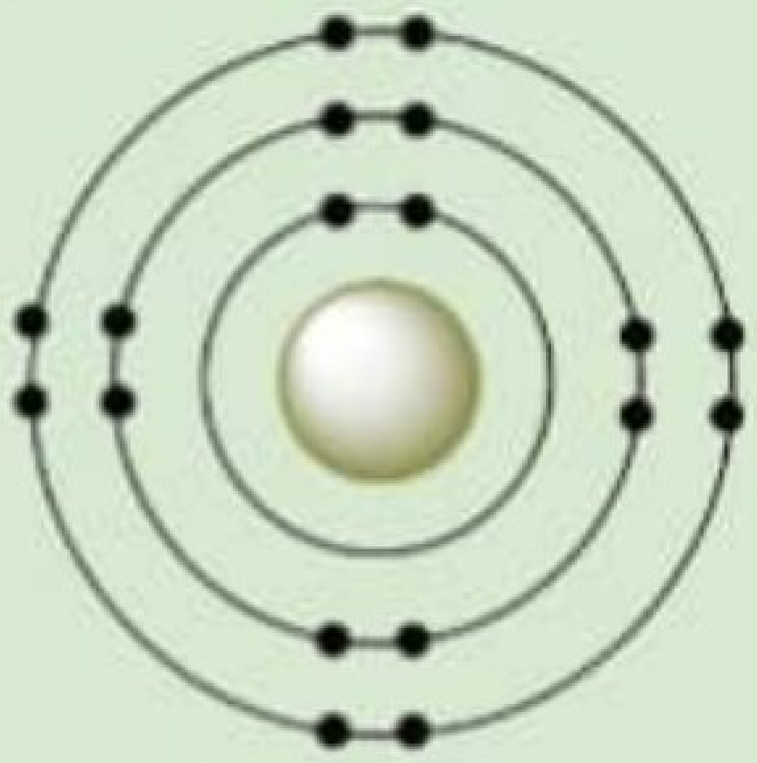
   (d) Which area is most likely to contain elements with both metallic and non-metallic characteristics?

9. (a) Given the elements calcium, sulphur, chlorine, helium and neon, write down their:
   - (i) period numbers.
   - (ii) group numbers.
   - (iii) atomic numbers.
   - (iv) number of electrons in one atom.
   - (v) electronic configurations.

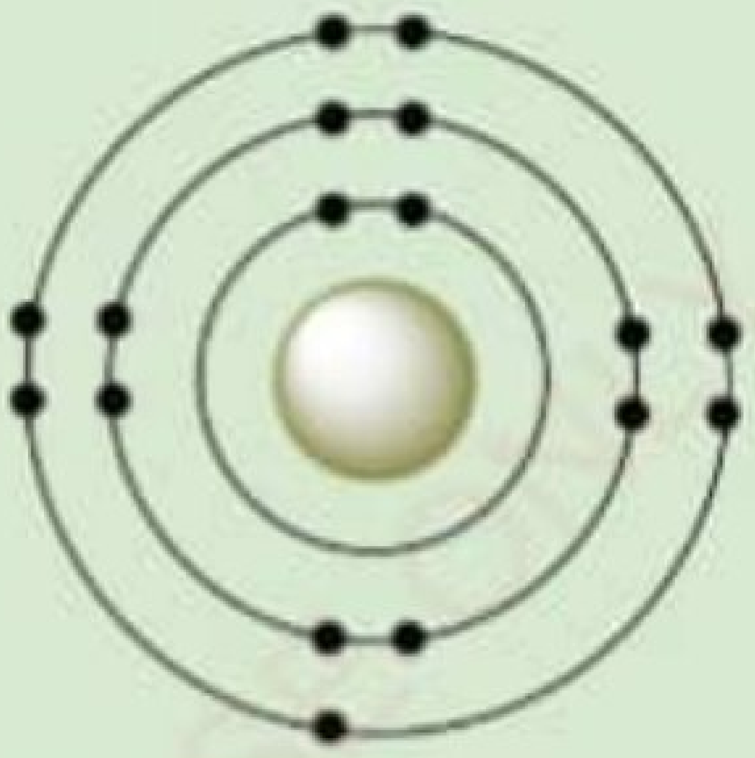
   (b) Which of the above elements would you expect to have similar properties? Give reasons.

10. Element R belongs to Period 3 and Group VI in the Periodic Table.
   - (a) Draw its atomic structure.
   - (b) Determine its atomic number.

11. The following diagrams show the structures of atoms of elements X and Y:



X



Y

   - (a) Which one is stable?
   - (b) Which of the two elements can conduct electricity?
   - (c) Which of the two elements is chemically more reactive? Explain.
   - (d) Identify elements X and Y.

12. Why do the thermal conductivities and electrical conductivities of elements in the Periodic Table decrease across the periods?

13. Write any three physical properties of alkali earth metals.

14. What happens when alkali earth metals burn in oxygen?

15. Consider elements $_{11}T$ and $_{17}Q$. Which of the two elements is more electronegative? Explain.

Chapter Seven

Chemical bonding, formula and nomenclature

Introduction

Bonding involves holding atoms together to form molecules or compounds. A chemical formula refers to symbols and numbers that represent the composition of a certain chemical substance. Nomenclature means naming. In this chapter, you will learn about chemical bonding, valencies, chemical formulae, oxidation states, radicals, covalent bonding, electrovalent bonding, and nomenclature of chemical substances. The competencies developed will enable you to identify the chemical substances used in daily life activities that are formed when different materials are bonded together. This will enable you to study the relationships existing among chemical substances and how to apply them in the learning and in different activities related to Chemistry.

Bonding

A *bond* is anything that holds two or more substances together. Many things used in our daily lives are constructed using different materials joined together by some bonds. For example, in a brick wall (Figure 7.1), each brick is joined to the other by a bond made of mortar. Similarly, chemical substances are made of atoms that are held together by chemical bonds.

Figure 7.1: *Bricks bonded using mortar to form a wall*

A *chemical bond* is a force of attraction that holds atoms or ions together to form molecules or compounds. The bond may result from forces of attraction between oppositely charged ions or through the sharing of electrons.

A *molecule* is the smallest particle of an element or compound which can normally exist separately.

Chemical bonding

Chemical bonding involves electrons in the outermost shells of atoms. When the outermost shells are completely filled with electrons, the atoms are said to be stable, otherwise they are unstable. Table 7.1 shows the electronic arrangements of some elements and the stability of atoms of the elements.

Table 7.1: *Electronic arrangements and stability of the first 20 elements of the Periodic Table*

Element	Number of electrons	Electronic arrangement	Stability of an atom
Hydrogen	1	1	Unstable
Helium	2	2	Stable
Lithium	3	2:1	Unstable
Beryllium	4	2:2	Unstable
Boron	5	2:3	Unstable
Carbon	6	2:4	Unstable
Nitrogen	7	2:5	Unstable
Oxygen	8	2:6	Unstable
Fluorine	9	2:7	Unstable
Neon	10	2:8	Stable
Sodium	11	2:8:1	Unstable
Magnesium	12	2:8:2	Unstable
Aluminium	13	2:8:3	Unstable
Silicon	14	2:8:4	Unstable
Phosphorus	15	2:8:5	Unstable
Sulphur	16	2:8:6	Unstable
Chlorine	17	2:8:7	Unstable
Argon	18	2:8:8	Stable
Potassium	19	2:8:8:1	Unstable
Calcium	20	2:8:8:2	Unstable

From Table 7.1, only helium, neon, and argon have stable electronic arrangements. Helium has a maximum number of 2 electrons in its outermost shell. Neon and argon have a maximum of 8 electrons in each of their respective outermost shells. These types of stable atoms are generally unreactive and can exist freely as single atoms. Unstable atoms cannot exist freely as single atoms. For unstable atoms to become stable, they should acquire electronic arrangements similar to those of noble gases. This means that they can either lose, gain or share electrons through chemical bonding.

Formation of ions

When an atom loses an electron to acquire a stable electronic arrangement, the lost electron is transferred to another atom which then becomes stable. The resulting species become *ions*. The one that loses an electron becomes a *positively charged ion*, while the one that gains an electron becomes a *negatively charged ion*. The number of electrons gained or lost will be equal to the charge of an ion. The positively charged ion is called a *cation* and the negatively charged ion is called an *anion*. Electrons are negatively charged while protons are positively charged; thus, the charge of the ion is due to the unbalanced number of electrons and protons. For example, a sodium atom has eleven protons and eleven electrons. Its charge is 0 because (+11) + (–11) = 0. After sodium loses one electron it will have 10 electrons. Its charge will become +1 because (+11) + (–10) = +1 (Figure 7.2).

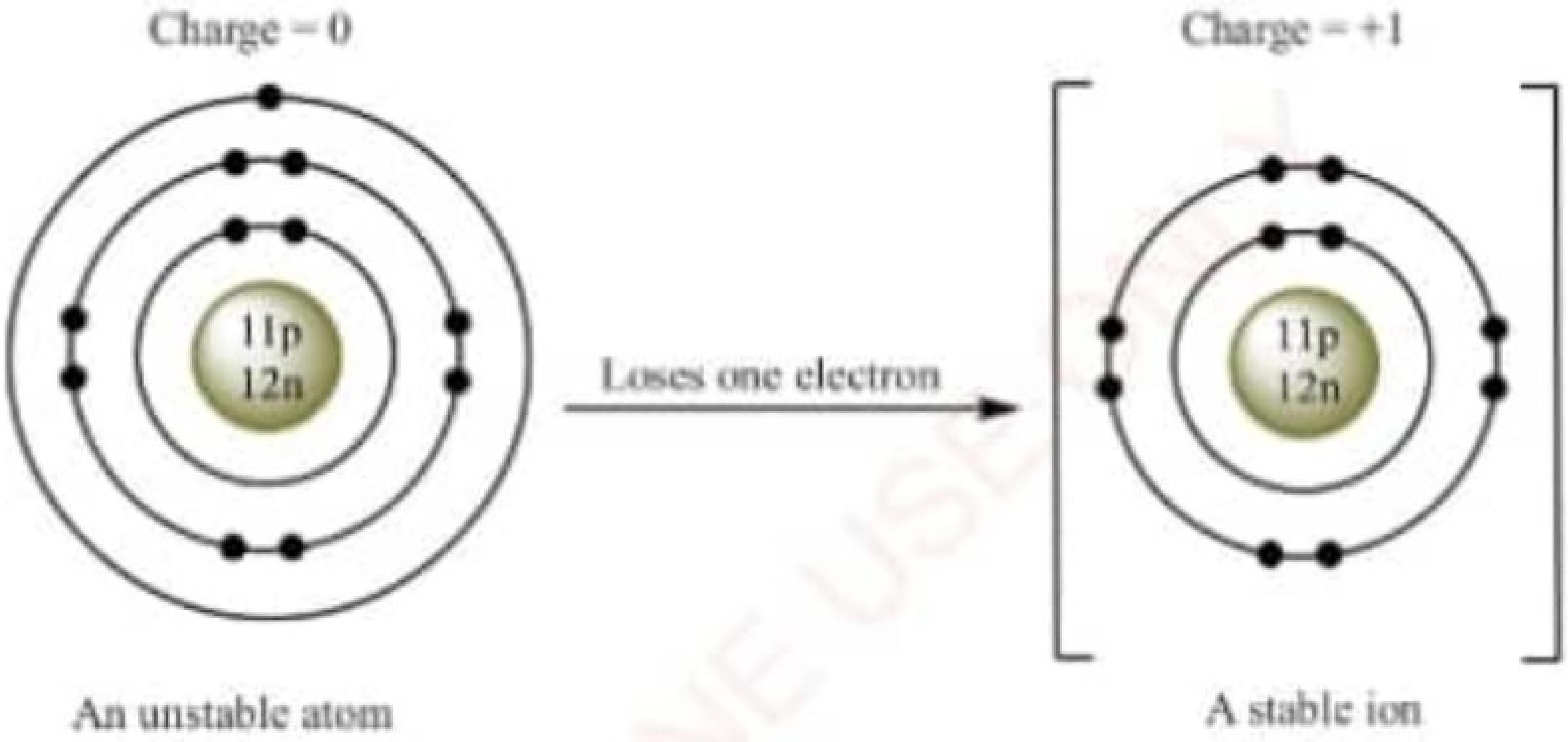


Figure 7.2: *Sodium atom and its ion*

Electrovalent bonding

Electrovalent (ionic) bonding is the attraction between ions of opposite charges. When an atom loses an electron, it becomes a positively charged ion, while the atom that accepts the electron becomes a negatively charged ion. These charged ions attract each other due to their opposite charges.

Electrovalent bonding usually occurs between *a metal* and *a non-metal*. The metal loses electron(s) and the non-metal gains electron(s). For example, when sodium and chlorine react to form sodium chloride, the sodium atom must lose an electron to acquire a stable noble gas structure of neon, while the chlorine atom must gain an electron to acquire a stable noble gas structure of argon. This results in a positively charged sodium ion and a negatively charged chloride ion which attract each other to form a sodium chloride crystal as shown in Figure 7.3. For the sake of this chapter, electrons are represented by dots with different colours.

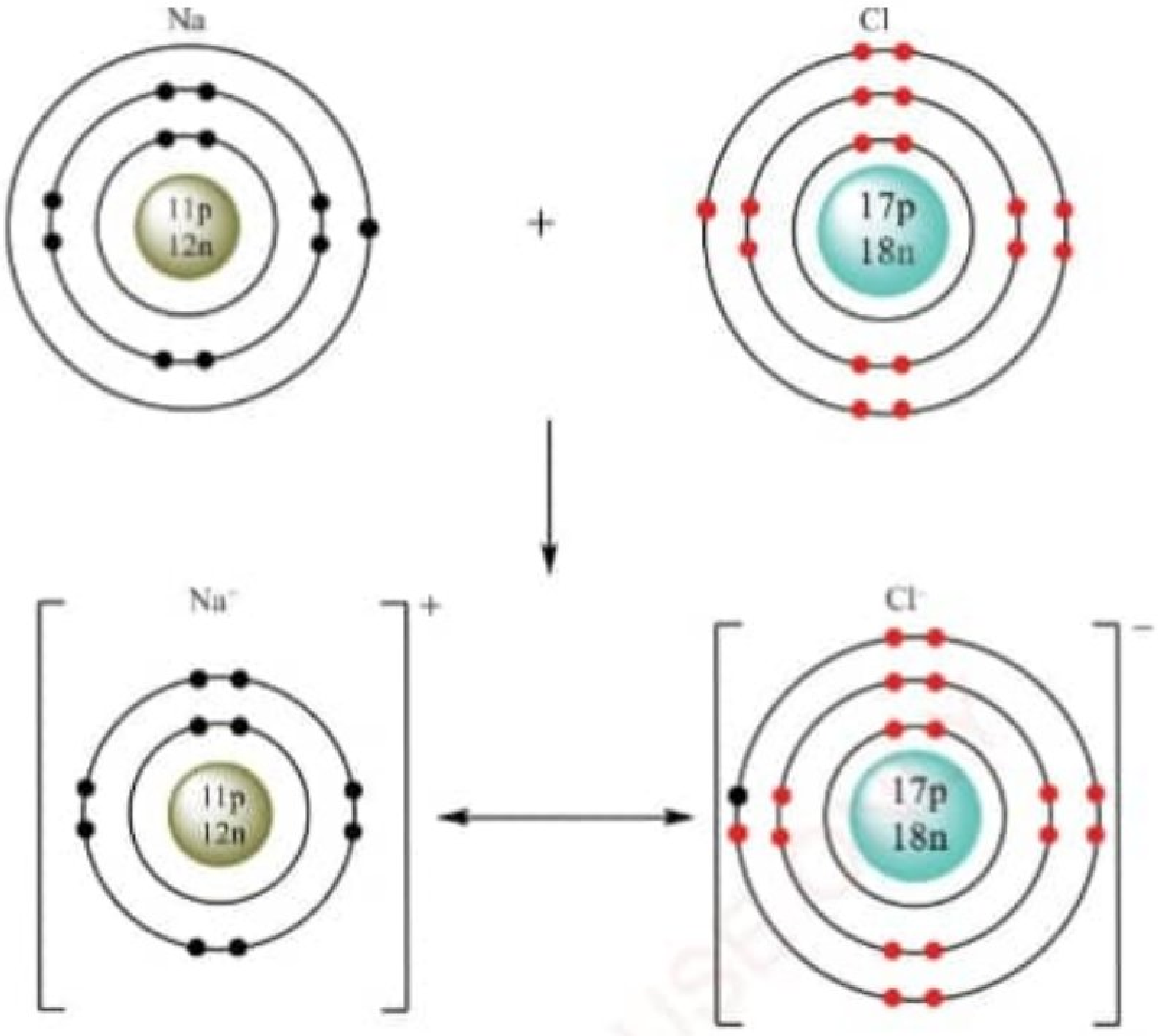


Figure 7.3: *Electrovalent bonding in sodium chloride*

Molecules of electrovalent compounds are *not discrete*. This means that the electrons are not localized or attached to particular ions of opposite charges to form pairs. Instead, a group of cations surround an anion and a group of anions surround a cation. This means that, ions can move freely around each other, especially when in molten or in solution forms.

Properties of electrovalent compounds

Electrovalent (ionic) compounds have the following properties:

(i) They are generally soluble in water.

(ii) They conduct electricity in solution or molten forms, but not in their solid forms.

(iii) They are usually crystalline solids at room temperature.

(iv) They have high melting and boiling points.

(v) They are generally insoluble in non–polar solvents such as carbon tetrachloride and hexane. Non–polar solvents contain bonds between atoms with similar electronegativities.

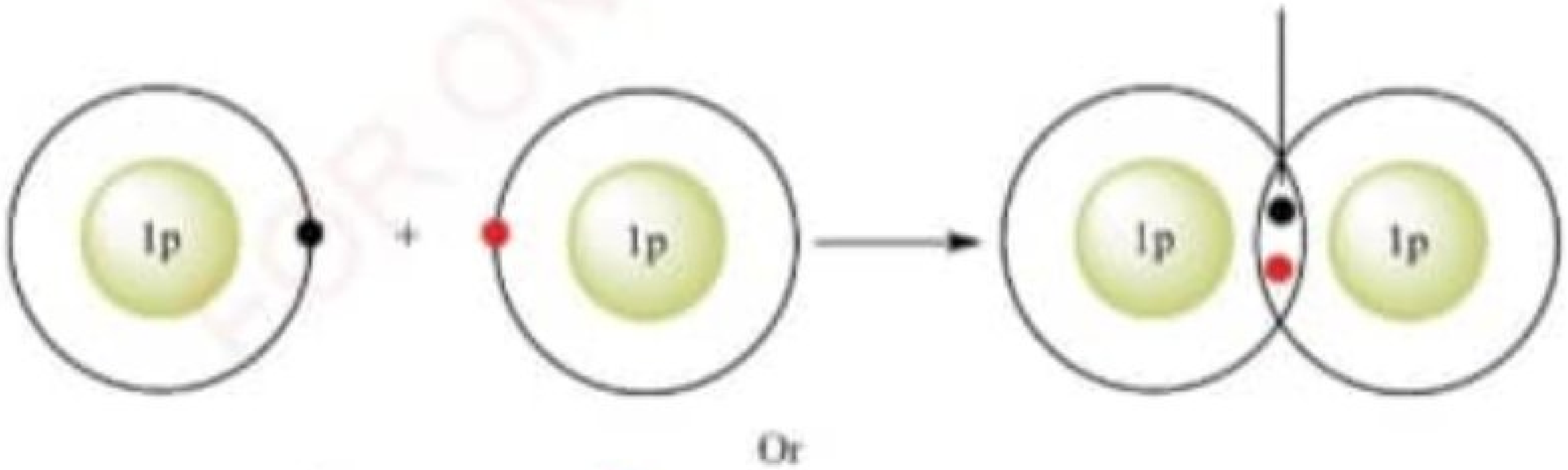
Exercise 7.1

1. Explain why cations are slightly smaller than their neutral atoms, whereas anions are slightly larger than their neutral atoms.
2. Magnesium and oxygen atoms combine to form magnesium oxide.
    (a) What is the charge on the magnesium ion in the oxide?
    (b) What is the charge on the oxide ion?
    (c) Illustrate the electron transfer using dots.

Covalent bonding

Covalent bonding involves the sharing of electrons between atoms of the same or different elements. Atoms may gain the noble gas electronic structures without becoming ions but by sharing of the outermost electrons. For example, a hydrogen atom needs one electron to acquire the helium electronic configuration. It can combine with another hydrogen atom and share the electrons so that each acquires a stable helium configuration. This results in the formation of a hydrogen molecule (Figure 7.4). The kind of the bond formed between the two atoms is called a *covalent bond*. A covalent bond is a chemical bond formed by the sharing of one or more electrons between atoms.

The two electrons are shared to form one covalent bond



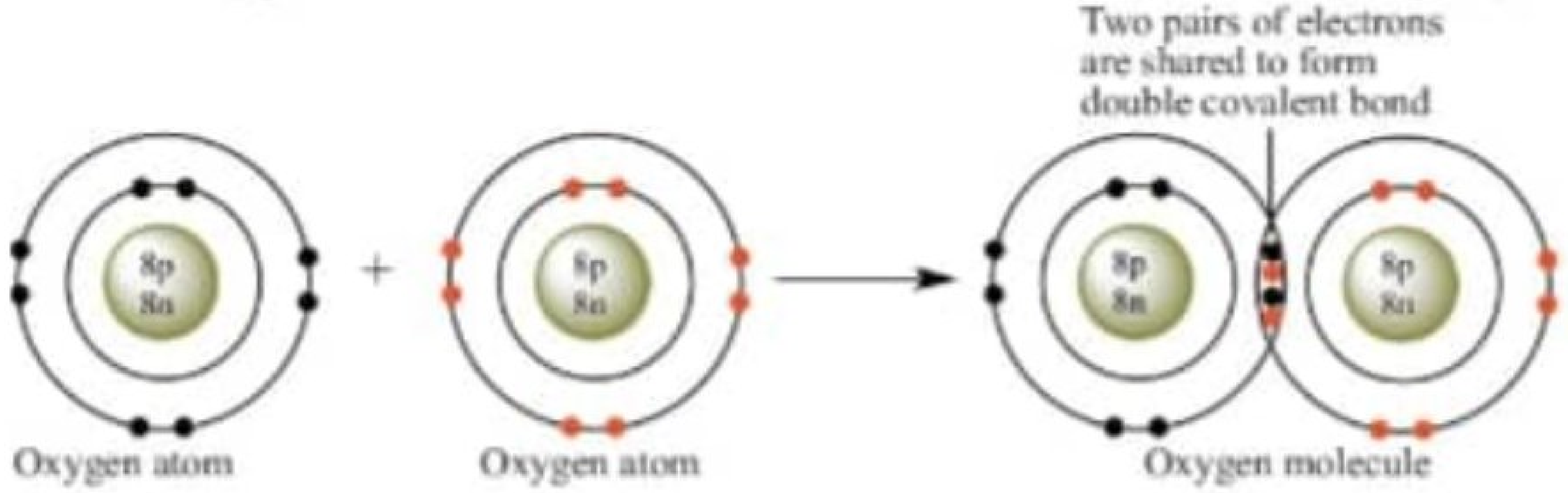
Or

$$H^{\bullet} + H^{\bullet} \longrightarrow H{:}H \text{ or } H-H$$

Figure 7.4: *Covalent bonding in a hydrogen molecule*

Note: A single line between two atoms indicates a single covalent bond carrying an electron pair. Two lines are used to represent bonding of two pairs, three lines represent three pairs, and so on.

Oxygen requires two electrons to acquire the stable atomic structure. Thus, two atoms of oxygen combine and share four electrons between them (Figure 7.5).



Or, considering only the outermost shells:

Figure 7.5: Double covalent bond in oxygen

Covalent bonding also occurs between atoms of different elements such as between hydrogen and chlorine. Hydrogen requires one electron to attain a stable helium electronic configuration, while chlorine requires one electron to acquire the argon electronic configuration. The two atoms combine by sharing two electrons between them (Figure 7.6).

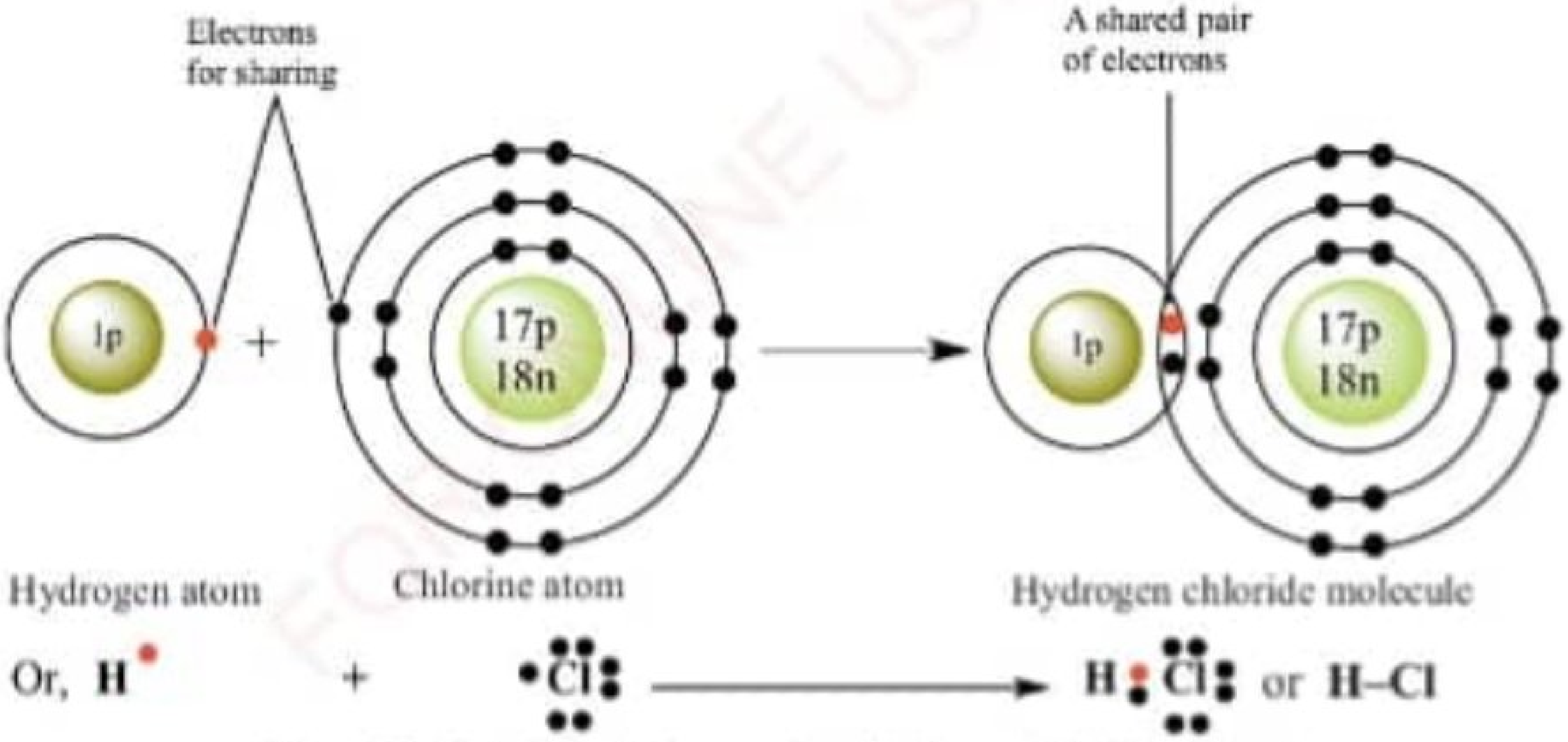


Figure 7.6: *Sharing of electrons in a hydrogen chloride molecule*

The same happens to the combination between oxygen and hydrogen. Oxygen requires two electrons to acquire a stable structure, while hydrogen requires one electron. Thus, one oxygen atom combines with two hydrogen atoms as shown in Figure 7.7, whereby each hydrogen provides one electron for sharing with oxygen.

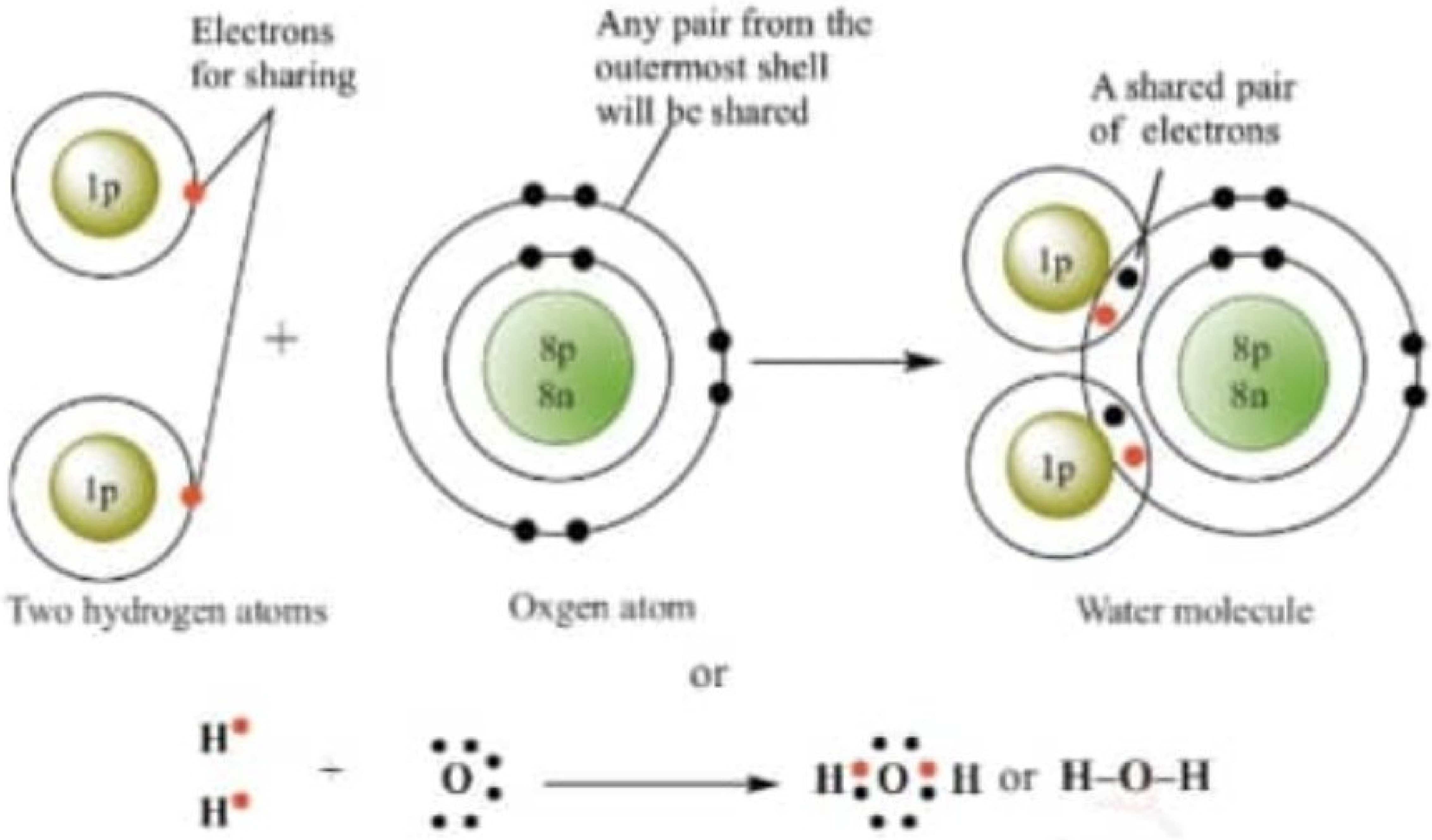


Figure 7.7: *Covalent bonding in a water molecule*

Covalent bonds are usually formed between non-metals. Molecules of covalent compounds are discrete or distinct. This means that the atoms forming the molecules cannot exist freely in the compounds. They remain bound together in molecules and their electrons are not free to form bonds with other atoms.

Properties of covalent compounds

The properties of covalent compounds include the following:

(i) Their melting and boiling points are usually low.

(ii) They are usually liquids or gases at room temperatures.

(iii) They do not conduct electricity.

(iv) They are generally soluble in non-polar solvents such as carbon tetrachloride and hexane.

(v) They are generally insoluble in water.

Activity 7.1

Aim: To illustrate the formation of covalent bonds.

Requirements: Manila paper, pens, coloured markers and a pair of compasses

Procedure

1. In groups, illustrate using circles and dots the covalent bonds in methane (CH_4), ammonia (NH_3) and carbon dioxide (CO_2).
2. In each case, indicate the number of covalent bonds present and the pairs of electrons involved.

Question

What determines the number of covalent bonds that an atom can form?

Valency

Valency refers to the ability of an atom of a given element to combine with other atoms, and is measured by the number of electrons that the atom will donate, receive or share to form a chemical bond. It is the *combining power/capacity* of an element or a radical. The combining capacity of an atom of a given element is determined by the number of hydrogen atoms it combines with. For example, the valency of chlorine is 1 because one atom of hydrogen combines with one atom of chlorine to form hydrogen chloride (HCl). The valency of zinc is 2 because two atoms of hydrogen are displaced from dilute acids by one atom of zinc.

It is easy to predict the valencies of elements from the Periodic Table. Group I elements have one electron in their outermost shells, and so, their valency is 1. Group II elements have two electrons in their outermost shells, hence their valency is 2. Valencies are rarely above 4. For elements with more than four electrons in the outermost shells, the valency number is usually obtained by subtracting the number of electrons from eight. For example, sulphur with six electrons in the outermost shell has a valency of $8 - 6 = 2$, and the electrons found in it are called *valence electrons*. Some elements have more than one valency. For example, iron has valencies of 2 and 3, copper has valencies of 1 and 2, lead has the valencies of 2 and 4, and manganese has valencies of 2, 4 and 7. Table 7.2 shows the valencies of some elements and radicals.

Radicals

A *radical* is a group of atoms which behaves as a single unit and has a positive or negative charge. It contains at least one unpaired electron. Such a group maintains its identity throughout any chemical reaction. Radicals can gain or lose electrons to form ions. Most radicals form the non-metallic part of a compound, so their ions are negatively charged. Examples are CO_3^{2-} and SO_4^{2-} ions. An exception is for the ammonium radical, NH_4^+, which behaves like the metallic part of a compound and forms a positive ion. The valency of the radical is the same as the numerical value that the group acquires when it loses or gains an electron to form an ion. The common radicals with their formulae and valencies are shown in Table 7.2.

Table 7.2: *Valencies of some elements and radicals*

Category	Valency 1		Valency 2		Valency 3	
	Element	*Ion/ radical*	*Element*	*Ion/ radical*	*Element*	*Ion/ radical*
Metals	Potassium (K)	K^+	Barium (Ba)	Ba^{2+}	Aluminium (Al)	Al^{3+}
	Silver (Ag)	Ag^+	Calcium (Ca)	Ca^{2+}		Fe^{3+}
	Sodium (Na)	Na^+	Iron (Fe)	Fe^{2+}	Iron (Fe)	
			Lead (Pb)	Pb^{2+}		
			Magnesium (Mg)	Mg^{2+}		
			Mercury (Hg)	Hg^{2+}		
			Zinc (Zn)	Zn^{2+}		
Non-metals	Chlorine (Cl)	Cl^-	Oxygen (O)	O^{2-}		
	Hydrogen (H)	H^+	Sulphur (S)	S^{2-}		
	Fluorine (F)	F^-				
Radicals	*Ammonium radical	NH_4^+	Carbonate	CO_3^{2-}	Phosphate	PO_4^{3-}
			Dichromate	$Cr_2O_7^{2-}$		
	Chlorate	ClO_3^-	Sulphate	SO_4^{2-}		
	Cyanide	CN^-	Sulphite	SO_3^{2-}		
	Hydroxide	OH	Thiosulphate	$S_2O_3^{2-}$		
	Hydrogencarbonate	HCO_3^-				
	Hydrogensulphate	HSO_4^-				
	Nitrate	NO_3^-				
	Nitrite	NO_2^-				
	Permanganate	MnO_4^-				

Note: Ammonium radical (NH_4^+) has a valency of 1 and can react like metals. Its compounds are similar to those of group I elements.

Role play

Assume each of your hands is a valence electron that can be used for covalent bond formation. Participate in a role play with your fellow students to construct the following molecules:

(a) O_2 (b) SO_2

Oxidation state

Oxidation state (also called *oxidation number*) is the total number of electrons that an atom either gains or loses in order to form a chemical bond with another atom. It is the measure of the electron control that an atom has in a compound compared to the atom in the pure element. The neutral atom has no charge. The following are the rules used to assign oxidation states of the elements:

1. The oxidation number of free elements is zero. For example, all elements in the Periodic Table have oxidation number of zero.
2. The sum of the oxidation states of all atoms forming a molecule or ion is the net charge of that species. For example, nitrogen (N_2), hydrogen (H_2) and oxygen (O_2) molecules have the oxidation number of zero.
3. In simple ions that consist of only one atom, the oxidation number is equal to the charge on the ion. For example, the oxidation number of a sodium ion (Na^+) is +1, aluminium (Al^{3+}) is +3, iron(II) (Fe^{2+}) is +2, and iron(III) (Fe^{3+}) is +3. In an oxide ion (O^{2-}), the oxidation number of oxygen is –2.
4. In their compounds, Group I metals have an oxidation number of +1. Group II metals have an oxidation number of +2, while, Group III metals have an oxidation number of +3.
5. In their compounds, halogens always have an oxidation number of –1.
6. Hydrogen has an oxidation state of +1 in most compounds. The exception is in hydrides of active metals where the oxidation number is –1. For example, the hydrogen atom gains an electron from the lithium atom in lithium hydride (LiH).
7. Oxygen has an oxidation state of –2 when present in most compounds, except:
    (i) in peroxides, e.g. H_2O_2, where the oxidation number is –1.
    (ii) when bonded with fluorine to form F_2O, the oxidation number is +2.

All oxidation numbers must be consistent with the conservation of charge.

This means that for all neutral molecules, the oxidation number of all the atoms must add up to zero. For example, in H_2O, two hydrogen atoms each of charge +1 combine with one oxygen atom of charge –2. The charge of the H_2O molecule is +2 – 2 = 0.

Note: There is a close relationship between valency and oxidation state, however, they are not the same. Valency is a fixed value, but oxidation state is an arbitrary value (it may vary).

Example 7.1

Find the oxidation state of chlorine in $KClO_3$

Solution

The oxidation number of potassium is +1

The oxidation number for oxygen is –2

For the three oxygen atoms, the oxidation number is $(-2 \times 3) = -6$

$KClO_3$ is a neutral compound, and therefore, the oxidation number of the compound is zero.

Therefore, $+1 + Cl - 6 = 0$

$$Cl = 6 - 1 = +5$$

The oxidation number of chlorine in $KClO_3$ is +5

Example 7.2

Find the oxidation number of sulphur in SO_4^{2-}

Solution

The total charge on the sulphate ion is –2

The oxidation number of oxygen is –2

Therefore, $S + (-2 \times 4) = -2$

$$S - 8 = -2$$

$$S = 8 - 2$$

$$S = +6$$

The oxidation state of sulphur in SO_4^{2-} is +6

Example 7.3

Give the oxidation number of Cr in $Cr_2O_7^{2-}$

Solution

Total charge on the dichromate ion is –2

For oxygen, $-2 \times 7 = -14$

Therefore, $2Cr - 14 = -2$

$$2Cr = +12$$
$$Cr = +6$$

Therefore, the oxidation number of Cr is +6

Exercise 7.2

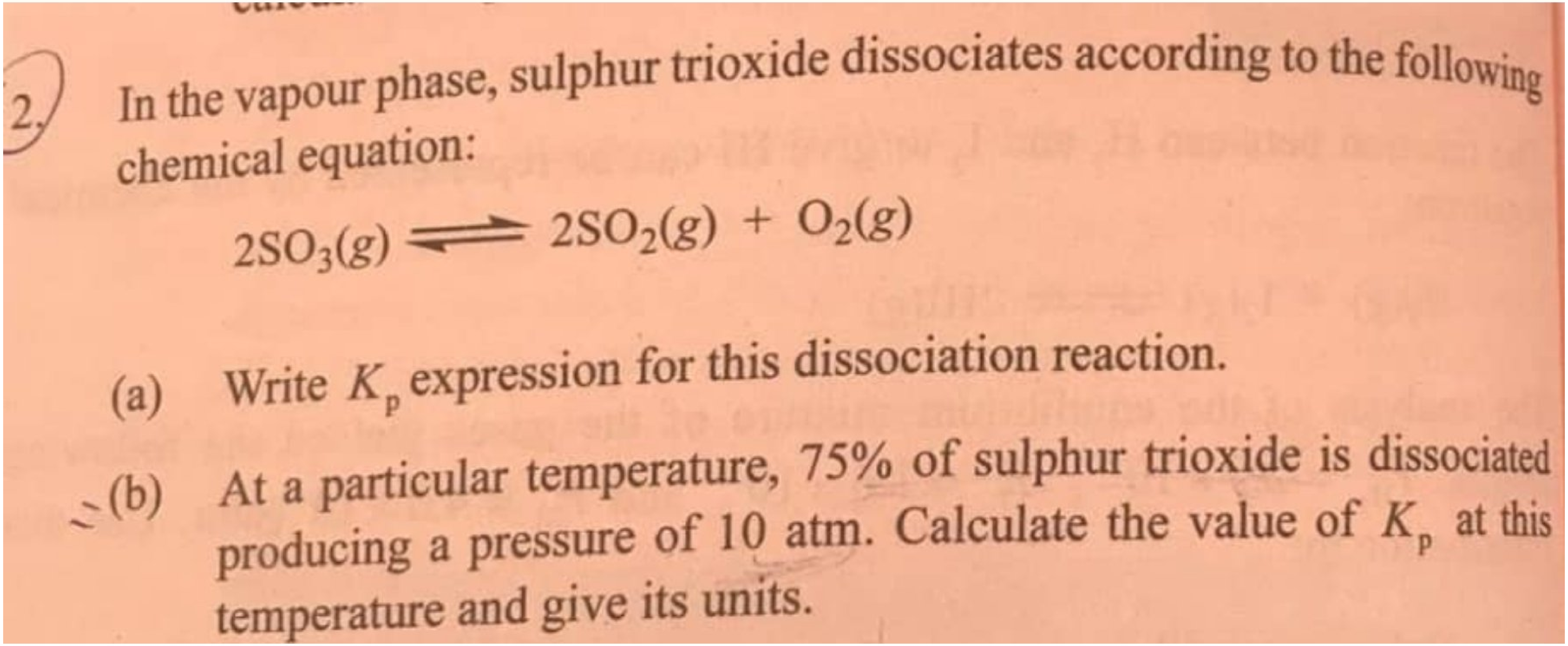
Calculate the oxidation number of each of the underlined elements in the following chemical substances:

(a) $Na_3\underline{P}O_4$ (b) $\underline{S}O_3^{2-}$ (c) $\underline{N}O_2^-$

Chemical formulae

A chemical formula is a representation that uses symbols to show the proportions of the elements present in a chemical compound. The number of atoms or groups of atoms are shown by number subscripts. For example, the chemical formula for sodium sulphide is Na_2S, which clearly shows that two atoms of sodium combine with one atom of sulphur to form the molecule of sodium sulphide. For groups of atoms (i.e. radicals), a bracket is used to show that they are being considered as a unit under one valency. For example, in calcium nitrate, $Ca(NO_3)_2$ the NO_3^- radical is in brackets. There are some points to remember when writing chemical formulae.

1. Positively charged ions (cations) are written before the negatively charged ions (anions).
2. A radical must be treated as a unit.
3. The name of the cation is the same as the neutral element from which it is derived (e.g. Na^+ is sodium ion).
4. The mono-atomic anions are named by adding *-ide* to the root of the name of the non-metal that forms the anion. For example, Cl^- is chloride.



2. In the vapour phase, sulphur trioxide dissociates according to the following chemical equation:

$$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$$

(a) Write K_p expression for this dissociation reaction.

(b) At a particular temperature, 75% of sulphur trioxide is dissociated producing a pressure of 10 atm. Calculate the value of K_p at this temperature and give its units.

Polyatomic anions (negatively charged radicals) follow different patterns depending on their compositions.

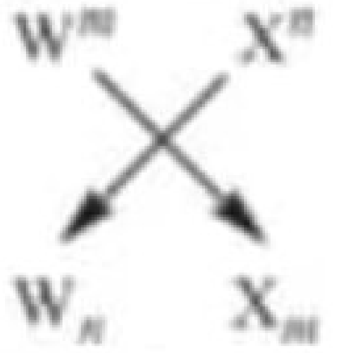
5. Brackets are not used for single elements.
6. The valency 1 is simply assumed and not written in the formula.

The symbols and valencies of the atoms and radicals are important in writing a chemical formula. For example, for arbitrary elements W and X with valencies *m* and *n*, respectively, and where X can be a radical or an atom, the following steps can be used to come up with a chemical formula of their compounds:

Step 1: Write the symbols of the elements and radicals, in this case W and X.

Step 2: Write down the ions used, with their valencies as superscripts, that is W^m X^n.

Step 3: Interchange the valencies of W and X and write them as subscripts.



With practice, the crossing lines are left out since the valencies are exchanged. The formula of the chemical compound is W_nX_m.

Note: When *m* and *n* are equal, there is no need for the exchange, and therefore, are not written, since they are in a ratio of 1:1.

Example 7.4

1. Give the formula of the compound of calcium and chlorine atoms.

Step 1: Write down the symbols for the elements or radicals.
Ca Cl

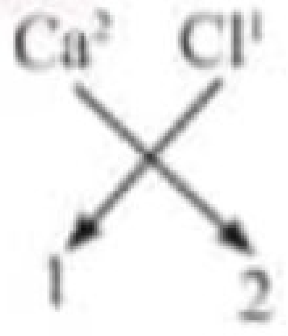
Step 2: Write down the ions used in the compound with their valencies as superscripts.

Ca^{2+} Cl^-

Step 3: Write the valencies as superscripts.

Ca^2 Cl^1

Step 4: Interchange the valencies and write them as subscripts.



$CaCl_2$

The formula of the chemical compound is $CaCl_2$. The compound contains two chlorine atoms for every one calcium atom.

Example 7.5

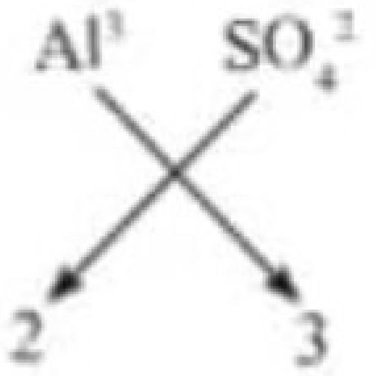
Give the formula of the compound of aluminium and sulphate.

Step 1: Al SO_4

Step 2: Al^{3+} SO_4^{2-}

Step 3: Al^3 SO_4^2

Step 4:



$Al_2(SO_4)_3$

The chemical formula is $Al_2(SO_4)_3$.

Exercise 7.3

1. Write the steps used to arrive to the following chemical formulae of the compounds:

   (a) $MgCl_2$ (b) Na_2SO_4 (c) NH_4NO_3

2. Write down the chemical formula of each of the compounds formed by the combination between the following elements:

   (a) Potassium and chlorine

   (b) Calcium and sulphur

   (c) Lithium and fluorine

Types of chemical formulae

Chemical formulae can basically be divided into three types, namely *empirical formula*, *molecular formula*, and *structural formula*.

Empirical formula is the formula which represents the simplest ratio of the atoms or ions in a compound. The simplest formula is usually determined by considering experimental data. That is why it is called 'empirical' which means 'based on experimentation'. For example, CH_2 shows there are twice as many hydrogen atoms as carbon atoms. It does not show the exact number of each atom of the element in the compound.

A **molecular formula** shows the actual number of each atom in a molecule. It is a multiple of the empirical formula. For example, if the empirical formula is CH_2, its molecular formula may be C_2H_4, C_4H_8, C_6H_{12}, and so on. Therefore, a molecular formula is equal to n × empirical formula, where n is a whole number. Note that when n is 1, the empirical formula equals the molecular formula.

A **structural formula** is a graphic representation of molecular structure showing how the atoms are arranged. At this level, only the empirical and molecular formulae will be studied.

Formula calculations

When the percentage compositions of the elements that make up a compound are known, it is possible to obtain both the empirical and molecular formulae of such a compound. The following are the steps considered when calculating the empirical formula:

Step 1: Obtain the mass of each element in the sample compound. If expressed in percentages, convert the percentage of each of the elements to mass. If the mass or relative molecular mass (R.M.M.) of the compound is not given, an arbitrary mass of 100 g is usually used.

Note: The R.M.M. is the sum of the relative atomic masses (R.A.M.) of all the atoms in a molecule of the compound. The R.A.M. can also be obtained from the Periodic Table (See Appendix 1).

Step 2: Divide the mass of each element by its R.A.M.

Step 3: Divide each of the values obtained in step 2 by the lowest value among them.

Step 4: Convert the ratios in step 3 to whole numbers. These whole numbers give the ratio of each element in the compound. This is the empirical formula.

Example 7.6

What is the empirical formula for a compound of mass 8.1 g if it consists of 4.9 g of magnesium and 3.2 g of oxygen?

Solution

Step 1: Obtain the mass of each element in the compound. These are already given:

Mass of magnesium = 4.9 g

Mass of oxygen = 3.2 g

Step 2: Divide the mass of each element by its R.A.M.

Magnesium, $\frac{4.9}{24} = 0.20$

Oxygen, $\frac{3.2}{16} = 0.20$

Step 3: Divide by the lowest quotient.

$$\begin{array}{ccc} Mg & : & O \\ \frac{0.20}{0.20} & : & \frac{0.20}{0.20} \end{array}$$

Step 4: Obtain their whole number ratios directly or by approximation.

Mg : O

1 : 1

The empirical formula is MgO.

Example 7.7

Given that a certain compound is 69.59% barium, 6.09% carbon and the rest is oxygen. Calculate the empirical formula of this compound.

Solution

Step 1: Assuming that you have 100 g of the compound, then the mass of each element will be:

Ba = 69.59 g

$C = 6.09$ g

The mass of oxygen will be 100 g – (mass of barium + mass of carbon)

$= 100\text{ g} - (69.59\text{ g} + 6.09\text{ g}) = 24.32\text{ g}$

Note that the 100 g is arbitrary for simplifying the calculations.

Step 2: Divide the mass of each element in the sample by its R.A.M.

$$Ba = \frac{69.59}{137.3} = 0.51$$

$$C = \frac{6.09}{12} = 0.51$$

$$O = \frac{24.32}{16} = 1.52$$

Step 3: Divide each value by the smallest number,

$$\frac{0.51}{0.51} = 1$$

$$\frac{0.51}{0.51} = 1$$

$$\frac{1.52}{0.51} = 2.98$$

Step 4: Obtain the whole number ratios.

Ba : C : O

1 : 1 : 3

The empirical formula is therefore, $BaCO_3$.

The empirical formula together with the relative molecular mass can then be used to establish the molecular formula of the compound.

Molecular formula = n × empirical formula, where n is a whole number.

Hint: R.M.M. = n × sum of R.A.M.

Example 7.8

A compound contains 15.8% carbon and 84.2% sulphur. Calculate its empirical formula. If its relative molecular mass is 76, what is its molecular formula?

Solution

Step 1: Assume that you have 100 g of the compound, then find the mass of each element in grams.

C = 15.8 g

S = 84.2 g

Step 2: Divide the mass of each element in the sample by its R.A.M., i.e.

$$\frac{\text{Mass in sample}}{\text{R.A.M.}}$$

$$C = \frac{15.8\text{ g}}{12\text{ g}} = 1.32$$

$$S = \frac{84.2\text{ g}}{32\text{ g}} = 2.63$$

Step 3: Divide throughout by the smallest value, in this case 1.32:

$$\frac{1.32}{1.32} : \frac{2.63}{1.32} = 1:1.99$$

Step 4: Obtain whole number ratios:

C : S

1 : 2

The empirical formula is therefore, CS_2.

The molecular formula = n × empirical formula = $n(CS_2)$

Now, R.M.M. = n × sum of R.A.M.

$$76 = n \times [12 + (2 \times 32)]$$

$$76 = n \times (12 + 64)$$

$$76 = 76\,n$$

$$n = 1$$

Therefore, the molecular formula is CS_2.

Exercise 7.4

1. A compound has 1.121 g of nitrogen, 0.161 g of hydrogen, 0.480 g of carbon and 0.640 g of oxygen. What is its empirical formula?

2. A compound has an empirical formula CH. If it has a relative molecular mass of 78.11, what is its molecular formula?

3. A compound is analysed and found to contain 52.17% carbon, 13.04% hydrogen, and 34.78% oxygen. Calculate:

   (a) its empirical formula.

   (b) its molecular formula, if its relative molecular mass is 46.

Nomenclature of binary inorganic compounds

Everything in the universe bears a name to differentiate it from others. Chemical substances also bear names that range from those of elements to those of compounds. The name of a substance can occur due some factors such as the place of origin, founder, use, and type or classification. Items or substances that fall under a particular group or classification are named systematically. A systematic way of assigning names to items that belong to a particular group or classification is called *nomenclature*.

Binary inorganic compounds

While an *inorganic compound* is any substance in which two or more chemical elements (usually other than carbon) are combined, always in definite proportions, a *binary compound* is the one which is formed by two chemical substances. For example, CaO, NaCl, and PCl, are *binary inorganic compounds*.

Inorganic compounds are categorised into *ionic* and *covalent*. The nomenclature of ionic compounds differs slightly from that of covalent compounds.

Nimenclature of binary ionic compounds

Ionic compounds are formed when a metal combines with a non-metal. The following are the steps considered when naming binary ionic compounds:

1. Name the metallic ion that appears first in the formula using the name of the element itself.

2. The second part of the formula which is usually an anion in the compound will end with a suffix "ide". For example, oxygen becomes oxide, hydrogen becomes hydride and chlorine becomes chloride.

Note:

Some metals always have fixed charges when they form ions, that is,

(a) Group I metals have a charge of +1.

(b) Group II metals have a charge of +2.

(c) Group III metals have a charge of +3.

(d) Silver (Ag) has a charge of +1.

(e) Zinc (Zn) has a charge of +2.

Other metals are multivalent and can thus form more than one ion. For example; iron (Fe) is bivalent; it has valencies of 2 and 3, copper (Cu) is also bivalent; it has valencies of 1 and 2. Compounds formed from these metals must be distinguished by stating which valency has been used in the compound. The valency of the respective metal is indicated by capital Roman numbers in parentheses (brackets)

Example 7.9

What is the name of the compound with the formula $FeCl_3$?

Solution

The total charge of the molecule is zero and Cl^- has a negative charge.

(i) Let x be the charge of Fe

(ii) $1(x) + 3(-1) = 0$

(iii) $x = +3$

(iv) So, the Fe is in the +3 oxidation state. Write the name 'iron' and place III in brackets beside it.

(v) Use the name chlorine but change the last three letters to "ide": So the name is iron(III) chloride.

Example 7.10

What is the name of the compound with the formula CuS?

Solution

(i) Let x be the charge of Cu.

(ii) Sulphur has a charge of –2.

(iii) $1(x) + 1(-2) = 0$
$x = +2$ for Cu

(iv) Write the name copper and place II in brackets beside it.

(v) Use the name sulphur but change the last two letters to "ide". The name of the compound is copper(II) sulphide.

Other examples of binary ionic compounds are given below.

(a) MgO is named magnesium oxide.

(b) $AlCl_3$ is named aluminium chloride .

(c) MnO_2 is named manganese(IV) oxide.

Note: Manganese can have more than one charge, but each oxygen ion has a charge of –2. In order for the compound to be neutral, Mn must have a charge of +4.

Nomenclature of binary covalent compounds

Covalent compounds are formed between two non-metal elements. These compounds are named differently from ionic compounds. The number of atoms are presented by prefixes as shown in Table 7.3.

Table 7.3: *Examples of prefixes*

Number	Prefix	Number	Prefix
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

The following are the steps considered when naming binary covalent compounds:

1. Give the name of the first element.
2. Give the name of the second element with the ending changed to –ide.
3. If more than one compound is possible between the two elements, give prefixes to indicate the number of atoms of each element.

Example 7.11

Give the name for PCl_3.

Answer

(i) Since there is one phosphorus atom, use it as the first part of the name.

(ii) There are three chlorine atoms, so use 'tri' in front of chlorine, then drop the 'ine' in chlorine and replace with 'ide'.

The name is phosphorus trichloride.

Example 7.12

What is the name for N_2O_4?

Answer

(i) Use the prefix 'di' in front of nitrogen since there are two atoms.

(ii) Use the prefix 'tetra' in front of the oxygen since there are four atoms.

(iii) Drop '–ygen' and replace with 'ide:

(iv) The name is dinitrogen tetraoxide.

Table 7.4 gives the formulae and names of some binary covalent compounds.

Table 7.4: *Some binary covalent compounds*

Formula	Name
CO_2	Carbon dioxide
CO	Carbon monoxide
N_2O_5	Dinitrogen pentoxide
HCl	Hydrogen chloride
NO	Nitrogen monoxide
SF_6	Sulphur hexafluoride

Note: The names of compounds involving radicals according to the International Union of Pure and Applied Chemistry (IUPAC) are provided in Appendix 2.

Chemical names of common substances

Chemical names are usually used to give accurate descriptions of the compositions of substances including those we encounter daily. For instance, it is very rare to ask someone to give you some sodium chloride (common salt) to use in your food, instead, you will ask for common salt. That is why there are 'common names' for some substances. However, it is important to note that some *common names* are inaccurate and may vary from one place to another, and therefore, they cannot tell the chemical composition of a substance. Table 7.5 lists the common names for some chemicals and their respective chemical formulae.

Table 7.5 *Common names and formulae of some substances*

Common name	Chemical name	Chemical formula
Baking soda	Sodium hydrogen carbonate or sodium bicarbonate	$NaHCO_3$
Gypsum	Hydrated calcium sulphate	$CaSO_4 \cdot 2H_2O$
Marble	Calcium carbonate	$CaCO_3$
Caustic soda	Sodium hydroxide	$NaOH$
Caustic potash	Potassium hydroxide	KOH
Chalk	Calcium carbonate	$CaCO_3$
Common salt	Sodium chloride	$NaCl$
Soda ash	Sodium carbonate	Na_2CO_3
Lime water	Calcium hydroxide solution	$Ca(OH)_2$
Slaked lime	Calcium hydroxide solid	$Ca(OH)_2$
Quick lime	Calcium oxide	CaO
Plaster of Paris (POP)	Calcium sulphate	$CaSO_4$
Water	Dihydrogen monoxide	H_2O
Vitamin C	Ascorbic acid	$C_6H_8O_6$
Sugar	Sucrose	$C_{12}H_{22}O_{11}$
Sulphate of ammonia (SA)	Ammonium sulphate	$(NH_4)_2SO_4$
Urea	Carbamide	$CO(NH_2)_2$
Fluospar/fluorite	Calcium fluoride	CaF_2
Asbestos	Magnesium silicate	$Mg_3Si_2O_5(OH)_4$
Aspirin	Acetylsalicylic acid	$C_9H_8O_4$
Chloroform	Trichoromethane	$CHCl_3$
Dolomite	Calcium magnesium carbonate	$CaMg(CO_3)_2$
Limestone	Calcium carbonate	$CaCO_3$
Milk of magnesia	Magnesium hydroxide	$Mg(OH)_2$
Bleach	Sodium hypochlorite	$NaOCl$
Saltpetre	Potassium nitrate	KNO_3

Chapter summary

1. Chemical bonding involves electrons in the outermost shell of an atom. When the outermost shell is fully-filled, the atom is said to be stable.
2. Ions are formed when an atom gains or loses electron(s). Cations are positively charged ions that result from atoms losing one or more electrons. Anions are negatively charged ions that result from atoms gaining one or more electrons.
3. Ionic (electrovalent) bonding usually occurs between a metal and a non-metal. It involves the transfer of electron(s) from the atoms of the metal to the atoms of the non–metal.
4. Covalent bonding takes place between two or more non-metals. It involves atoms of the non-metals sharing electrons that are in their outermost shells.
5. The ability of an atom to combine with other atoms according to the number of electrons it can give, take or share is known as valency.
6. The oxidation state (oxidation number) of an element is the number of electrons that need to be added, shared or removed by its atom, to make a neutral molecule. The oxidation number is arbitrary and may be positive, negative or zero.
7. A radical is a group of atoms which behaves as a single unit and has a positive or negative charge. A radical can also be an atom, molecule or ion that has unpaired valence electron. Such a group maintains its identity throughout any chemical reaction.
8. A chemical formula is a representation that uses chemical symbols to show the proportions of the elements present in a chemical compound.
9. An empirical formula is the simplest way of writing a chemical formula and indicates the ratio of the atoms in a compound.
10. A molecular formula is a chemical formula that shows the total number of atoms of each element in a molecule.
11. A systematic way of naming items or substances of a particular category is known as nomenclature.

Revision exercise 7

Choose the correct answer to each of the questions 1 to 11.

1. An element with chemical symbol $^{39}_{19}K$ has a valency of

   (a) 1 (b) 2 (c) 3 (d) 4

2. What name is given to the force of attraction that holds atoms together to form a molecule?

   (a) Chemical change
   (b) Chemical bond
   (c) Friction
   (d) Centripetal

3. During the formation of ions,

   (a) non-metal atoms gain protons.
   (b) metal atoms lose their outermost electrons.
   (c) metal atoms gain electrons in their outermost shells.
   (d) non-metal atoms lose electrons of their outermost shells.

4. Usually, electrovalent bonding occurs between

   (a) metals and metals.
   (b) metals and non-metals.
   (c) metals and inert gases.
   (d) non-metals and non-metals.

5. The following atoms cannot exist freely as single atoms, **except**

   (a) sodium atom.
   (b) chlorine atom.
   (c) argon atom.
   (d) magnesium atom.

6. A covalent bond is formed due to

   (a) opposite charges of atoms.
   (b) transfer of electrons of atoms.
   (c) forces of attraction between atoms.
   (d) sharing of electrons between atoms.

7. Which of the following is a property of covalent compounds?
   - (a) They do not conduct electricity.
   - (b) They conduct heat easily.
   - (c) They are mostly soluble in water.
   - (d) Their melting and boiling points are very high.

8. What is the oxidation number of nitrogen in the radical NH_4^+?

   (a) +2 (b) –3 (c) +4 (d) +5

9. The valency of the phosphate radical is

   (a) 1 (b) –3 (c) 3 (d) 4

10. Which of the following sets of symbols represent cations?
    - (a) K^+ and Mg
    - (b) Mg and Al
    - (c) Al^{3+} and Cl^-
    - (d) K^+ and Al^{3+}

11. Identify the chemical formula for potassium chlorate(V).

    (a) K_2ClO_3 (b) $K(ClO_3)_2$ (c) K_3ClO_3 (d) $KClO_3$

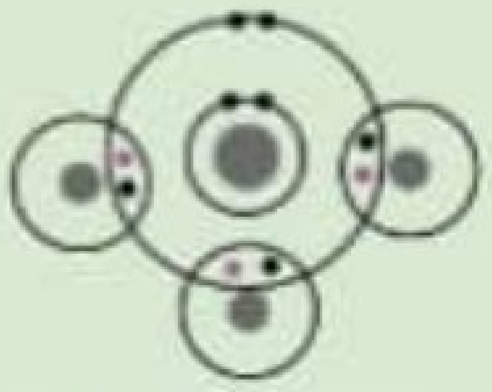
12. (a) Identify the compounds with the correct or incorrect IUPAC names.
    - (i) KCl – Potassium chloride
    - (ii) $Fe_2(SO_4)_3$ – Iron(II) sulphate(VI)
    - (iii) $CaCl_2$ – Calcium dichloride
    - (iv) Na_2SO_4 – Sodium tetraoxosulphate(VI)

    (b) Rename the incorrectly named compound(s) in (a) using the IUPAC rules.

13. What do you understand by the following terms?
    - (a) Valency
    - (b) Oxidation state
    - (c) Radical
    - (d) Empirical formula
    - (e) Ionic bond
    - (f) Binary compounds

14. You are provided with the following list of substances: Mg, Ca^{2+}, Cl, N_2, Al^{3+}, and H_2. Which of these substances are:
    (a) atoms?
    (b) molecules?
    (c) ions?

15. A molecule of a certain gas can be represented by the following diagram:



    (a) What is the name of the molecule?
    (b) What is the molecular formula of the gas?
    (c) What type of bonding holds the atoms of the molecule?
    (d) Name other compounds with this type of bonding.

16. Write the electronic configuration of each of the following species:
    (a) Sodium
    (b) Aluminum ion
    (c) Magnesium
    (d) Chlorine ion
    (e) Sulphur
    (f) Neon

17. Choose the correct formula for the combination of the following ions:
    (i) Mg^{2+} and PO_4^{3-}
        (a) $Mg_3(PO_4)_2$ (b) Mg_3PO_4 (c) $Mg_2(PO)_3$ (d) $MgPO_4$
    (ii) Ba^{2+} and N^3
        (a) BaN (b) BaN_2 (c) Ba_2N_3 (d) Ba_3N_2
    (iii) Al^{3+} and SO_4^{2-}
        (a) $AlSO_4$ (b) $Al(SO_4)_2$ (c) $Al_2(SO_4)_3$ (d) $Al_3(SO_4)_2$
    (iv) Zn^{2+} and $C_2H_3O_2^-$
        (a) $ZnC_2H_3O_2$ (b) $Zn_2C_2H_3O_2$ (c) $Zn(C_2H_3O_2)_2$ (d) $Zn_2(C_2H_3O_2)_2$

18. Name the following compounds:

(a) MgI_2 (b) N_2O_5 (c) CCl_4 (d) $FeBr_2$ (e) CuI_2

(f) H_2S (g) Hg_2Cl_2 (h) K_2O (i) PCl_5 (j) SF_6

19. The colour of CrO_4^{2-}*(aq)* is yellow and that of $Cr_2O_7^{2-}$*(aq)* is orange. Calculate the oxidation number of Cr in both of these radicals.

20. A compound of sulphur and oxygen is 40.1% sulphur by mass. What is the empirical formula for the compound? The R.A.M. of S and O are 32.07 and 16.00, respectively.

21. Write the chemical formulae of the following compounds:
    (a) Sodium hydrogen carbonate
    (b) Silver trioxonitrate(V)
    (c) Copper(I) oxide
    (d) Aluminium tetraoxosulphate(VI)

22. Write down any two different ionic states in each of the following elements:

(a) Fe (b) Cu (c) Pb (d) Mn

23. An atom of element X (atomic number 11) and an atom of element Y (atomic number 9) combine to form a compound.
    (a) Write the formula of the compound.
    (b) State the type of the bond present in the compound.

24. A hydrocarbon contains 88.88% carbon and 11.12% hydrogen by mass. Calculate the empirical formula of the molecule.

25. A compound consists of calcium 40%, carbon 12% and oxygen 48% by mass. Calculate the empirical formula of the compound.

26. (a) Calculate the oxidation number of the underlined elements in the following compounds:

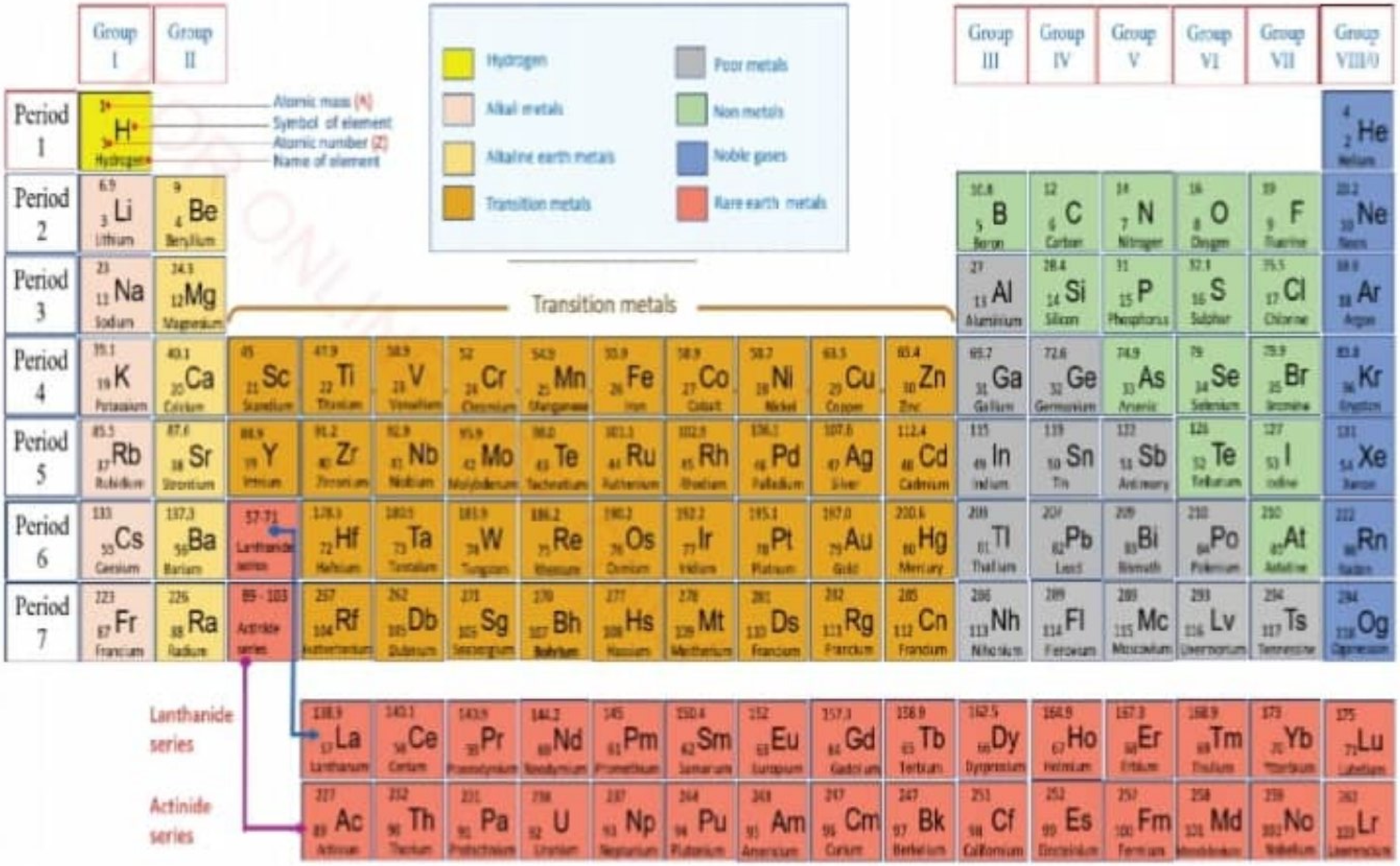
(i) $\underline{N}H_4Cl$ (ii) $Na_2\underline{S}O_4$ (iii) $\underline{Al}_2O_3$ (iv) $H_2\underline{O}_2$

27. Elements *T* and *Q* have atomic numbers 12 and 17, respectively. Use the two elements to answer the following questions:
    (a) Write the electronic configuration of element *Q*.
    (b) What is the valency of element *T*?
    (c) Write the chemical formula of a compound formed when *T* and *Q* combine.
    (d) Mention the type of the bond formed by the combination of elements *T* and *Q*.
    (e) In which group and period in the Periodic Table does element *Q* belong?

28. An organic compound contains 26.70% carbon, 2.20% hydrogen and 71.10% oxygen. If its vapour density is 45. Determine its molecular formula.

Appendices

Appendix 1: *The Periodic Table*



	Group I	Group II											Group III	Group IV	Group V	Group VI	Group VII	Group VIII/0
Period 1	1 1 H Hydrogen																	4 2 He Helium
Period 2	6.9 3 Li Lithium	9 4 Be Beryllium											10.8 5 B Boron	12 6 C Carbon	14 7 N Nitrogen	16 8 O Oxygen	19 9 F Fluorine	20.2 10 Ne Neon
Period 3	23 11 Na Sodium	24.3 12 Mg Magnesium											27 13 Al Aluminium	28.4 14 Si Silicon	31 15 P Phosphorus	32.1 16 S Sulphur	35.5 17 Cl Chlorine	39.9 18 Ar Argon
Period 4	39.1 19 K Potassium	40.1 20 Ca Calcium	45 21 Sc Scandium	47.9 22 Ti Titanium	50.9 23 V Vanadium	52 24 Cr Chromium	54.9 25 Mn Manganese	55.9 26 Fe Iron	58.9 27 Co Cobalt	58.7 28 Ni Nickel	63.5 29 Cu Copper	65.4 30 Zn Zinc	69.7 31 Ga Gallium	72.6 32 Ge Germanium	74.9 33 As Arsenic	79 34 Se Selenium	79.9 35 Br Bromine	83.8 36 Kr Krypton
Period 5	85.5 37 Rb Rubidium	87.6 38 Sr Strontium	88.9 39 Y Yttrium	91.2 40 Zr Zirconium	92.9 41 Nb Niobium	95.9 42 Mo Molybdenum	98.0 43 Te Technetium	101.1 44 Ru Ruthenium	102.9 45 Rh Rhodium	106.4 46 Pd Palladium	107.8 47 Ag Silver	112.4 48 Cd Cadmium	115 49 In Indium	119 50 Sn Tin	122 51 Sb Antimony	128 52 Te Tellurium	127 53 I Iodine	131 54 Xe Xenon
Period 6	133 55 Cs Caesium	137.3 56 Ba Barium	57-71 Lanthanide series	178.5 72 Hf Hafnium	180.9 73 Ta Tantalum	183.9 74 W Tungsten	186.2 75 Re Rhenium	190.2 76 Os Osmium	192.2 77 Ir Iridium	195.1 78 Pt Platinum	197.0 79 Au Gold	200.6 80 Hg Mercury	208 81 Tl Thallium	207 82 Pb Lead	209 83 Bi Bismuth	210 84 Po Polonium	210 85 At Astatine	222 86 Rn Radon
Period 7	223 87 Fr Francium	226 88 Ra Radium	89 - 103 Actinide series	257 104 Rf Rutherfordium	262 105 Db Dubnium	271 106 Sg Seaborgium	270 107 Bh Bohrium	277 108 Hs Hassium	278 109 Mt Meitnerium	281 110 Ds Francium	282 111 Rg Francium	285 112 Cn Francium	286 113 Nh Nihonium	289 114 Fl Flerovium	289 115 Mc Moscovium	293 116 Lv Livermorium	294 117 Ts Tennessine	294 118 Og Oganesson

Transition metals

Lanthanide series	138.9 57 La Lanthanum	140.1 58 Ce Cerium	140.9 59 Pr Praseodymium	144.2 60 Nd Neodymium	145 61 Pm Promethium	150.4 62 Sm Samarium	152 63 Eu Europium	157.3 64 Gd Gadolinium	158.9 65 Tb Terbium	162.5 66 Dy Dysprosium	164.9 67 Ho Holmium	167.3 68 Er Erbium	168.9 69 Tm Thulium	173 70 Yb Ytterbium	175 71 Lu Lutetium
Actinide series	227 89 Ac Actinium	232 90 Th Thorium	231 91 Pa Protactinium	238 92 U Uranium	237 93 Np Neptunium	244 94 Pu Plutonium	243 95 Am Americium	247 96 Cm Curium	247 97 Bk Berkelium	251 98 Cf Californium	252 99 Es Einsteinium	257 100 Fm Fermium	258 101 Md Mendelevium	259 102 No Nobelium	262 103 Lr Lawrencium

- Hydrogen
- Alkali metals
- Alkaline earth metals
- Transition metals
- Poor metals
- Non metals
- Noble gases
- Rare earth metals

Appendix 2: *IUPAC names for common compounds*

Formula	Common name	IUPAC name
H_2SO_4	Sulphuric acid	Sulfuric acid or Tetraoxosulphate(VI) acid
HNO_3	Nitric acid	Nitric acid or Trioxonitrate(V) acid
HCl	Hydrochloric acid	Hydrochloric acid
H_2CO_3	Carbonic acid	Trioxocarbonate(IV) acid
HNO_2	Nitrous acid	Dioxonitrate(III) acid
H_2SO_3	Sulphurous acid	Trioxosulphate(IV) acid
Na_2CO_3	Sodium carbonate	Sodium trioxocarbonate(IV)
$NaNO_3$	Sodium nitrate	Sodium trioxonitrate(V)
CO_2	Carbon dioxide	Carbon(IV) oxide
CO	Carbon monoxide	Carbon(II) oxide
SO_2	Sulphur dioxide	Sulphur(IV) oxide
SO_3	Sulphur trioxide	Sulphur(VI) oxide
N_2O	Dinitrogen oxide or nitrous oxide	Nitrogen(I) oxide
NO	Nitrogen monoxide or nitric oxide	Nitrogen(II) oxide
N_2O_4	Dinitrogen tetraoxide	Dinitrogen tetraoxide
Na_2SO_4	Sodium sulphate	Sodium tetraoxosulphate(VI)
$CuSO_4$	Cupric sulphate/Copper sulphate	Copper(II) tetraoxosulphate(VI)
$Al(NO_3)_3$	Aluminium nitrate	Aluminium trioxonitrate(V)
$CaCO_3$	Calcium carbonate	Calcium trioxocarbonate(IV)
CO_3^{2-}	Carbonate ion	Trioxocarbonate(IV) ion
$FeCO_3$	Ferrous carbonate	Iron(II) trioxocarbonate(IV)
PbO	Lead monoxide	Lead(II) oxide
FeO	Ferrous oxide	Iron(II) oxide
Fe_2O_3	Ferric oxide	Iron(III) oxide
SO_4^{2-}	Sulphate ion	Tetraoxosulphate(VI) ion
SO_3^{2-}	Sulphite ion	Trioxosulphate(IV) ion
HCO_3^-	Hydrogencarbonate ion	Hydrogen trioxocarbonate(IV) ion
MnO_2	Manganese dioxide	Manganese(IV) oxide
MnO_4^-	Permanganate ion	Tetraoxomanganate(VI) ion
$Cr_2O_7^{2-}$	Dichromate ion	Heptaoxodichromate(VI) ion

Glossary

Acidic oxide	an oxide that reacts with water to form an acid, or with a base to form salt and water. Acidic oxides are oxides of either non-metals or of metals in high oxidation states
Anode	a positively charged electrode to which negative ions move towards it
Anhydrous	it contains no water
Basic oxide	an oxide that shows basic properties and that can either react with water to form a base; or reacts with an acid to form salt and water
Biogas	a mixture of methane and other minor gases such as carbon dioxide
Catalyst	a substance that alters the rate of a chemical reaction but it remains unchanged at the end of the reaction
Cathode	a negatively charged electrode to which positive ions move towards it
Decomposition	breaking down of a chemical compound into elements or simpler compounds
Distillation	the action of purifying a liquid by a process of heating and cooling
Electrode	a positive or negative pole to which negative or positive ions move, respectively, during electrolysis
Electrolysis	the breaking down of compound in solution form by means of electricity
Fractional distillation	the separation of liquid components from a liquid mixture with closer boiling points of the components
Globe	World, Earth, Universe or Planet
Haber process	an industrial method of manufacturing ammonia from nitrogen and hydrogen, using a metal catalyst at high temperature and pressure

Liquefaction	a process that generates a liquid from a solid or a gas
Mortar	a mixture of cement, sand and water
Oxidation	addition of oxygen to a substance or removal of hydrogen from a substance
Oxidizing agent	a substance that adds oxygen to another substance or removes hydrogen from another substance
Orbit	passing through/revolve or a path through which to revolve
Ores	a natural rock or sediment that contains one or more valuable minerals, typically metals, that can be mined and treated for different purposes
Oxy-hydrogen flame	a very hot flame produced by the combustion of a mixture of oxygen and hydrogen
Reduction	the addition of hydrogen to a substance or removal of oxygen from a substance
Reducing agent	a substance which adds hydrogen to another substance or removes oxygen from another substance
Respiration	a process in living organisms involving the production of energy, typically with the intake of oxygen and the release of carbon dioxide from the oxidation of complex organic substances
Saline water	the water containing dissolved salts (salty water)
Sod	a rectangular piece that has been cut from an area of grass or a surface on the ground with the grass growing on it
Steam reforming	a method of producing hydrogen from organic compounds such as methane, whereby the compound decomposes into carbon monoxide and hydrogen
Transpiration	a process by which plants lose water through the stomata of their leaves to the atmosphere
Universal solvent	a solvent capable of dissolving all or many chemical substances

Bibliography

Chang. R. (2010). *Chemistry* (10th ed.). New York, NY: McGraw-Hill Companies Inc.

Jespersen, N. D & Brady, J. E. (2012). *Chemistry, The Molecular Nature of Matter* (6th ed.). River street, Hoboken, John Wiley & Sons, Inc.

Kenkel, J. (2011). *Basic Chemistry Concepts and Exercises*. Broken Sound Parkway, NW: Taylor and Francis Group LLC.

SCSU & MoEVT-Zanzibar. (2008). *Chemistry for Zanzibar Secondary Schools, Forms 1 &2*. Dar es Salaam: Oxford University Press Tanzania Ltd.

Tanzania Institute of Education. (2019). *General Chemistry Textbook for Advanced Secondary Education*, TIE, Dar es Salaam, Tanzania.

Tanzania Institute of Education. (2019). *Inorganic Chemistry Textbook for Advanced Secondary Education*, TIE, Dar es Salaam, Tanzania.

Tanzania Institute of Education. (2019). *Organic Chemistry Textbook for Advanced Secondary Education*, TIE, Dar es Salaam, Tanzania.

Tanzania Institute of Education. (2019). *Physical Chemistry Textbook for Advanced Secondary Education*, TIE, Dar es Salaam, Tanzania.

Tanzania Institute of Education. (1995). *Secondary School Chemistry Book One* (1st ed.). Dar es Salaam: Dar es Salaam University Press.

Tanzania Institute of Education. (1995). *Secondary School Chemistry Book Two* (1st ed.). Dar es Salaam: Dar es Salaam University Press.

Zumdahl, S. S. & Zumdahl, S. A. (2014). *Chemistry* (9th ed.). Belmont: Mary Firich.

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ISBN 978-9987-09-287-1

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