

CHAPTER ONE

INTRODUCTION TO CHEMISTRY

Learning objectives

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

1. (a) *Define Chemistry.*
 (b) *State: (i) The branches of chemistry*
 (ii) *The aims and objectives of learning chemistry.*
 2. *Define the following terms:*
 - *Hypothesis,*
 - *Scientific Theory and*
 - *Scientific Laws (Principles).*
 3. (a) *Name and state the common Laboratory Apparatus.*
 (b) *Draw and label the diagram of the Bunsen burner.*
 (c) (i) *Name the types of flames.*
 (ii) *Differentiate between the types of flames.*
 4. *State the Laboratory rules.*
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1.1 INTRODUCTION

Chemistry is one of the Science subjects. At secondary level science is divided into two main branches, namely:

- Biological Science and
- Physical Science.

Biological Science - *is a branch of science which deals with the study of living things.*

Physical Science - *is a branch of science which deals with the study of non-living things.*

Physical Science is further divided into Physics and Chemistry

Definition

Chemistry is a branch of physical science which deals with the study of the composition, structure, properties, and interactions of matter.

Branches of Chemistry

Chemistry is mainly divided into three branches. These are:

Physical chemistry *is fundamental to all chemistry and deals with the application of physical laws to chemical systems and chemical change.*

Organic chemistry *branch of chemistry in which carbon compounds and their reactions are studied.*

Inorganic chemistry *is the study of the structure, properties, and reactions of the chemical elements and their compounds.*

Other branches are:

Analytical chemistry - *is the science of separating complex materials into simpler ones and detecting and measuring the constituents.*

Photochemistry - *the study of chemical changes that are initiated by light.*

Biochemistry - *is the chemistry of living organisms and life processes.*

Geochemistry - *is the application of chemistry to processes taking place in the earth, such as mineral formation, the metamorphosis of rocks, and the formation and migration of petroleum.*

Note: People who study Chemistry are called *Chemists*.

1.2 Aims of Teaching Chemistry in Schools

- (i) To make a society that knows about chemistry and enjoys the fruits of chemistry. E.g. Chemistry teachers, Doctors, Nurses, Physicians (specialists in medicine) etc.
- (ii) To enable a student to think critically in a learning situation.
- (iii) To help a student to appreciate the importance and the application of chemistry in our daily life. E.g. The application of chemicals such as pesticide, insecticides (like DDT) to fight diseases.
- (iv) To help a student to develop an experimental attitude of the mind by performing experiments, making correct observations and draw conclusions from the observations made.
- (v) To help a student to communicate chemical information clearly.
- (v) To help a student to pass examinations so as to prepare him/her for further studies in chemistry or training in which the knowledge of chemistry is required.
- (vi) To familiarize the student with the scientific hypothesis, theories and law or principles.

1.3 Hypothesis, Theory and Laws (Principles)

(a) Hypothesis:

Hypothesis is a scientific idea put forward and is still in the stage of experimental investigation. However, if proved to be correct it then becomes a law or principle.

(b) Scientific Law (Principle):

*A scientific law or principle is a generalized statement of observed facts.
E.g. Graham's law of diffusion, Boyle's law, Graham's law etc.*

(c) Scientific Theory:

*A scientific theory is a scientific idea put forward to explain the existence of one or more laws.
E.g. Dalton's atomic Theory.*

NB: *A theory may vary from time to time due to advancement in technology.*

1.40 How to Study Chemistry

The study of Chemistry as a science subject is based on *scientific principles*. These principles can best be understood by using experimental approach. These experiments are carried out in Chemistry laboratory.

1.41 Chemistry Laboratory

Science courses typically include laboratory sections in which students conduct experiments that replicate or illustrate scientific principles. Chemistry laboratory therefore refers to a room or a building for carrying out scientific experiments. In the laboratory, a number of laboratory apparatus are used in carrying out the experiments.

(NB: *In general terms laboratory is a room or building for scientific experiments, research, or teaching, or for the manufacture of drugs or chemicals).*

1.42 Laboratory Apparatus

These are equipments/instruments used for carrying out experiments in the laboratory. Some of the apparatus commonly used include:-

- Beakers,
- Flasks,
- Tubes,
- Funnels,
- Heating apparatus and
- Others.

The Diagrams and uses of the Common Laboratory Apparatus are shown below

(a) Beakers

A beaker is a lipped cylindrical glass or plastic container used for measuring volumes of liquids and solutions. They are of various capacities. The common capacities are: 100 cm^3 , 200 cm^3 , 250 cm^3 , 500 cm^3 and 600 cm^3 .

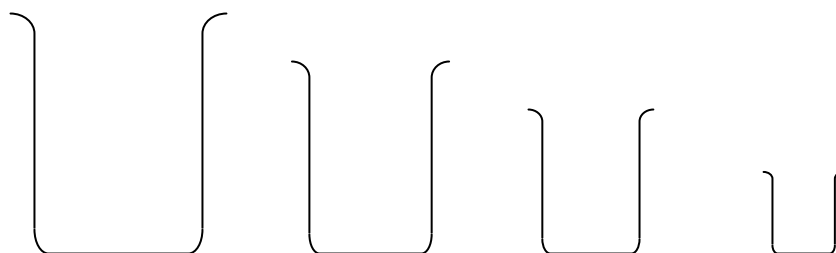


Figure 1 Diagram of beakers of various sizes

(b) Flasks - A flask is a narrow-necked conical or spherical bottle.

They are classified according to their shapes and the purpose they serve. They include:

- (i) *A conical flask*
It is used in titration experiments in volumetric analysis.
- (ii) *A Flat Bottom Flask*
It is used for carrying out experiments in gas preparations.
- (iii) *A Round bottom Flask*
It is used for carrying out experiments in gas preparations which involve heating.
- (iv) *A distillation flask*
It is used for distillation of liquids.
- (v) *A volumetric flask*
It has a flat bottom. It is used for preparation of standard solutions (i.e. solutions of known concentrations) in volumetric analysis.

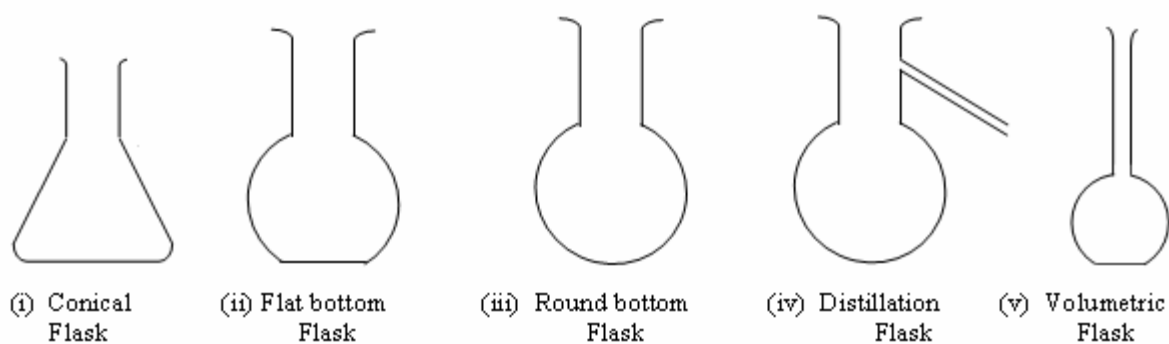


Figure 2 Diagrams of flasks

(c) Funnels

A funnel is an apparatus that is wide at the top and narrow at the bottom, used for guiding liquid or powder into a small opening or used for separating immiscible liquids.

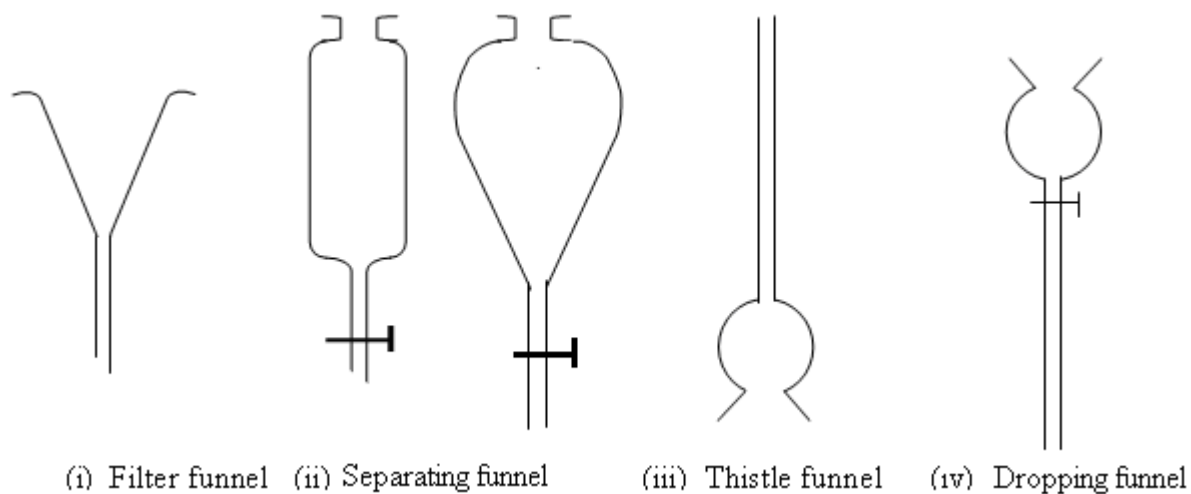
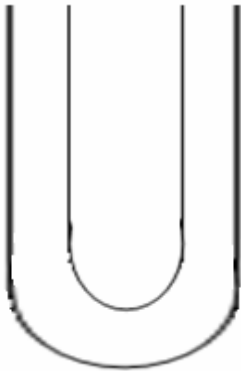

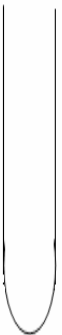


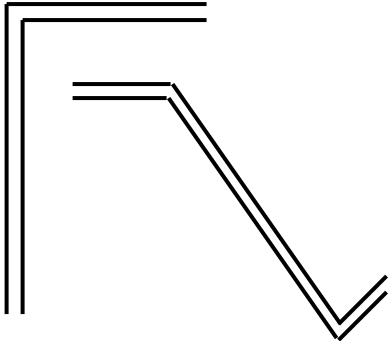


Figure 3 Diagrams of the different types of funnels

(d) Tubes

A tube is a glass container open at one end and sealed at the other end. Or
A tube is a long, hollow cylinder used for conveying fluids (liquids or gases).
The common tubes used in the chemistry laboratory are:

<p>(i) U-Tubes</p>  <p>It is used mainly for parking solid drying agents for dry gases.</p>	<p>(ii) Boiling tube</p>  <p>It is used for heating test solutions to boiling point.</p>
<p>(iii) Test tube</p>  <p>It is smaller in diameter compared to boiling tube. It is used for testing chemicals in the laboratory.</p>	<p>(iv) Ignition tube</p>  <p>It is smaller than a test tube and is used for heating solids strongly.</p>
<p>(v) Combustion tube</p>  <p>It is used mostly in reduction reactions where a current of gas is to be passed through over a heated compound.</p>	<p>(vi) Delivery tubes</p>  <p>They are used for delivering gas into the collection point during gas preparation.</p>

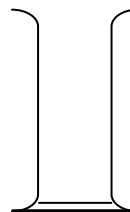
(e) Other apparatus include:

(i) Trough



It is made of glass. It is mostly used to place gas jar stand/bee hive during collection of gases.

(ii) Gas jar



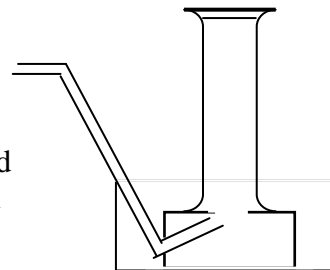
It is used for collecting gases during preparation of gases

(iii) A gas jar stand or a bee hive shelf



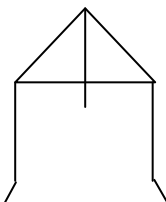
It is used for supporting a gas jar during gas preparations

(iv) A gas jar standing on a bee-hive shelf in a trough



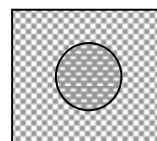
This set up is used when collecting a gas over water.

(v) Tripod Stand



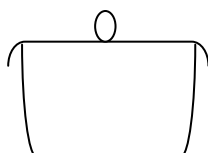
This helps to support apparatus during heating

(vi) Wire Gauze



It is made of iron wires. It is placed on a tripod stand to support apparatus when heating. The central part helps to spread the heat evenly.

(vii) Crucible



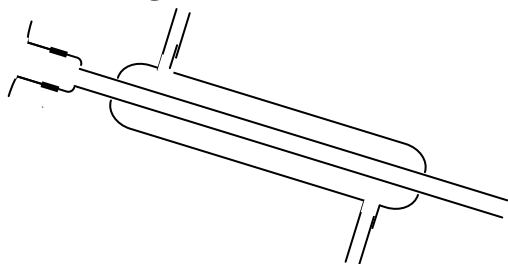
It is used in the heating of solids

(viii) Evaporating Dish



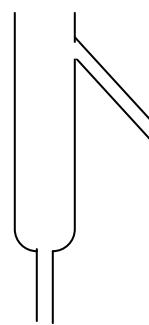
It is made of glass. It is used in evaporating solutions by heating.

(ix) Lie big Condenser:



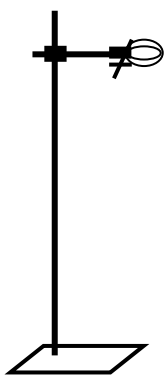
It is used to condense vapour given out during distillation experiment.

(x) Fractionating Column



It is used in *fractional distillation* to separate miscible liquids especially those whose boiling points are very close.

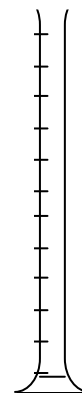
(xi) Retort Stand:



It is made up of either wood or iron.

It is used for holding or supporting apparatus in the right position during an experiment.

(xii) Measuring Cylinder



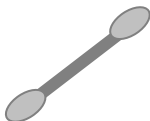
It is either made of glass or plastic.

It is used for measuring volumes of solutions or liquids of specific volumes.

The common capacities are: 10 cm^3 , 100 cm^3 , 250 cm^3 , 500 cm^3 , 1000 cm^3 ,

(xiii) Spatula

It is used for scooping solid substances from their stock containers. It is also used in direct heating of solids if made of metal.



(xv) Stirring rod

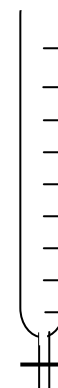
It is used for stirring solutions. It is either made of plastic or glass.



(xiv) Burette

It is used in titration experiments for adding accurately small solution volumes of solutions on solution in a conical flask

Its capacity is 50 cm^3 beginning with 0 cm^3 from top.



(xvi) Test tube holder

It is used for holding test tubes when the substance under test is being heated.

(xvii) Pipette

It is used for measuring a specific volume of solutions. The capacities of pipettes which are common in use are: 10 cm^3 , 20 cm^3 , 25 cm^3



Figure 5

Diagrams of other laboratory apparatus

1.43 The Bunsen burner

A Bunsen burner is a heating device widely used in school laboratories. When used properly, it provides a hot, steady, smokeless flame.

(a) Structure of the Bunsen burner

The burner is a short, vertical tube of metal connected to a gas source and perforated at the bottom to admit air. The flow of air in to the burner is controlled by an adjustable collar or metal ring on the tube.

The Diagram of Bunsen burner

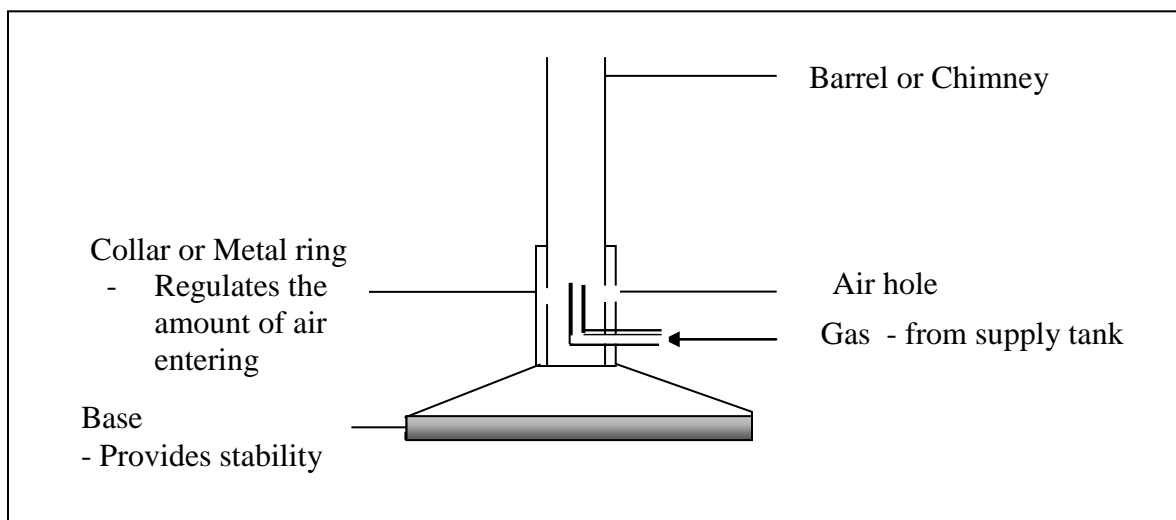


Figure 1

(b) How to light a Bunsen burner

In lighting the Bunsen burner, the following steps are followed:

- (i) Close the air holes by rotating the metal ring.
- (ii) Connect the Bunsen burner to the gas supply tank.
- (iii) Light a match stick keeping it near the top of the barrel.
- (iv) Turn on the gas and open the air holes slowly until the preferred flame is obtained.

(c) Burning back or striking back

Burning back occurs when the air holes are fully opened. Too much air enters through the air holes such that the rate at which the gas burns exceeds the rate at which it is supplied from the jet. As such the flame moves down the barrel and burns at the jet. When a flame burns back, turn off the gas, close the air holes and light it again using the usual procedure.

1.44 **Flames**

A flame refers to a glowing body of mixed gases that produce heat and light when undergoing the process of combustion.

Flames generally consist of a mixture of oxygen (or air) and other gases.

Type of Flames

A Bunsen burner produces two types of flames, namely: -

- Luminous flame &
- Non-luminous flame.

Formation of the above types of flames depends on whether the air hole is open or closed.

(a) Luminous flame

A luminous flame is a type of flame, which produces much light.

It is formed when the air hole of the Bunsen burner is closed.

Characteristics of Luminous Flame

- (i) It has four parts/zones, which have different colours.
- (ii) It burns silently.
- (iii) It is not hot enough.
- (iv) It produces soot (black smoke)
- (v) It is mainly yellow in colour.
- (vi) It is unsteady.

(b) Non-luminous Flame

A non-luminous flame is a type of flame, which produces very little light. It is formed when the air hole of the Bunsen burner is open.

Characteristics of Non-luminous flame

- (i) It has three parts/zones, which have different colours.
- (ii) It is mainly blue in colour.
- (iii) It does not produce soot.
- (iv) It is very hot and noisy.
- (v) It is steady.

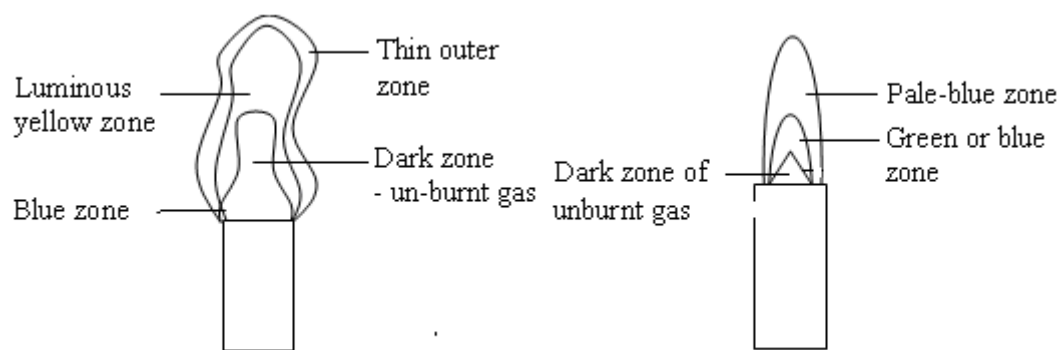


Figure 2 Luminous and Non-luminous flames

Differences between Luminous and Non-luminous flame

Luminous flame	Non-luminous flame
(i) Is produced when the air hole is closed.	- Is produced when the air hole is open
(ii) Has four zones or regions	- Has three zones or regions
(iii) Is mainly yellow in colour	- Is mainly blue/green in colour
(iv) Produces much light	- Produces little light
(v) Is not hot enough	- It is very hot
(vi) Produces soot	- Does not produces soot
(vii) It is large and wavy	- It is short and steady
(viii) Does not produce noise	- It is noisy

Table 1 Differences between Luminous and Non-luminous flames

1.45 Fire Extinguishers

These are apparatus used to fight fire out break. They should always be available in the laboratory. And they should be placed at strategic positions for easy access. The common laboratory fire extinguishers are:

- (i) Water gas
- (ii) Sand
- (iii) Blanket and
- (iv) Water

The last three are easy to use except the first one which requires the following procedures.

Procedure

- Remove the safety pin.
- Aim or direct horn at the base of fire.
- Squeeze or Press the handle to discharge and
- Sweep.

Memory Aid:

Use the word “**RASS**” to remember the principle of operating the water gas extinguisher.

R – Remove pin **A** – Aim **S** - Squeeze **S** – Sweep

Laboratory Values (Safety Precautions)

When an experiment is being done, it **SHOULD NEVER** be assumed to be completely harmless. Otherwise it may lead to accidents. Most common accidents in the laboratory include:

- Glass cuts, burns, eye damage, fire, gas poisoning & explosions.

To avoid these accidents from occurring, the following laboratory rules should be followed strictly.

1. Do **NOT** enter the laboratory without the teacher's permission.
2. Do **NOT** enter the laboratory in slippers or open shoes.
3. Enter the laboratory orderly and do **NOT** scramble over space.
4. **LABEL** all the chemicals you are using to avoid confusion.
5. **READ** and **FOLLOW** instructions carefully before you perform any experiment.
6. **ALWAYS** replace the stopper of the reagent bottle after you have finished using the required reagent from it and store it in its proper place.
7. **NEVER** run while you are in the laboratory.

8. **ALWAYS** wash off any chemical that gets on your skin immediately with a lot of water.
9. **REPORT** accidents involving cuts and burns immediately to the teacher or laboratory technician.
10. **NEVER** taste, eat, drink, or smoke in the laboratory.
11. **NEVER** look directly into reaction vessels where a chemical reaction is taking place.
12. **NEVER** smell substance or gases directly. Instead direct the fumes carefully toward your nose with your hand.
13. Experiments in which poisonous gases are produced **MUST** be carried out in a fume chamber or in open air.
14. When heating a substance in a test tube/boiling tube **NEVER** direct the open end of the tube towards your self or anybody else.
15. **ALWAYS** hold the tube using a test tube holder when heating.
16. Put off flames that are **NOT** in use or if in use and keep it away from inflammable substances, especially volatile liquids.
17. **ALWAYS** work in a clean environment e.g. any spillage of water/liquid should be cleaned immediately.
18. **ALWAYS** consult your teacher before trying any experiment of your choice.
19. After completing your experiments, **DISPOSE** the used chemicals, clean all the apparatus you have used and return them to their correct places.
20. **ALWAYS** wash your hands thoroughly with detergent before you leave the laboratory.

Self-Check 1.0

1.
 - (a) Define *Chemistry* ?
 - (b) Give **three reasons** why it is important to study Chemistry?
2. State any **five** laboratory rules.
3. Explain why:
 - (a) A student should not carry out an experiment without approval from the teacher concerned?
 - (b) It is dangerous for students to fight or play while in the Chemistry Laboratory.
 - (c) It is not allowed for students to mix up any chemicals to see what happens without the approval from the concern teacher?
 - (d) A student should not taste/eat anything he or she finds in the Chemistry Laboratory without permission from the concerned teacher?
4. Name and draw the diagrams of any **four** common laboratory apparatus you know.
5. Define the following terms.
 - (a) Hypothesis.
 - (a) Scientific Law (Principle).
 - (b) Scientific Theory.
6. Name two types of apparatus used for heating in a Chemistry laboratory.
7.
 - (a) What is a flame?
 - (b) Name two types of Bunsen flames you know.
 - (c) Give four differences between the flames you have named in (b) above.
8. The following diagram in figure 1 is of a Bunsen flame use it to answer the questions that follow.

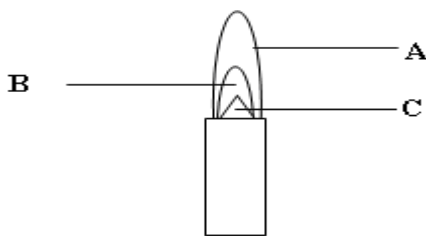


Figure 1

- (a) Name the zones A, B and C.
- (b) Identify the Bunsen flame shown.
- (c) State the condition under which such a flame is produced.
- (d) Give any **two** advantages of using the flame for heating chemicals in the Chemistry laboratory.

9. The following diagram in figure 2 represents a flame of a Bunsen burner. Use it to answer the questions that follow.

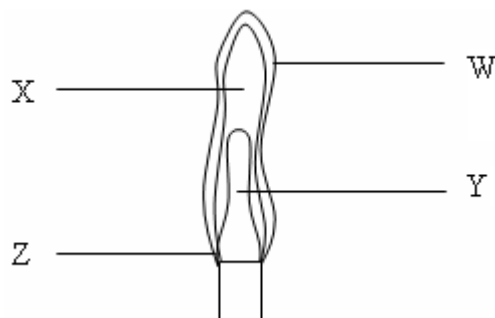


Figure 2

- (a) Name the four zones marked W, X, Y and Z.
- (b)
 - (i) Name the type of the flame.
 - (ii) State the condition under which this type of flame you have named is produced.
 - (iii) Give any *two* disadvantages of using this type of flame for heating chemicals in the Chemistry laboratory.

CHAPTER TWO

STATES OF MATTER

Learning objectives

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

1. (a) State the three states of matter.
(b) Differentiate between the three states of matter.
 2. (a) State: (i) Dalton's Atomic theory of Matter.
(ii) Kinetic Theory of Matter
 3. (a) Explain the effect of heat on matter.
(b) Define the following terms:
 - Boiling, melting, evaporation, boiling point, melting point, freezing point.
 - Sublimation.
 4. Graphically represent changes in state of matter during:
 - Heating and
 - Cooling
-

2.0 MATTER

Matter is anything that occupies space and has mass. Matter can not be destroyed nor created. That is it can not be obtained from nothing nor can it be changed to nothing. However, it can change from one state to another.

2.10 States (Phases) of Matter

Matter exists in three states namely:

- | | | | |
|-------|--------|------|-----------------------|
| (i) | Solid, | e.g. | stone, wood etc. |
| (ii) | Liquid | e.g. | water, paraffin etc. |
| (iii) | Gas | e.g. | oxygen, nitrogen etc. |

Each of these states is made of so many tiny particles called *atoms*. The arrangement of the particles and the magnitude of the forces holding them make one state different from other states.

2.20 Differences in the three states of matter

(a) Solids

- (i) The molecules/particles are closely packed and are arranged in a regular pattern called *lattice*.
- (ii) The forces holding the particles are strong.
- (iii) The molecules are not free to move because of the strong forces of attraction.
- (iv) Solids have a definite shape and volume.
- (v) There is very little diffusion in solids.
- (vi) Due to the close packing of the particles, solids cannot be compressed.
- (vii) The molecules vibrate about a fixed position (i.e. the particles cannot move from one position to another). Heat makes the vibration stronger until, at the melting point, particles break free of the forces holding them in fixed positions and then become liquid.

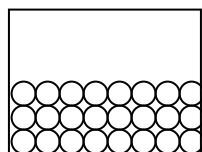
(b) Liquids

- (i) The molecules in liquid are fairly close to each other and are in irregular pattern.
- (ii) The forces of attraction are weak but are strong enough to hold them together. The molecules move randomly throughout the liquid. However, at the surface of the liquid, these forces are weaker, some energetic molecules move fast enough, and leave the liquid, i.e. evaporation occurs.
- (iii) Liquids have definite volume, but no definite shape. They take the shape of the container and then acquire a definite volume.
- (v) Liquids have slow diffusion.
- (vi) They cannot be compressed.
- (vii) Heat makes the movement of the particles faster and stronger, and evaporation occurs more easily. At the boiling point, the particles break free from the forces and the liquid becomes a gas.

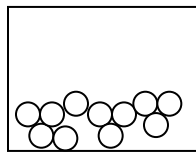
(c) Gases

- (i) The particles are far apart from each other.
- (ii) The forces holding the molecules are very weak.
- (iii) The particles move randomly and at comparatively high velocities.
- (iv) Gases have no definite shape and volume. They fill the whole container in which they are placed.
- (v) There is quick diffusion in gases.
- (vi) Gases are quite easy to compress. This is because there are large spaces between the particles.

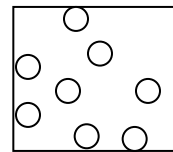
The diagrams in figure 2 show a model representation of the three states of matter



(a) Solid



(b) Liquid



(c) Gas

Figure 2 Arrangement of particles in the three states of matter

(d) Dalton's Atomic theory of Matter.

The theory states that:

1. *Matter is made up of tiny particles called atoms.*
2. *Atoms can neither be created nor destroyed.*
3. *The atoms of the same element are alike (e.g. in mass, colour, density) but they differ from those of other elements.*
4. *Two or more atoms may combine to form a molecule and they do so in small whole numbers.*

2.30 Kinetic Theory of Matter

States that:

1. *Matter is made of small particles called atoms.*
2. *The particles in liquid and gases are in a constant random motion.*
3. *Increases in temperature, causes the average speed of particles to increase, that is, the average kinetic energy increases.*
The particles (molecules) in gases move faster than those of liquids and those of liquids move faster than of solids do.

Assumptions of Kinetic Theory of Gases

1. *Molecules of gases move in straight line with a very great speed or velocity until they collide with each other or with the walls of the container. Pressure of a gas is due to collisions of the gas molecules with the walls of the container. This pressure increases as the temperature increases.*
2. *The volume of the gas molecules is negligible compared to the total volume of the container; so the molecules can be taken to be points of negligible volume.*
3. *The forces of attraction between the gas molecules are negligible.*
4. *The average kinetic energy of the molecules is a measure of the temperature of the gas.*
5. *The collisions of the molecules are perfectly elastic. (That is, when molecules collide, there is no loss in kinetic energy).*

2.40 Change of State

Change of state refers to change of matter from one form to another. For example, change from solid to liquid, Liquid to gas or vice versa. Change of state is caused by addition heat energy to or subtraction of heat energy from a body.

(a) Effect of Heat on Matter

(i) Heating a Solid

When a solid is heated, its particles acquire kinetic energy and vibrate more violently. The particles continue to vibrate until a point is reached when the vibration overcome the binding force (forces of attraction between the particles). The crystalline structure corrupts and the particles become mobile. At this point, the solid *melts* or *fuses* (i.e. changes to liquid) at a constant temperature called *melting point* and the process is called *melting* or *fusion*.

On further heating of the liquid, the molecules of the liquid acquire increased kinetic energy. The kinetic energy continues to increase until they overcome the forces of attraction between the molecules. At this point, the liquid *boils* (i.e. changes to vapour) at a constant temperature called the *boiling point* and the process is called *boiling* or *evaporation*.

(ii) Effect of Cooling

When the temperature of a gas or vapour decreases by cooling, the kinetic energy of the particles reduces. Since the kinetic energy of the particles is directly proportional to their velocity, the average velocity (speed) of the particles gradually decreases until the forces of attraction between the particles build up and the particles come closer i.e. they condense. At this point, the gas changes to liquid at a constant temperature. The process is called *condensation*.

Cooling the liquid further causes more loss in kinetic energy until eventually the particles settle to form a solid at a constant temperature called *freezing point*. The process involved is called *freezing* or *solidifying*.

2.41 Sublimation

When some solids are heated, the particles acquire more kinetic energy and overcome almost all the attractive forces between them so they escape from their neighbours and go into a gaseous state (i.e. the solid changes to vapour). We say the solid sublimates. At low temperature, the molecules lose their kinetic energy and slow down to form solid. Again, we say the vapour sublimates, and the process is called *sublimation*.

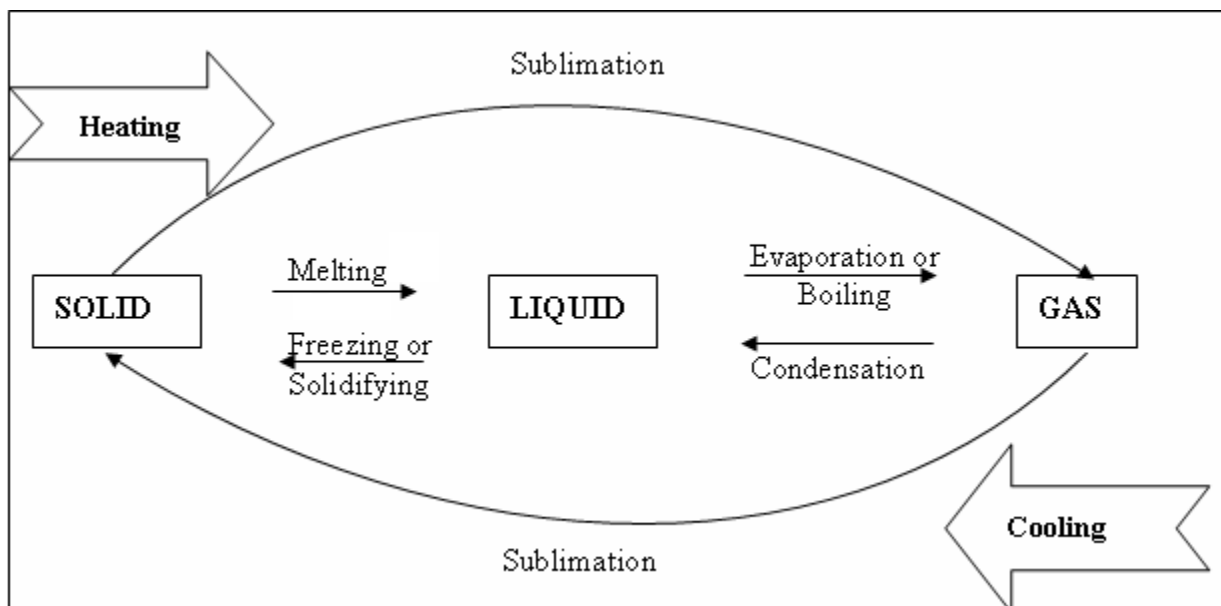
Definition

Sublimation is the change of state from solid to vapour and vice versa, without passing through the liquid state.

Examples of substances that sublime on heating are:

- (i) Ammonium chloride.
- (ii) Iodine.
- (iii) Iron III Chloride.
- (iv) Aluminium chloride.
- (v) Mercury II chloride.
- (vi) Dry ice (Solid carbon dioxide).

Figure 2.1 Diagrams showing interconversions between the three states of matter

**2.50 Graphical Representation of Physical changes during heating and cooling**

The graphs of temperature against time intervals for heating and cooling are called *heating* and *cooling curves* respectively.

- (a) **Heating Curve** - is a graph obtained by plotting temperature of a substance as it heats up against time interval.

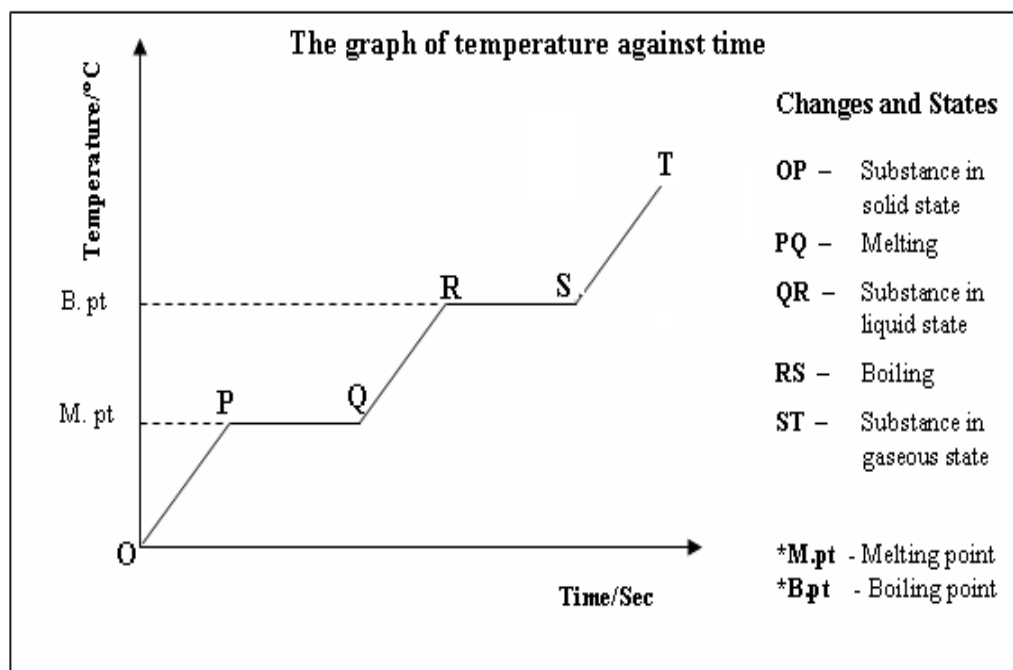
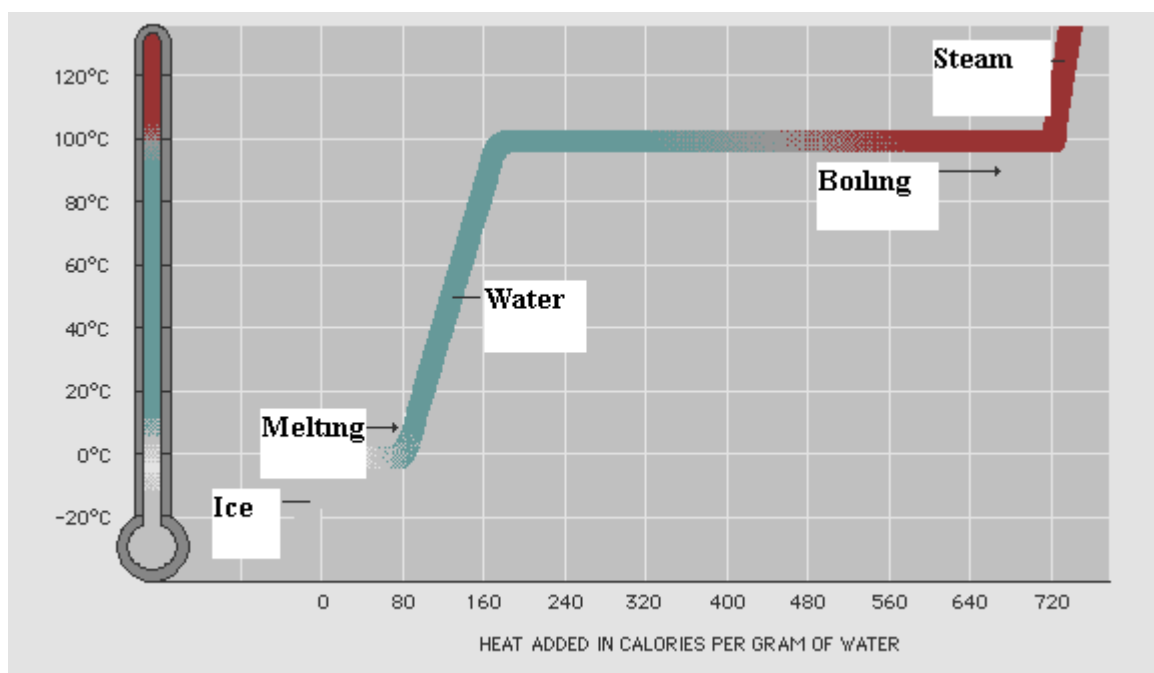


Figure 2.2 Effect of heat on a pure solid substance at a temperature of 0°C

An example of such a graph is shown below.

Temperature versus heat added to water in form of ice



The graph represents the temperature change that occurs when heat is added to water. At 0°C and at 100°C , you can add heat to water without changing its temperature. This “*latent heat*” breaks bonds that hold the molecules together but does not increase their kinetic energy.

In heating water from ice to steam, it passes through *melting* and *boiling or evaporation*. The important terms are defined as follows:

1. **Melting:** *is a process by which solid changes into a liquid.*
E.g. ice changing into water. The change in state occurs at a constant temperature called *melting point*.
2. **Melting Point:**
Melting point is a specific temperature at which a solid changes into a liquid.
3. **Boiling:** *Boiling is a rapid change of a liquid into vapour on heating.*
It occurs at a specific temperature called *boiling point* and takes place throughout the liquid.
4. **Boiling Point:**
Boiling point is a specific temperature at which a liquid changes into gas on heating. E.g. the boiling point of water is 100°C .
5. **Evaporation:**
Evaporation is a process by which a liquid changes to gas.
It takes place on the surface of a liquid at all temperatures.

(b) **Cooling Curve** - *is a graph obtained by plotting temperature of a substance against time interval as it cools down.*

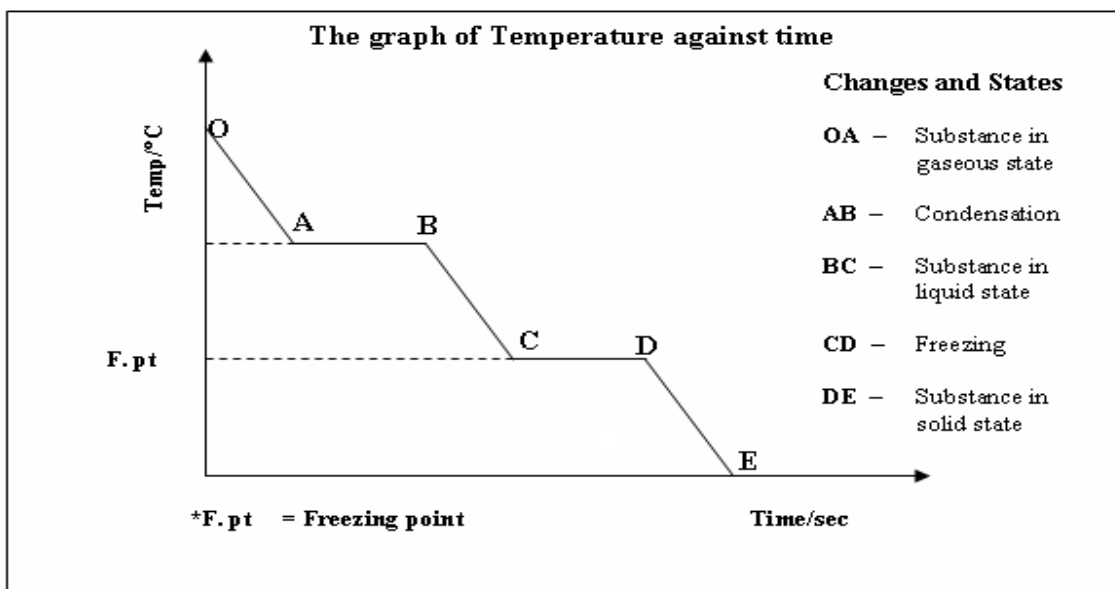


Figure 2.3 Cooling Curve for a pure substance in a gaseous state

To determine the state of matter at a given temperature when the melting point and the boiling point are given

Arrange the temperature values (the given temperature, the melting point and the boiling point) in ascending order.

- (i) *If the given temperature is before the melting point, the substance is in solid state.*
- (ii) *If the given temperature is after the melting point but before the boiling point, the substance is in liquid state.*
- (iii) *If the given temperature is after the boiling point, the substance is in gaseous state.*

Example

Consider the table below in which the melting and boiling points of substances P, Q and R are given.

Substance	Melting Point °C	Boiling Point °C
W	+0	+100
X	+1083	+2567
Y	-33	+13

- (a) In what state are the substances at a temperature of 50 °C?
- (b) Which substances are in liquid state at a temperature of 10 °C?
- (c) State the substance which is a metal and give a reason to support your answer.

Solution

- (a) W - In liquid state.
X - In solid state.
Y - In gaseous state.
- (b) W and Y
- (c) X - Because the metals have high melting and boiling points.

2.6 Types of Changes of Matter

Matter undergoes two types of changes, namely:

- (i) Physical change (temporary change).
- (ii) Chemical change (permanent change).

(a) Physical Change

A physical change is a change where no new substance is formed.

The change is reversibly. The properties of the substance remain the same.

Examples of physical changes are:-

- Melting of ice, fat etc.
- Evaporation of liquid.
- Subliming of iodine.
- Dissolving of sugar or salt in water.
- Magnetizing iron filings or nails.

(b) Chemical Change

A chemical change is a change where a new substance is formed.

The change is irreversible. The products have different properties from those of the reactants.

Examples of chemical changes:-

- (i) Rusting of iron.
- (ii) Burning of magnesium, charcoal, wood, paper, etc.
- (iii) Acid reacting with a base or metal.

The table showing the differences between physical and chemical changes

Physical change	Chemical change
(i) No new substance is formed.	- New substance is formed.
(ii) No heat or light is absorbed or given out.	- Heat or light is absorbed or given out.
(iii) Not accompanied by change in mass occurs.	- Accompanied by change in mass.
(iv) The change is reversible or can be reversed to its original state.	- The change is irreversible or cannot change to its original state.

Table 2. Differences between Physical and Chemical changes

2.70 Properties of Matter

Matter has two types of properties, namely:

- (i) Physical properties
- (ii) Chemical properties

(a) Physical Properties

These are measurable properties of matter. Examples of physical properties of matter include:

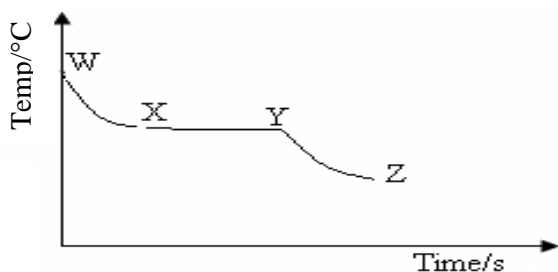
- Physical constants such as; melting point, boiling point and freezing point.
- Smell, colour, taste, density, solubility etc.

(b) Chemical Properties

These are properties of matter that tell us what happens to the substance when it undergoes a chemical change.

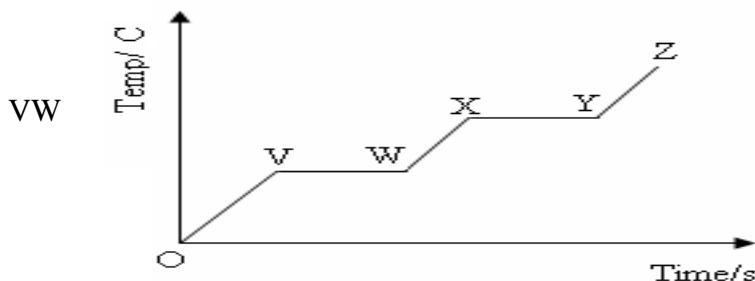
Self check 2.0

1. The particles in a solid at room temperature are
A. close together and vibrating. B. close together and stationary.
C. far apart and moving at random. D. close together and moving at random.
2. The graph in the figure shows a cooling curve of a pure substance. The substance is all in a solid state between



- A. W and X
- B. W and Y
- C. X and Y
- D. Y and Z

4. The graph in the figure is that of the temperature against time for a substance, which is heated at a constant rate. Which part of the graph corresponds to the situation when the molecules of the substance have the highest average kinetic energy?



- A. OV
- B. YZ
- C. VX
- D. YZ

4. At room temperature, air is less dense than water because air molecules
 - A. move faster
 - B. are smaller
 - C. have greater force of attraction
 - D. are more widely separated from each other.

5. When a liquid is heated
 - A. its density decreases
 - B. boiling occurs at all temperatures
 - C. its molecules move with the same speed
 - D. evaporation takes place throughout the liquid

6. It is more difficult to compress a liquid than a gas because
 - A. the speed of liquid molecules is lower than that of gas molecules.
 - B. liquid particles attract one another when compressed while gas particles repel each other.
 - C. the distances between liquid particles are less than those between gas particles.
 - D. liquid molecules repel one another when compressed while gas molecules attract one another.

The boiling points and the melting points of substances P, Q, R and S are given in the table below. Use it to answer questions 7 - 10.

Substance	Melting Point °C	Boiling Point °C
P	+17	+45
Q	+651	+1100
R	-33	+13
S	0	+100

7. Which one of the substances is a gas at room temperature (25°C)?
 - A. P
 - B. Q
 - C. R
 - D. S

8. Which one of the substances is a solid at room temperature (25°C)?
 - A. P
 - B. Q
 - C. R
 - D. S

9. Which one of the substances is a gas at room temperature (-10°C)?
 - A. P
 - B. Q
 - C. R
 - D. S

10. Which one of the substances is a metal?
 - A. P
 - B. Q
 - C. R
 - D. S

11. Which one of the following statements is true when a solid is changing into a liquid?
- A. The temperature increases as energy is increased.
 - B. The temperature remains constant as energy is increased.
 - C. The temperature drops as energy is increased.
 - D. The temperature remains constant as the energy decreases.
12. Which of the following statements is incorrect when a tin containing air tightly sealed is heated?
- A. The average speed of molecules increases.
 - B. The molecules of air hit the walls of the tin harder.
 - C. The molecules of the air strike the walls of the tin less often.
 - D. The pressure inside the tin increases.

SECTION B

13. Define the following terms in each case give an example.
- (a) Physical change.
 - (b) Chemical change.
 - (c) Indicate clearly whether a chemical or physical change is involved in the following processes. Give reason for your answer.
 - (i) Rusting of iron.
 - (ii) Addition of sodium chloride (common salt) into water.
 - (iv) The heating of ammonium chloride.
 - (v) Decaying of organic matter.
14. (a) Differentiate between physical and chemical change.
- (b) Sketch a graph of temperature against time when ice is heated until it changes to steam.

CHAPTER THREE

ELEMENTS, COMPOUNDS AND MIXTURES

Learning objectives

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

1. (a) Define the following terms:
 - Elements, Compounds, Atoms, Molecules and Mixture.
 - (b) Differentiate between Compounds and Mixtures.
 - (c) State: (i) the types of Mixtures and
(ii) the methods of separating mixtures.
 2. (a) Define: (i) Miscible liquids and
(ii) Immiscible liquids.
 - (b) Describe how to separate miscible liquids and immiscible liquids.
 3. (a) Define the term: Chromatography.
 - (b) Identify the different colours from a Chromatogram.
 4. Explain: - how to test purity of a material.
- Effect of impurity on physical properties of matter.
-

3.0 Introduction

All substances belong to one of the three classes of matter, namely:

- (i) Element
- (ii) Compound and
- (iii) Mixture.

Elements

An element is a substance that cannot be split into two or more simpler substances by any chemical means.

It is a single substance and therefore is said to be pure. Examples of elements are:

- Oxygen, hydrogen, nitrogen, carbon, copper etc.

Compounds

A compound is a substance which consists of two or more elements chemically combined together.

It can be split into simpler substances. The simpler substances may be elements or compounds or both.

Examples of compounds include:

- Water, carbon dioxide, carbon monoxide, acids, salts, bases, etc.

Elements and compounds are made up of atoms.

(a) Atoms

An atom is the smallest particle of an element that can take part in chemical reaction.

(b) Molecules

A molecule is a group of atoms that are chemically combined together and is capable of existing on its own.

A molecule may contain atoms of the same element (e.g. O₂) or atoms of different elements (e.g. H₂O). A substance made of molecules that include two or more atoms of the same kind is called an *element*. While a substance made of molecules that include two or more different chemical elements is called a *molecular compound*.

3.1 Mixtures

A mixture is substance that consists of two or more elements or compounds that are physically combined together.

(a) Types of Mixtures

There are three types of mixtures, namely:

- Natural mixtures: E.g. Air, milk, blood.
- Artificial mixtures: E.g. Mixture of salt and sand, mixture of iron and sulphur, ground-nut and simsim etc.
- Special mixtures: E.g. Alloy.

(i) Alloy

An alloy is a mixture of two or more different metals in a definite proportion.

Examples of alloys are: Brass, Bronze, Solder, Steel, Duralumin etc

Alloy	Composition of Alloy (Metals present)
Brass	Copper and zinc
Bronze	Copper and tin
Solder	Tin and lead
Steel	Iron and carbon plus small amount of chromium, nickel and manganese.
Duralumin	Aluminium, magnesium and copper

Table 3 Common alloys and their respective metals

Differences between compound and mixture

Compound	Mixture
(i) Is a substance that consists of two or more elements chemically combined.	- Is a substance that consists of two or more elements or compounds physically combined together
(ii) Constituents can be separated by chemical means.	- Constituents can be separated by physical means
(iii) The properties of the constituents e.g. colour, density, are quite different from the properties of the elements in it.	- The properties of the mixture, e.g. density, colour are average of the constituents in it.
(iv) Energy is usually absorbed or given out when a compound is being made.	- Energy is not usually given out or absorbed when a mixture is being made.
(v) Constituents are combined in definite proportion by mass.	- Constituents are combined in any proportion by mass.

Table 3.1 Differences between compound and mixture

(b) Classification of Mixtures

Mixtures are classified into six classes, namely:

- Solid - solid,
- Solid - liquid
- Solid - gas
- Liquid - liquid
- Liquid - gas
- Gas - gas

3.11 Separation of the Components of Mixtures

Mixtures can be separated using the following methods.

Filtration, evaporation, decantation, using a separating funnel, distillation, fractional distillation, crystallization, magnetic separation, sublimation, paper chromatography, winnowing and hand picking. The method applied depends on the type of the mixture.

Type of Mixture	Appropriate Method (s) that can be used
Solid – Solid	Magnetic, Winnowing, Hand picking,
Solid – Liquid	Filtration, Decantation, Crystallization
Liquid – Liquid	Distillation, Fractional Distillation
Liquid – Gas	Boiling e.g. boiling water with a dissolved gas
Gas – Gas	Fractional distillation

Table 3.2 Types of mixtures and the major methods used to separate them.

(a) To separate solid from a liquid

A solid is separated from a mixture of solid and liquid by the method of *filtration* and *decantation*.

(i) Filtration

Filtration is the process of separating a suspended solid from a liquid or solution in which it is already suspended by straining it through a porous medium, such as filter that can be penetrated easily by liquids or solution.

The mixture to be filtered is called the suspension. The liquid that passes through the filter is called the *filtrate*; and the solid material that remains on the filter is known as the *residue*.

Experiment 3.0 To separate sand particles from muddy water

Apparatus Retort stand, filter funnel and 2 beakers

Procedure

- Clamp up a filter funnel
- Fold a filter paper in the shape of a filter funnel and fix it in the funnel.
- Place a clean beaker below the filter funnel.

- Stir the mixture of sand and water thoroughly and pour a reasonable amount into the filter paper in the filter funnel.
- Leave the apparatus to stand for some time.

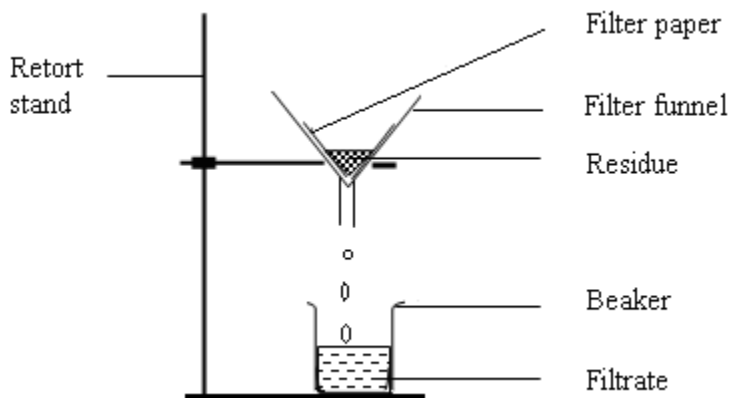


Figure 3 Separation of solid from liquid by filtration method

Observation

Droplets of colourless liquid (filtrate) collect in the beaker while sand (residue) is trapped on the filter paper.

Conclusion A mixture of sand and water can be separated by filtration.

Procedure to fold the filter paper

Half fold the filter paper two times and then open one layer to obtain a hollow cone as in diagram 3.

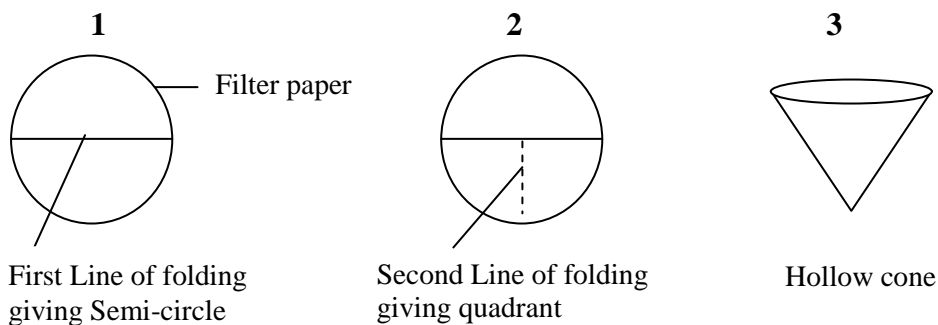


Figure 3.1 Folding a filter paper

Applications of Filtration on large scale

Filtration is used in:

- the treatment of sewage,
- the chemical-manufacturing industry,
- the brewing of coffee and
- the clarification of sugar syrup.

(ii) Decantation.

This is the separation of solid particles from liquid by gently and carefully pouring off the clear liquid into another container so as not to disturb the sediment. The settlement of the solid particles can be fastened by using a *centrifuge*.

Centrifuge:

A centrifuge is a mechanical device that uses the principle of centrifugal force to separate substances of different densities. It has a set of containers (tubes) that is spun rapidly. The solid particles settle at the bottom of the tube and the clear liquid on top. The liquid can be removed by decantation or by pipette.

(b) To separate a solute from a solution

A solution consists of a solute dissolved into a solvent e.g. sea water is a solution of salt (solute) and water (solvent). The solute is separated from the solution by a process called *evaporation* and *crystallization*.

(i) Evaporation

The solution is heated in an evaporation dish. The water evaporates and the salt remains in the evaporation dish. The final heating may be carried out by using indirect heating such as steam bath or sand bath in order to avoid loss of salt by spitting.

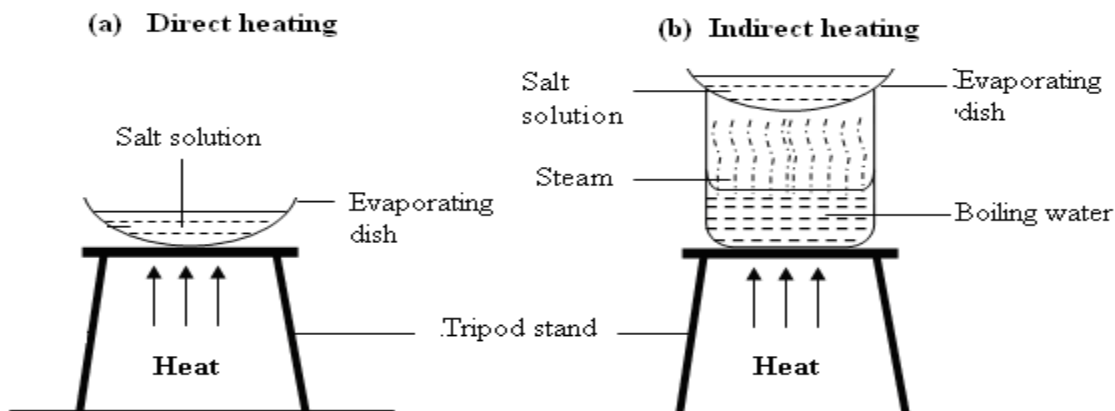


Figure 3.2 Evaporating a solution of common salt.

(ii) Crystallization.

This is the process by which a pure and large crystal of salt is obtained from a solution. The salt is heated to crystallization point (the point at which crystals begin to form on cooling). It is then allowed to cool. The large crystals formed are then filtered, washed and dried.

(c) To separate a solvent from a solution or miscible liquids

A solvent is separated from a solution by a process called *distillation*.

Distillation

Distillation is the process of heating a solution until its more volatile constituent reaches its boiling point and pass into the vapour phase, and then cooling the vapour to recover the constituent in liquid form by condensation. The main purpose of distillation is to separate a mixture of two liquids by taking advantage of their different boiling points. The condensed liquid is called a *distillate*.

Experiment 3.1

To separate solvent from a solution e.g. Pure water from impure water

Apparatus/Requirements

Retort stand, Distillation flask, Tripod stand, Thermometer, Lie big condenser, Conical flask, Running cold water, Source of heat

Procedure

- Pour impure water (e.g. sea water) into a distillation flask.
- Insert a thermometer through a cork and then cork the distillation flask with it.
- Pass the cold water from the lower part of the condenser until it flows out from the top of the condenser into a sink. (This is to ensure that there is complete condensation).
- Heat the water in the flask until it boils.

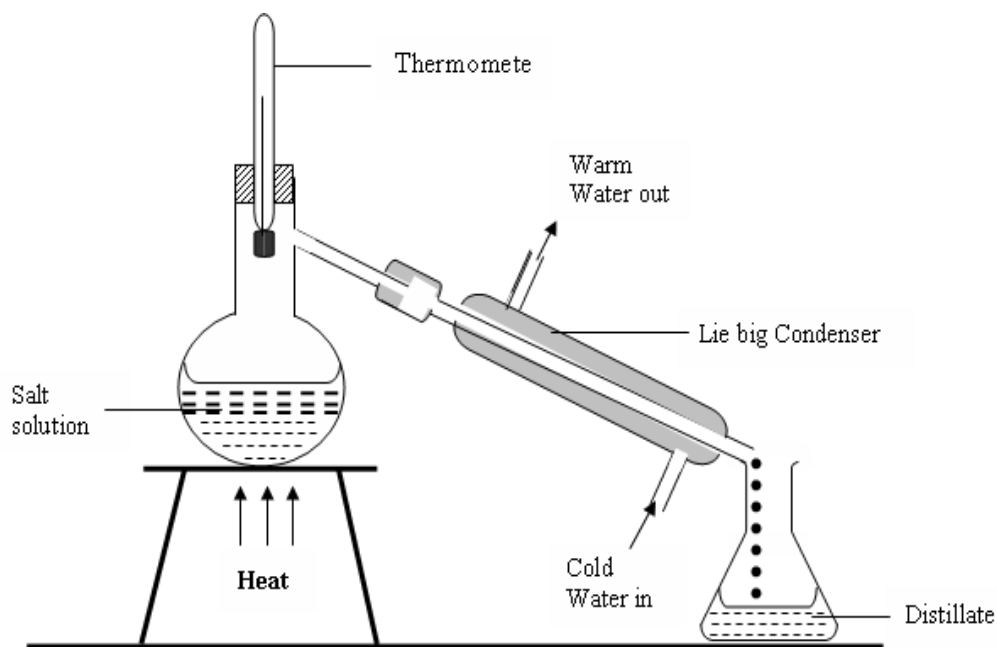


Figure 3.3 Laboratory setup for simple distillation

Observation

- The level of the thermometric liquid rises for some time and then remains constant.
- The steam rises up and on reaching into the condenser, condenses to (i.e. changes to) a liquid.

Conclusion: A solvent can be separated from a solution by distillation.

(d) To separate a mixture of two liquids

A mixture of two liquids may be *miscible* or *immiscible*.

(i) Miscible liquids

These are liquids that mix completely and form no layers. Examples of miscible liquids are:

- Ethanol & water
- Ethanol & propane
- Liquid air and
- Petroleum.

Miscible liquids are separated by a method called *fractional distillation*.

Fractional Distillation

Definition

Fractional distillation is a process of separating a mixture of two or more miscible liquids which have different but close boiling points by distillation, the distillate being collected as fractions which boil at different temperatures.

Fractional distillation is the same as simple distillation except the fractionating column is packed with glass beads. The glass beads provide large surface area for condensation of the incoming vapour. The liquid with the lowest boiling point distils off first and collects as a *distillate*.

Experiment 3.2

To separate a mixture of ethanol and water by fractional distillation

Apparatus/Requirements

A round bottom flask, fractionating column, thermometer, Liebig condenser, tripod stand, source of heat, glass beads, source of running cold water, 2 conical flasks.

Procedure

- Mix some ethanol and water and pour it into a round bottom flask with glass beads.
- Fit a fractionating column corked with a thermometer and connect to a Liebig condenser.
- Heat the mixture and collect the distillate each distillate as shown by their boiling points.

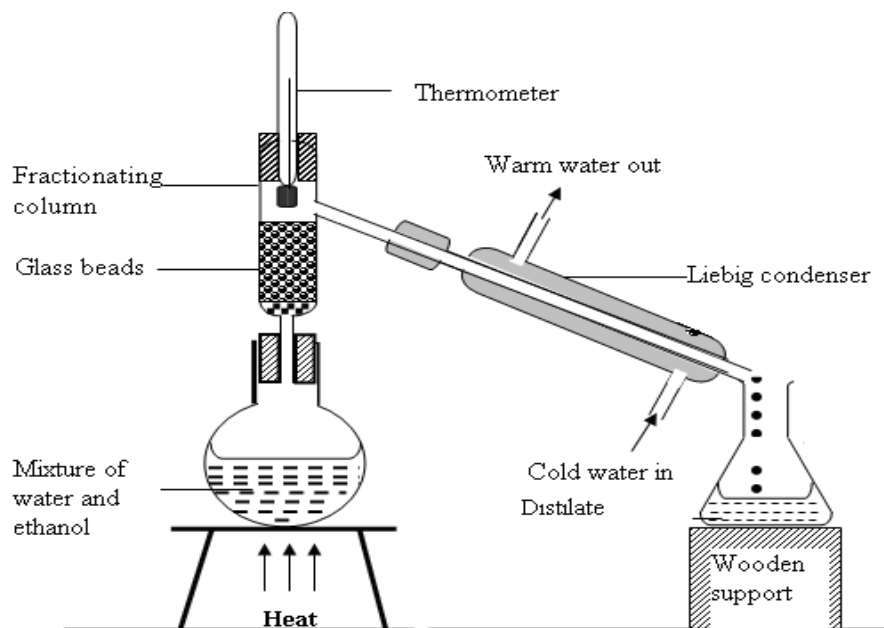


Figure 3.4 The diagram showing fractional distillation of miscible liquids

Observations

- The level of the thermometric liquid rises for some time and then remains constant.
- Steam (vapour) rises up and on reaching into the condenser, condenses to a colourless liquid called distillate and collects in the conical flask.
- Some vapour condenses on the surfaces of the broken glasses in the fractionating column and drips back into the round bottom flask.
- After some time the temperature on the thermometer increases to a new level and again becomes constant.

Conclusion: A mixture of miscible liquids can be separated by fractional distillation.

Applications of fractional distillation

Fractional distillation is applied in:

- The extraction of oil from nuts.
- The purification of water.
- Oil refinery (oil refining)

NB:

- ❖ *The glass beads provide large surface area for the vapour to condense.*
- ❖ *The conical flask reduces the loss of distillate by evaporation since it provides small surface area.*

(ii) Immiscible liquids

These are liquids which do not mix together to form a single homogenous substance. They form separate layers. Examples of liquids that are immiscible are:

- Water & oil,
- Water & paraffin,
- Mercury & water,
- Benzene & water and
- Carbon tetrachloride and water.

In the mixture of immiscible liquids, the less dense liquid floats on top of the denser liquid. Therefore the idea used in their separation is the difference in their densities. Immiscible liquids are separated by use of a separating funnel.

**Experiment 3.3 To separate a mixture of immiscible liquids
(e.g., water & paraffin)**

Apparatus

A separating funnel, 3 beakers, filter funnel if necessary, retort stand

Procedure

- Clamp up the separating funnel using the retort stand.
- Close the tap of the separating funnel.
- Pour the mixture of water and paraffin into a separating funnel using the filter funnel.
- Leave the mixture to settle for a few minutes and open the tap to collect the water and paraffin respectively in separate beakers.

The diagram showing the separation of immiscible liquids

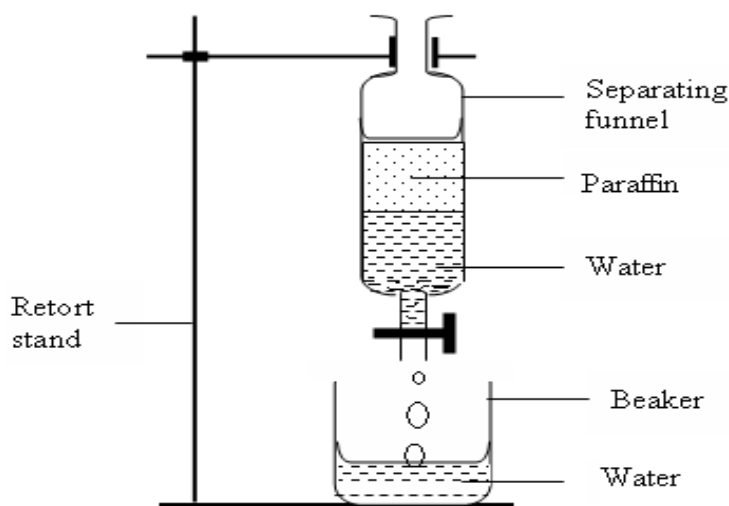


Figure 3.5 Separation of immiscible liquids by using a separating funnel

Observation: The paraffin floats on top of the water.

Conclusion: A mixture of paraffin and water can be separated by using a separating funnel.

(e) Chromatography

Chromatography is a method of separating a mixture of similar coloured solutes dissolved in a solvent by use of difference in movement of the substances in the solvent over an adsorbent material. The technique depends on the principle of selective adsorption (not to be confused with absorption), a type of adhesion and the rate of movement of the substance. The adsorbent material can be a filter paper or aluminium oxide packed in long column.

Types of Chromatography

There are two types of Chromatography, namely:

- (i) Paper Chromatography
- (ii) Column Chromatography

Chromatography is used to separate dyes or pigments in plant leaves.

Paper Chromatography

Paper chromatography is the separation of a mixture of coloured solutes dissolved in a solvent by using a porous paper as an adsorbent material.

Experiment 3.4 To separate dyes in ink by paper chromatography.

Apparatus A beaker, an ink, a strip of filter paper, solvent (ethanol), a glass rod.

Procedure

- Put a drop of ink about one centimeter from the lower end of a strip of filter paper and allow it to dry.
- Suspend the strip of filter paper vertically in a beaker containing very little solvent (Ethanol or water) such that its end near the spot is dipped into the solvent.
- Leave the apparatus to stand for some time.
- When the ethanol has nearly reached the top, remove the filter paper and allow it to dry.

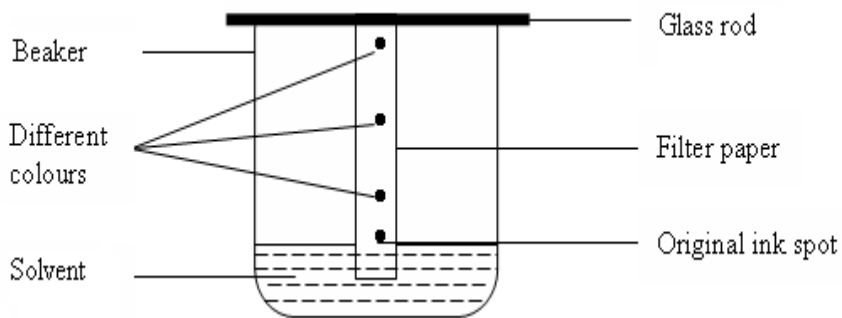
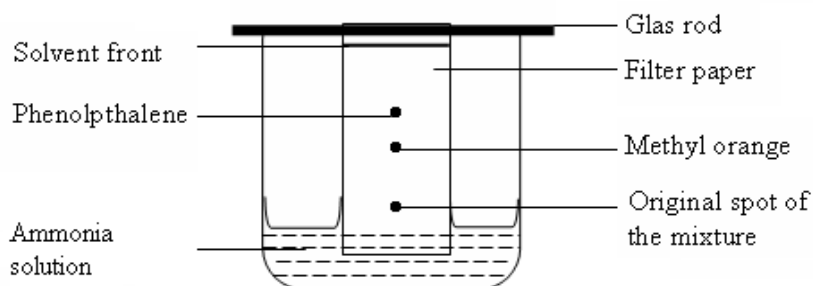


Figure 3.6 separation of dye in ink using paper chromatography

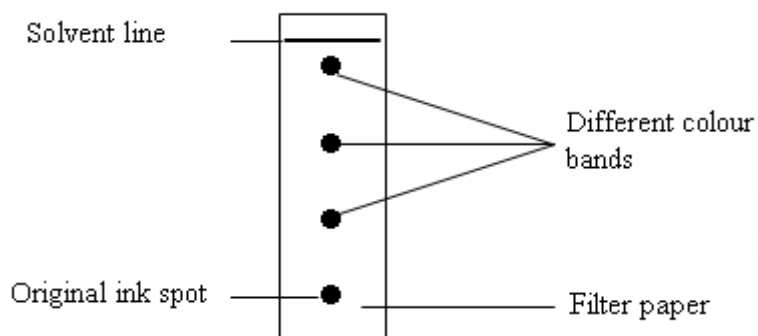
Observation

The solvent rises past the original spot and dissolves the components in the ink spot which move up the paper at different speeds giving different bands of colours.

This method is also used to separate methyl orange and phenolphthalein indicators. The mixture of phenolphthalein indicator and methyl orange is applied towards the end cut of the filter paper and then dipped into a container containing some little ammonia solution. The phenolphthalein indicator moves faster than the methyl orange. It is higher up methyl orange as shown in the diagram below.



A sample of chromatogram



- NB:**
- The higher the speed, the more soluble the dye is.
 - The coloured bands as they appear on the filter paper constitute what we call **Chromatogram**.

(f) Fractional Crystallization

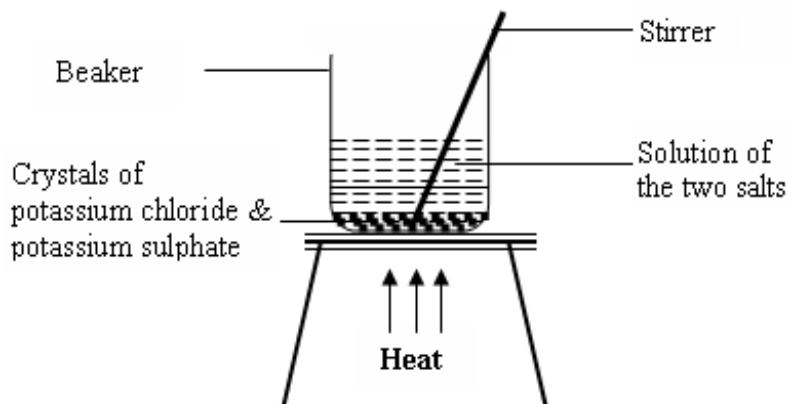
This is a method of separating a mixture by repeated partial crystallization of the solution. This method takes the advantage of different solubilities of solutes in a given solvent. Example of a mixture that can be separated by fractional crystallization is a mixture of potassium chloride and potassium sulphate.

Experiment 3.5 To separate a mixture of potassium chloride and potassium sulphate.

Apparatus A beaker, distilled water, stirrer, cold ice, source of heat.

Procedure

- Add the mixture of the potassium chloride and potassium sulphate in distilled water in a beaker.
- Warm the solution gently while stirring continuously until the solutes dissolve.



Cool the solution by placing the beaker in a larger beaker half filled with pieces of ice.

Observation

The salt with the lowest solubility, potassium sulphate, crystallizes out first.

NB: For more information about solubility of salts, see section Chapter 13 section 13.53.

(g) To separate mixture of two solids

Mixtures of two solids may be separated by the following methods.

- Sublimation, magnet, winnowing and hand picking

(i) Sublimation

Sublimation is a process where by a solid substance changes directly into the vapour phase and back into the solid state without a liquid being formed at any time. Examples of substances that sublime are:

- Ammonium chloride,
- Iodine,
- Anhydrous iron (III) chloride,
- Anhydrous aluminium chloride and
- Dry ice (solid carbon dioxide).

As a result of this property, solids which sublime can be separated or purified from those which do not sublime. Therefore, the following mixtures can be separated or purified by sublimation.

- Sodium chloride and ammonium chloride,
- Sand and iodine and
- Sodium chloride and Iron (III) chloride

Experiment 3.6

To separate a mixture of Ammonium chloride and Sodium chloride

Apparatus

A beaker, source of heat, tripod stand, cold water and round bottom flask.

Procedure

- Place the mixture in a beaker.
- Stand the beaker on a tripod stand.
- Fill cold water into a round bottom flask and place it on the beaker.
- Heat the beaker gently as shown in the diagram 3.10.

Diagram showing the separation of a mixture using sublimation method

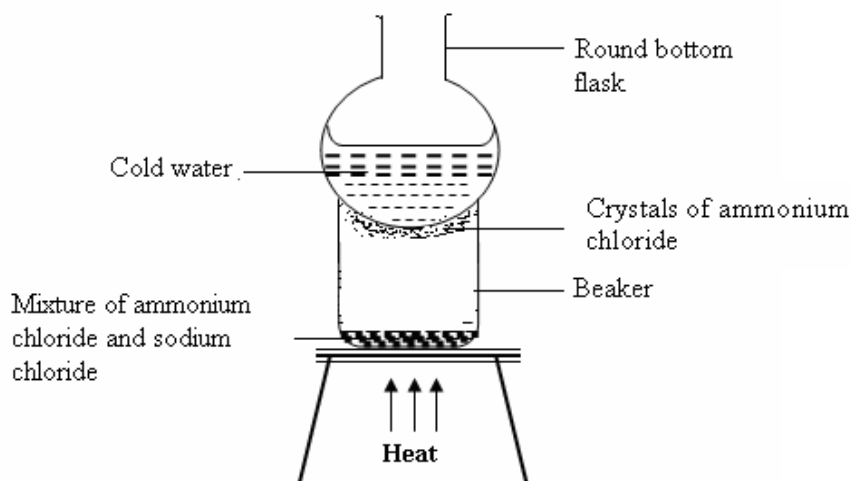


Figure 3.10

Observation

The ammonium chloride sublimates and solidifies in the cooler part of the flask as white solid.

The Sodium chloride remained in the beaker.

Conclusion

A mixture of Sodium chloride and Ammonium chloride can be separated by sublimation.

(ii) Magnetic method

A magnet can be used to separate a mixture containing magnetic and non-magnetic substances. On large scale, this method is used in industries where large and powerful electro magnets are used to separate scrap of iron from pieces of non magnetic substance.

Experiment 3.7 To separate a mixture of iron fillings and sulphur

Apparatus

A bar magnet, a piece of paper & a mixture of iron fillings and sulphur

Procedure

- Place a mixture of iron fillings and sulphur on a piece of paper.
- Gently move a bar magnet over the mixture.

Observation

The iron fillings are attracted to the bar magnet leaving behind the sulphur (yellow) powder.

Conclusion

Magnetic substance can be separated from non magnetic substance by using a magnet.

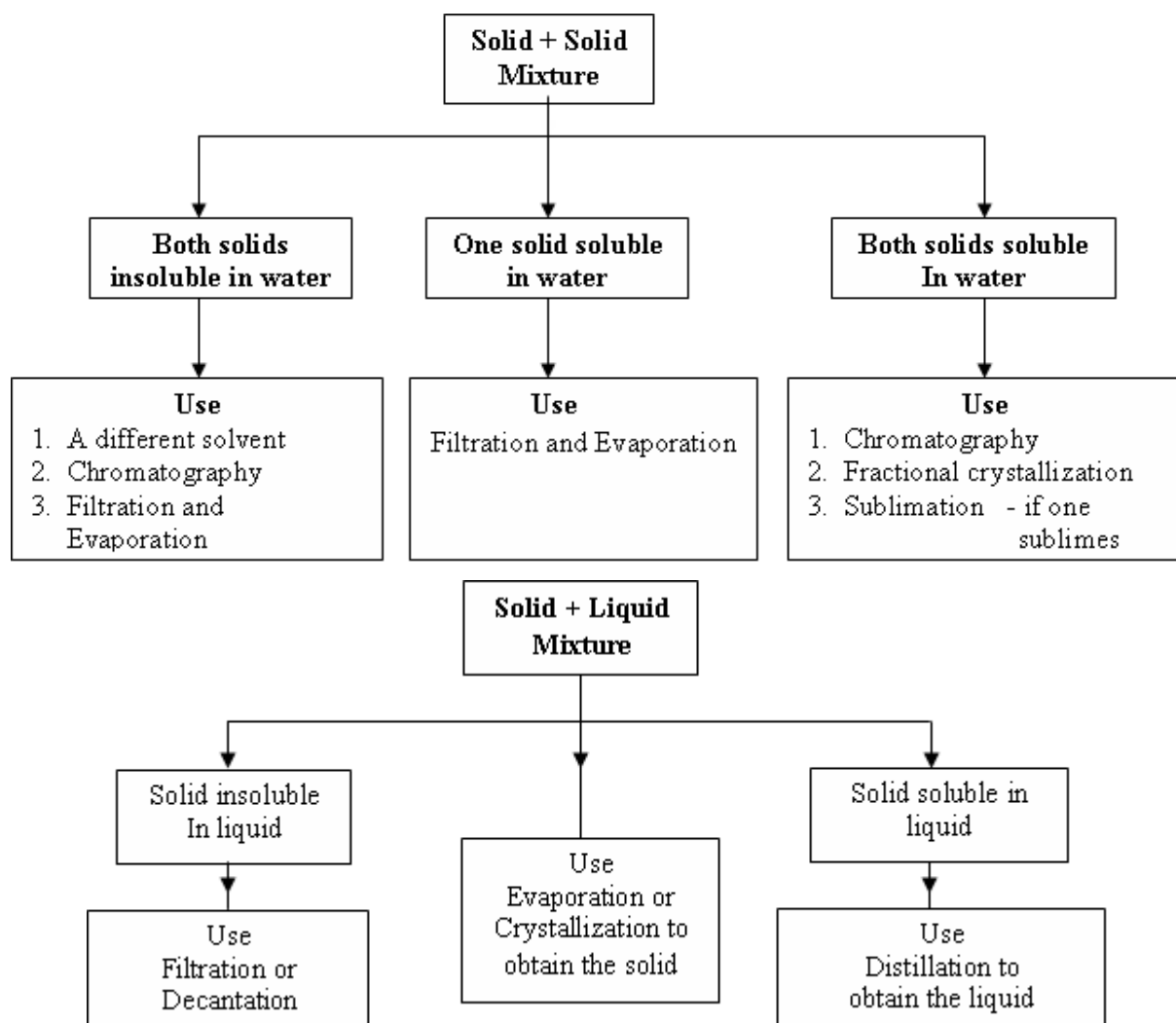
(iii) Winnowing,

This is a method used for separating a mixture of two solids one of which is light enough to be blown away by wind. e.g. a mixture of rice and husks. Wind is used as a medium of separation. The apparatus used is called a Winnowing.

(iv) Hand picking

This is the method of separating a mixture of two solids by employing the hand. In this method one of the components of the mixture which is large enough is picked off.

Summary of methods of separating substances



3.2 Pure materials

When a material exists as a single substance it is said to be pure. It is often difficult to tell whether a material is pure or impure by looking at its appearance or taste.

(a) Simple Test for purity

The purity of a substance is determined by measuring its physical constants. These are:

- (i) Freezing point.
- (ii) Boiling point and
- (iii) Melting point.

If these above constants are sharp (exact), the substance is pure and if are not exact, the substance is impure. Impure substances contain some other substances. Therefore, they are said to contain impurities.

Experiment 3.8 To determine the melting point of naphthalene

Apparatus A test tube, thermometer, source of heat, retort stand/clamp

Procedure

Place a small amount of naphthalene in a test tube.

Clamp the test tube vertically by a clamp.

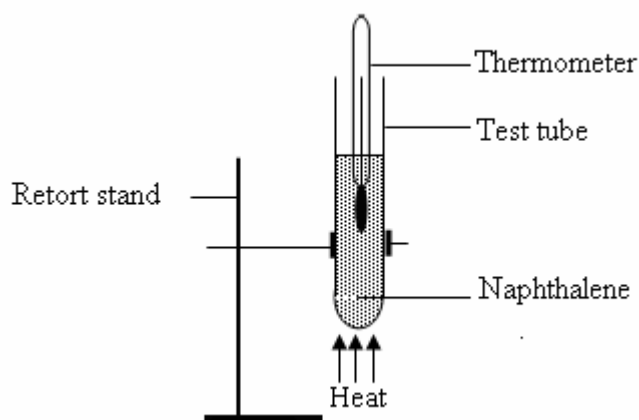
Heat the naphthalene gently using a non-luminous flame until it just melts.

Insert a thermometer in the molten naphthalene and continue to heat it until the temperature is about 100 °C.

Put off the flame and take the readings of the thermometer at minute intervals as the tube and its content are cooling.

Note and record the constant temperature value.

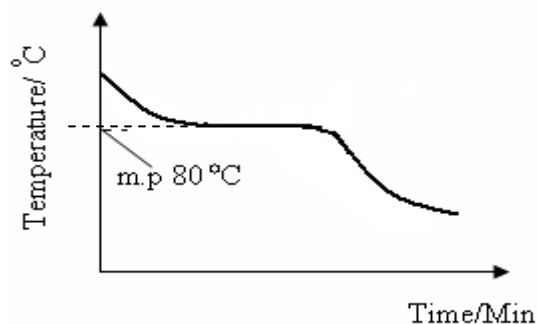
Diagram showing the determination of melting point of naphthalene



Observation

The temperature continues to drop for some few minutes and then remained constant at 80 °C (freezing point which is equal to the melting point) for some time until all the naphthalene solidifies after which the temperature begins to fall again.

The temperature changes are illustrated most strikingly by plotting a graph of temperature against time.



(b) Effect of impurity on Physical properties

The presence of an impurity alters (changes) the physical properties. An impurity:

- (i) Elevates (raises) the boiling points of liquids and
- (ii) Lowers the melting and freezing points of substances.

Self-Check 2.0

1. Which one of the following liquids is miscible with water?
A. Ethanol. B. Methyl benzene.
C. Kerosene. D. Petrol.
2. Which of the following pairs of metals is used to make solder?
A. Zinc and lead. B. Copper and aluminium.
C. Tin and lead. D. Copper and zinc.
3. A chromatogram of substances P, Q, R, S, T and a mixture **M**, was developed as shown in figure 1 below.

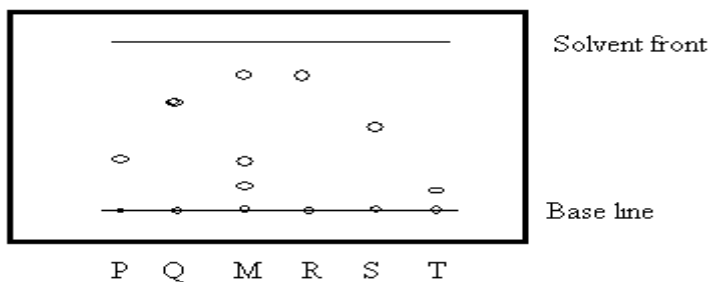
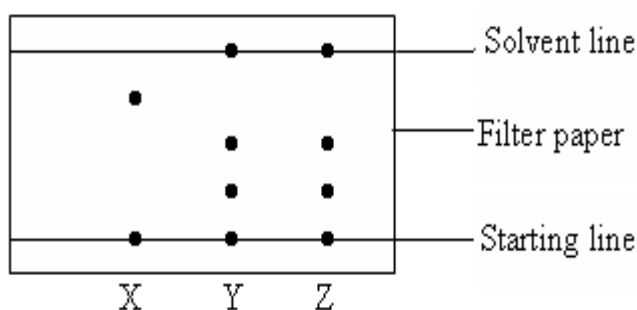


Fig. 1

- Figure 1 shows that the mixture **M**, consists of
- A. P, R and T.
 - B. R, S and T.
 - C. P, Q and R.
 - D. Q, R and S.
4. The components of ink can be separated by
A. distillation. B. chromatography.
C. filtration. D. electrolysis.
 5. Which one of the following mixtures can be separated by sublimation?
A. A mixture of ammonium chloride and magnesium chloride.
B. A mixture of sodium chloride and sodium carbonate.
C. A mixture of calcium chloride and calcium carbonate.
D. A mixture of lead (II) chloride and iron filling.
 6. Fractional crystallisation can be used to separate a mixture in solution. The substances in the mixture must have
A. a large difference in freezing points. B. very low boiling points.
C. a hygroscopic property. D. very high molecular masses.

7. The separation of substances that make up ink by chromatography depends on the
- solubilities of the substances in the solvent.
 - size of the chromatography paper.
 - freezing points of the substances.
 - osmotic pressure of the solution of ink.
8. The best method for separating a mixture of ammonium chloride and sodium chloride is
- decantation.
 - filtration.
 - distillation.
 - sublimation.
9. The results of the chromatography of the dyes used in making three sweets X, Y and Z is shown in figure 1 below.



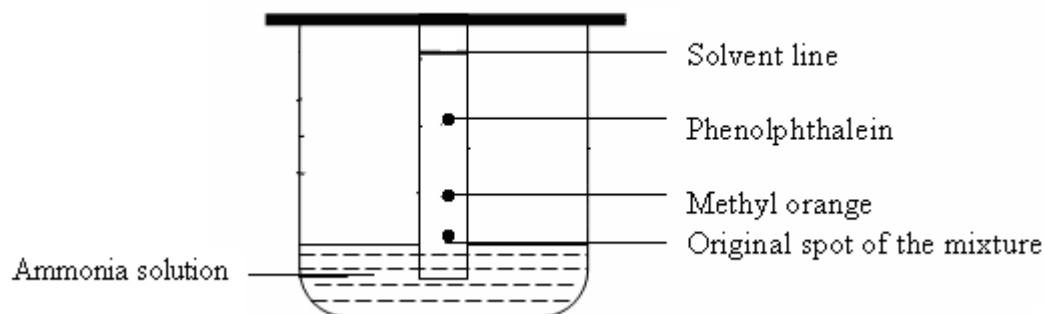
The number of dyes used in making the sweets is

- 3
 - 4
 - 5
 - 9
10. Which one of the following substances does **not** sublime when heated?
- Ammonium chloride.
 - Iron (III) chloride.
 - Iodine.
 - Copper (II) oxide.

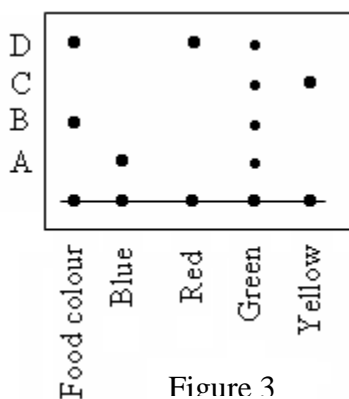
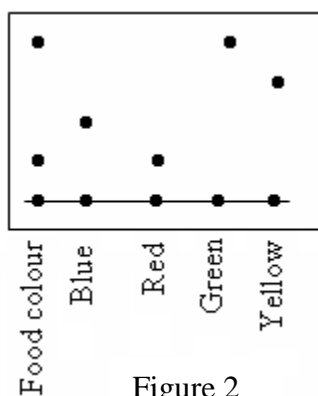
SECTION B

11. Name **one** process by which the components of the following mixtures can be separated:
- Pigments of green leaf.
 - Water and ethanol.
 - Iodine and potassium chloride.
 - Copper (II) sulphate and sand.
 - A mixture of sulphur and iron fillings.

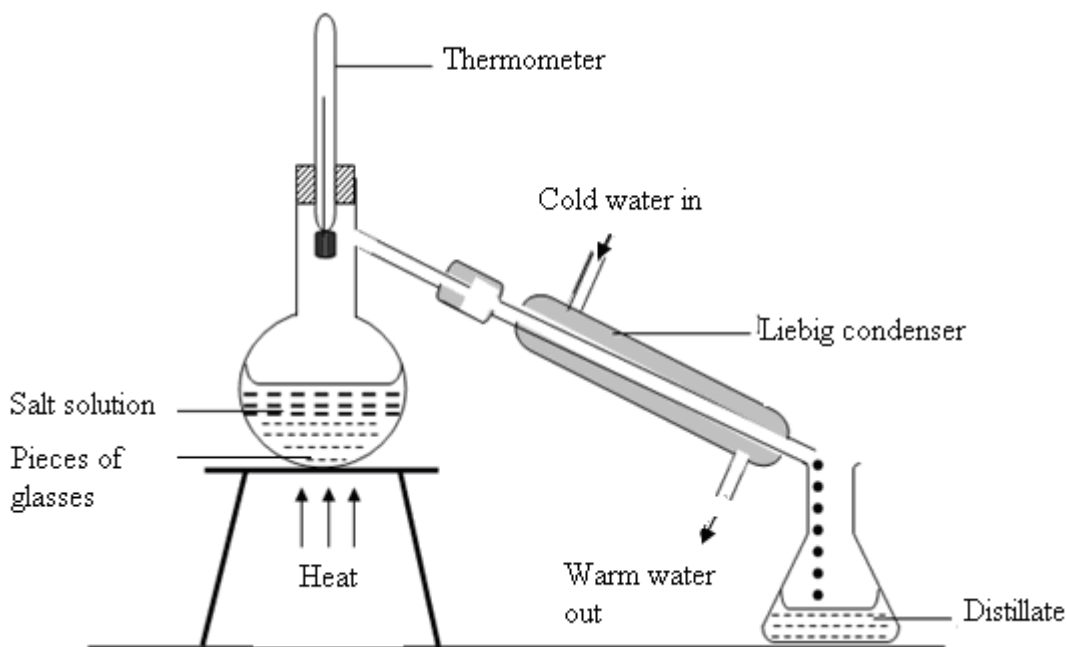
12. A mixture of phenolphthalein and methyl orange was separated as shown in the diagram in figure 1.



- (a) Name the method used in the separation of the mixture.
- (b) (i) State the colour of methyl orange in this experiment.
(ii) Phenolphthalein is usually colourless. Explain why it is purple in this experiment.
- (c) Give one other mixture that can be separated by the method you have named in (a).
13. The following chromatographs were obtained when a food-colour dye was analyzed using two different solvents:



- (i) Which pure dyes are present in the food-colour? Explain your answer.
- (ii) How far would pure green dye go in a chromatography with solvent two?
- (iii) Suggest any precautions that are necessary. What contributes to the separation of dyes?
- (iv) State one industrial use of chromatography.
14. Pure water can be obtained from sea water by using the apparatus shown below.



- (a) The condensation of vapours is not satisfactory. Why do you think it is so?
- (b) Why are there a few pieces of broken porcelain placed in sea water in the flask?
- (c) Why is the thermometer placed in the position shown in the figure?
- (d) Suggest one reason for collecting the distillate in a conical flask rather than a beaker.
- (e) Why is the distilled water collected in the conical flask unlikely to contain any dissolved salts? How can you show that this is true?
- (f) Suggest two reasons why this apparatus would be unsuitable for obtaining pure ethanol from ethanol- water mixture. Which would be the first distillate in this case and why?
- (g) Name **two** industrial processes where the improved method as suggested in (f) is applied in separating substances.

15. A pure sample of iodine crystals can be separated from its mixture with sand by heating.

- (a) What would be observed during the heating?
- (b) What property of iodine makes this separation possible?
- (c) Give **one** other mixture which can be separated by the above method.

CHAPTER FOUR

CHEMICAL SYMBOLS AND FORMULAE OF ELEMENTS AND COMPOUNDS

Learning objectives

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

1. (a) Define: (i) Metal and (ii) Non-metal.
(b) Differentiate between a metal and a non-metal.
 2. (a) Name and list the Chemical Symbols of the common elements.
(b) Write the Chemical Formulae of the common elements
 3. (a) Define the following terms:
(i) Atomicity
(ii) Valency
(iii) Radicals
(b) Write the Chemical Formulae of Compounds.
(c) Identify the valences of elements in a compound.
(d) Write word equations for the most common reactions.
-

4.0 Elements

An element is a substance that cannot be split into two or more simpler substances by any chemical means.

There are 116 elements occurring, 92 of which are natural and 24 are artificially made by Chemists.

4.10 Classification of Elements

Elements are classified into two classes namely:

- (i) Metals
- (ii) Non-metals

(a) Metals

A metal is an element which ionizes by loss of electron(s).

Examples of metals are: Iron, copper, silver, zinc, aluminium etc.

(b) Non-metals

A non-metal is an element which ionizes by gain of electron(s).

Examples of non-metals include: Carbon, sulphur, oxygen, nitrogen, hydrogen etc.

Metals	Non-metals
1. Ionize by loss of electron(s).	- Ionize by gain of electron(s).
2. Good conductors of heat and electricity.	- Bad conductors of heat and electricity.
3. Malleable (can be rolled into sheets).	- Not malleable (can not be rolled into sheets).
4. Ductile (can be drawn into wire).	- Not ductile (can not be drawn into wire).
5. Lustrous (shiny when polished).	- Not lustrous (can not shine when polished).
6. Sonorous (produce sound when hit).	- Not sonorous (do not produce sound when hit).
7. Generally have high tensile strength.	- Generally have low tensile strength.
8. Generally have high melting point.	- Generally have low melting point.
9. Generally have high densities.	- Generally have low densities.
10. Generally exist in solid form at ordinary temperature except mercury.	- Generally exist in gaseous state except a few exist in solid form. E.g. carbon
11. Have valences 1, 2, 3 and 4.	- Have valences 1 – 7.

Table 4.1 Differences between Metals and Non-metals

4.20 Chemical Symbols

A chemical symbol is a letter or two letters which represent one atom of an element.

Most symbols are derived from the chemical names of the elements. However, a few symbols are derived from the Latin names of the elements or the names of the Chemists who discovered them.

The symbols of the first twenty elements in the Periodic Table are shown in table 4.2 below.

Chemical Name	Latin Name	Symbol
1. Hydrogen		H
2. Helium		He
3. Lithium		Li
4. Beryllium		Be
5. Boron		B
6. Carbon		C

7. Nitrogen		N
8. Oxygen		O
9. Flourine		F
10. Neon		Ne
11. Sodium	Natrium	Na
12. Magnesium		Mg
13. Aluminium		Al
14. Silicon		Si
15. Phosphorous		P
16. Sulphur		S
17. Chlorine		Cl
18. Argon		Ar
19. Potassium	Kalium	K
20. Calcium		Ca

Table 4.2

Other common elements include:

Chemical Name	Latin Name	Symbol
1. Barium		Ba
2. Bromine		Br
3. Cobalt		Co
4. Copper	Cuprum	Cu
5. Iodine		I
6. Iron	Ferum	Fe
7. Lead	Plumbum	Pb
8. Manganese		Mn
9. Mercury	Hydrogerum	Hg
10. Nickel		Ni
11. Platinum		Pt
12. Silver	Argentum	Ag
13. Tin	Suntium	Sn
14. Zinc		Zn

Table 4.3

4.30 Chemical Formulae of Elements

The formula of an element is the symbol and a number which mean one molecule of the element.

(a) Representation of a molecule of an element

A molecule of element is denoted by writing the symbol of the element and to the right slightly below it a number expressing the number of the atoms in the molecule.

Generally a molecule of an element is represented as:



Where : X = Symbol of an element
 n = Number of atoms present in the molecule.

The value of n may be equal to: 2, 3, 4,

Examples of molecules of elements are:

O ₂	- Oxygen,
H ₂	- Hydrogen,
N ₂	- Nitrogen,
Cl ₂	- Chlorine.

The meaning of the above molecules

H₂ - Denotes one molecule of hydrogen containing two atoms of hydrogen chemically combined.
 O₂ - Denotes one molecule of oxygen containing two atoms of oxygen chemically combined.
 O₃ - Denotes one molecule of oxygen containing three atoms of oxygen chemically combined.

(b) Atomicity of Elements

Atomicity of an element is the number of atoms in one of its molecules.

The table 4..3 below shows the atomicity of some common elements.

Element(s)	Chemical Symbol/ formula	Atomicity	Prefix
Noble gases	-He, Ne, Ar, Kr, Xe	1	Monatomic
Hydrogen, Nitrogen, Oxygen	H ₂ N ₂ O ₂	2	Diatomic
Ozone	O ₃	3	Triatomic
Phosphorous vapour	P ₄	4	Tetra atomic
Sulphur	S ₈	8	Octatomic

Table 4.4

(b) Representation of two or more molecules of an element

The number of molecules present is shown by writing a whole number in front of the formula of the molecule.

Many molecules are represented as: mX_n

Where: X_n = symbol of the molecule of an element.

m = a numerical value (2, 3, 4) showing the number of molecules present.

The total number of the atoms of the element present is given by: mn

Note: m and n are as defined above.

E.g. $3H_2$ Denotes three molecules of hydrogen physically combined together.

That is each molecule exists on its own.

The total number of hydrogen atoms in the molecule

$$= 3 \times 2$$

$$= 6$$

4.4 Valency (Combining Number) of Elements

Valency is a number representing the capacity of a single atom or *radical* to combine with other atoms or radicals. The value is an expression of the number of electron(s) that an atom can give to or accept from another atom or radical during a chemical reaction. Hydrogen is the standard. Its valency is 1.

Definition:

Valence is the number of hydrogen atom(s) which combine with or displace one atom of an element or a group of radical.

The tables 4.4 to 4.7 show valence of the common elements and radicals.

(a) Valences of Metals

Valency 1		Valency 2		Valency 3	
Element	Symbol	Element	Symbol	Element	Symbol
Potassium	K	Calcium	Ca	Aluminium	Al
Ammonium	*NH ₄ ⁺	Magnesium	Mg	Iron (III)	Fe
Sodium	Na	Zinc	Zn		
Silver	Ag	Iron (II)	Fe		
		Lead (II)	Pb		
		Copper (II)	Cu		
		Barium	Ba		
		Mercury	Hg		

Table 4.5

- NB:**
1. *NH₄⁺ Ammonium ion is not a metal but behaves like metals.
 2. Use the words “**PASS**” and “**ALIR**” to recall that all metals have valency of 2 except Potassium, Ammonium, sodium and silver having valency of 1 and Aluminium and iron (III) having valency of 3.

(b) Valency of Non-metals

Valency 1		Valency 2		Valency 3	
Element	Symbol	Element	Symbol	Element	Symbol
Hydrogen	H	Oxygen	O	Nitrogen	N
Chlorine	Cl	Sulphur	S	Phosphorous	P
Bromine	Br				
Iodine	I				

Table 4.6

4.41 Radicals

A radical is a group of two or more atoms that exist in several compounds but do not exist on its own.

The common radicals are shown in table 4.6 below.

Valency of radicals

Valency 1		Valency 2		Valency 3	
Radical	Formula	Radical	Formula	Radical	Formula
Hydroxide	OH	Sulphite	SO ₃	Phosphate	PO ₄
Nitrite	NO ₂	Sulphate	SO ₄		
Nitrate	NO ₃	Carbonate	CO ₃		
Hydrogen Carbonate	HCO ₃	*Oxide	O		
Hydrogen Sulphate	HSO ₄				
Permanganate	MnO ₄				
Chlorate	ClO ₃				
*Chloride	Cl				

Table 4.7

NB: * Chloride (Cl) and Oxide (O) are examples of “Special radicals” in which each consists of one atom.

4.42 Variable Valences

Some elements have variable valences. They are shown in table 4.8 below.

Class of Element	Element	Symbol	Valences
Metals	Copper	Cu	1 or 2
	Iron	Fe	2 or 3
	Mercury	Hg	1 or 2
Non-metals	Nitrogen	N	3 or 5
	Phosphorous	P	3 or 5

Table 4.8

4.5 Writing Chemical Formulae of Compounds

We use **chemical** symbols of elements, formulae of radicals and their valences to write the chemical formulae of compounds.

Formation of Compounds

Compounds are formed when two or more elements combine together chemically.

They may be formed between;

- | | | | |
|-------|---------------------------|------|--|
| (i) | A metal and non-metal | e.g. | NaCl (Sodium Chloride) |
| (ii) | A metal and a radical | e.g. | CaSO ₄ (Calcium Sulphate) |
| (iii) | A non-metal and non-metal | e.g. | HCl (Hydrogen Chloride) |
| (iv) | A non-metal and radical | e.g. | H ₂ SO ₄ (Sulphuric acid) |

(a) Rules for writing Chemical formulae of Compounds

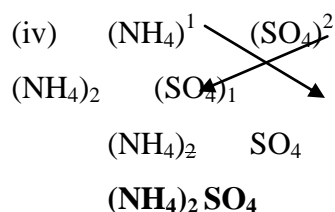
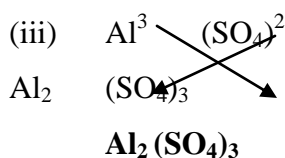
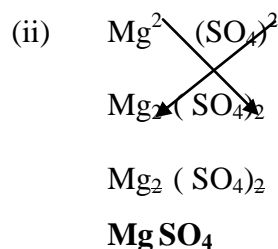
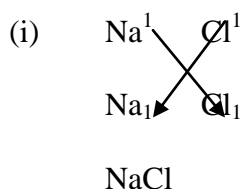
1. Write the symbols of the elements or the formula of the radical in the compound.
2. Write the valences above and to the right of the symbols or formula of radical in a bracket.
3. Re-write the symbols reversing their valences and writing figures 2, 3, 4, 5, 6 or 7 but not 1 below and to the right.
4. If the numbers in rule 3 are divisible, reduce them to their simplest form. If their simplest form is figure 1, remove the brackets from the radical.

Examples

Write the chemical formulae of the following

- (i) Sodium chloride.
- (ii) Magnesium sulphate.
- (iii) Aluminium sulphate.
- (iv) Ammonium sulphate.

Solution



(b) General Idea

Consider the compounds



In compound, AB_x , x = the number of atoms of element B and
= the valence of element A.

In compound, P_xQ_y , x = the number of atoms of element P and
= the valence of element Q.

y = the number of atoms of element Q and
= the valence of element P.

(c) The meaning of a formula of a compound

Since molecules are made up of atoms in simple whole-number ratios, molecular composition can be expressed by means of symbol of element, with subscript numbers used to indicate the relative number of atoms of each element in the compound.

Consider the compounds given below.

NaCl Means one molecule of sodium chloride containing one atom of sodium and one atom of chloride that are chemically combined together.

MgSO₄ Means one molecule of magnesium sulphate containing one group of sulphate attached chemically to one atom of magnesium.

Al(NO₃)₃ Means one molecule of Aluminium nitrate containing three groups of nitrate attached chemically to one atom of aluminium.

Ca₃(PO₄)₂ Means one molecule of calcium phosphate containing two groups of phosphate attached chemically combined to three atoms of calcium.

4.6 Nomenclature of Compounds

The word “*nomenclature*” means system of naming. The system enables us to state the composition of most chemical substances directly from the names of the elements in it.

(a) Binary Compounds (Compounds Containing Two Elements)

To binary compounds, the name ending – *ide* is assigned. The name of the compound is derived from the names of elements present in it. The name of the first element is written

fully and the name of the second element is ended using the suffix – *ide*. Examples of binary compounds that belong to this group include: -

- H_2S – Hydrogen sulphide
- NaCl – Sodium chloride
- CaCl_2 – Calcium chloride
- CuO – Copper (II) oxide

The chief exceptions to this rule include:

- water – H_2O
- ammonia – NH_3

Other exceptions are acid salts of hydrogen sulphide and ammonium chloride where there are three elements and the name ending is *ide*.

- NaHS - sodium hydrogen sulphide
- NH_4Cl - ammonium chloride

⇒ **Binary Compounds with oxide as the second element**

If oxide is used as the second element, the number of oxide atom(s) present is indicated by use of the following prefixes:

Number of oxides	Prefix	Example
1	Mono	CO - carbon monoxide
2	Di	CO_2 - carbon dioxide SO_2 - Sulphur dioxide NO_2 - Nitrogen dioxide
3	Tri	SO_3 - sulphur trioxide
5	Penta	P_2O_5 - Phosphorous pent oxide

Table 4.9

(b) **Acids**

Many acids contain hydrogen, Oxygen, and the third element. The most stable acids usually contain higher number of oxide (i.e. contain more oxygen atoms).

To this group is given the name ending – *ic*, and the name is derived from the name of the element it contains in addition to hydrogen and oxygen. e.g.

- H_2SO_4 - Sulphuric acid
- HNO_3 - Nitric acid
- H_3PO_4 - Phosphoric acid.

An acid containing the same element but less oxygen has the name ending – *ous*. Examples of acids that belong to this group are:

- H_2SO_3 - sulphurous acid.
- HNO_2 - nitrous acid.
- H_3PO_3 - phosphorous acid.

(c) Acid Radicals

The naming of the acid radicals depends on the acid from which they are derived. The acid radicals derived from – *ic* acids take the ending – *ate*. The acid radicals that belong to *ate* group are:

- SO_4 - sulphate from *sulphuric acid*.
- NO_3 - nitrate from *nitric acid*
- PO_4 - phosphate from *phosphoric acid*.

While those acid radicals derived from – *ous* acids take name ending – *ite*. *E.g.*

- NO_2 - nitrite from *nitrous acid*,
- SO_3 - sulphite from *sulphurous acid*.

(d) Metal Compounds (e.g. Metallic Salts)

(i) Metallic Compounds of metals with fixed valences

For metals which have fixed valency, the name of the metal is followed by the name of the acid radical.

- For example:
- NaNO_3 - sodium nitrate.
 - $\text{Al}_2(\text{SO}_4)_3$ - aluminium sulphate.
 - $\text{Mg}(\text{NO}_3)_2$ - magnesium nitrate.

(ii) Metallic Compounds of metals with variable valences

For metals with variable valences, roman numerals are included in the name to indicate the valency of the metal. For example:

- $\text{Cu}(\text{NO}_3)_2$ - Copper (II) nitrate.
- FeSO_4 - Iron (II) sulphate.
- $\text{Fe}_2(\text{SO}_4)_3$ - Iron (III) sulphate.

4.7 Chemical Equations

A chemical equation is an equation which represents a chemical reaction by means of symbols and formulae.

It tells us the *reactants* (substances that react), *products* (substances that are formed), their physical state and their relative amounts.

NB: *Chemical Reaction is a process by which atoms or groups of atoms are redistributed, resulting in a change in the molecular composition of substances.*

(a) Setting up a Chemical Equation

In setting up an equation, the following must be stated;

- (i) The physical nature of reactants and products.
- (ii) The direction of the reaction. This is done by use of;
 - Double barbed arrow, (\rightarrow) and
 - half barbed arrows in opposite directions (\rightleftharpoons).

A double barbed arrow is used if the reaction is one way reaction. (i.e. goes to completion). And half barbed arrows in opposite directions is used if the reaction is two way reaction (i.e. Reversible reaction).

(b) Steps Followed In Setting up a Chemical Equation

- (i) Write the formula (e)/symbol(s) of the reactant(s) to the left and product(s) to the right of an arrow. If there are more reactants and products, use (+) sign between them.
- (ii) Write to the right and slightly below the reactant(s) and product(s). The state symbols in brackets.
 - (s) - Solid state.
 - (l) - Liquid state.
 - (g) - Gaseous state.
 - (aq) - Aqueous state.

Consider the equation below;



The equation is read as;

- *P* reacts with *Q* forming *R* and *S*.
- *P* and *Q* are called **reactants** while *R* and *S* are called **products**.

- ❖ The plus sign (+) to the left of the arrow means “*reacts with*” and the plus sign (+) to the right of the arrow simply means “*and*”.
- ❖ The arrow shows that the reaction proceeds from left to right.
- ❖ The symbols in the brackets show the *physical states* of the reactants and the products.

4.8 Word Equation

A word equation is an equation which represents a chemical reaction by means of words with out the use of chemical symbols.

Example:

1. Metal + Oxygen → Metallic oxide
 Copper + Oxygen → Copper (II)oxide
 Cu (s) + O₂ (g) → CuO (s)

2. Non-metal + Oxygen → Non-metallic oxide
 Carbon + Oxygen → Carbon dioxide
 C (s) + O₂ (g) → CO₂ (s)

3. Base + Acid → Salt + Water
 Sodium + Hydrochloric → Sodium
 hydroxide acid Chloride + Water
 NaOH (aq) + HCl (aq) → NaCl (aq) + H₂O (l)

4. Metal + Acid → Salt + Hydrogen
 Magnesium + Sulphuric acid → Magnesium sulphate + Hydrogen
 Mg (s) + H₂SO₄ (aq) → MgSO₄ (aq) + H₂ (g)

5. Acid + Metallic carbonate → Salt + Carbon dioxide + water
 H₂SO₄ (aq) + MgCO₃ (s) → Mg SO₄ (aq) + CO₂ (g) + H₂O (l)

Self-Check 4

SECTION A

1. The valency of X in $X_2(SO_4)_3$ is
A. 2
B. 3
C. 4
D. 5
2. The valency of M in MSO_4 is
A. 2
B. 3
C. 4
D. 5
3. Element X reacts with chloride to form a compound with formula XCl_4 . The formula of the oxide of X is
A. XO_4
B. XO
C. X_2O
D. XO_2
4. The formula of the nitrate of metal P is PNO_3 . The formula of the oxide of P is
A. P_2O_3 .
B. P_2O_5 .
C. P_2O .
D. PO_2
5. The formula of the chloride of metal M is MCl_3 . The formula of the sulphate of M is
A. MSO_4 .
B. M_2SO_4 .
C. $M_2(SO_4)_3$.
D. $M_3(SO_4)_2$.

SECTION B

6. Define the following terms
 - (a)
 - (i) Element.
 - (ii) Metal.
 - (iii) Non-metal.
 - (b) Differentiate between metal and non-metal.
7. (a)
 - (i) Atomicity of an element.
 - (ii) Radical.
 - (iii) Valency.

(b) Complete the table below.

Name of Element or radical	Chemical Symbol/ formula	Valency
	Fe	
Sodium		
Sulphate		
	CO_3	
	OH	

8. State the valencies of elements W, X, Y and Z in the following compounds.
- (a) (i) $W_2(SO_4)_3$. (ii) XCl_2 (iii) YCl (iv) ZO
- (b) Write the full meaning of the following compounds.
- (i) $NaOH$ (ii) H_2SO_4 (iii) $MgSO_4$ (iv) $MgCO_3$
9. Write the chemical formulae of the following
- (i) Sodium sulphate.
- (ii) Magnesium chloride.
- (iii) Aluminium nitrate.
- (iv) Ammonium Sulphate.
10. Complete the following word equations
- (a) Metal + Oxygen \rightarrow
- (b) Non-metal + \rightarrow Non-metallic oxide
- (c) Base + Acid \rightarrow
- (d) Acid + Metallic carbonate \rightarrow

CHAPTER FIVE

ATMOSPHERE AND COMBUSTION

Learning objectives

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

1. (a) Define: the term atmosphere.
(b) State: (i) the importance (uses) of some of the gases in air to organisms.
(ii) the Uses of Noble Gases.
(c) Describe: (a) the separation of the components of air using the process called fractional distillation.
(b) An experiment:
(i) to find if any air is used up when a candle burns.
(ii) identify the products of a Burning Candle.
(ii) to determine the percentage of Oxygen in air:
– using rusting method and
– combustion of copper in a combustion tube.
 2. (a) Recall the major sources of Atmospheric Pollution.
(b) State: (i) the Harmful Effects of the Atmospheric Pollutants.
(ii) the measures taken to reduce atmospheric pollution.
 3. (a) Define rusting.
(b) Recall the chemical name and chemical formula of rust.
(c) State the conditions necessary for rusting to occur.
(d) Describe an experiment to investigate the conditions necessary for iron rusts to rust.
 4. State: (a) The disadvantages of rusting.
(b) The methods of preventing rusting.
-

5.0 The Atmosphere

On broader base, **atmosphere** refers to the air (mixture of gases) surrounding any celestial body that has a gravitational field strong enough to prevent the gases from escaping; especially the gaseous envelope of Earth.

But at the level of this book, atmosphere is defined as: *The air surrounding the earth.*

(a) Air

Air is a mixture of gases. It consists of the following proportions of gases:

- | | |
|---------|-------------------------------|
| - 78 % | - nitrogen (N ₂), |
| - 21 % | - oxygen (O ₂), |
| - 0.9 % | - argon (Ar), |

- 0.03 % - carbon dioxide (CO₂) and the remaining
- 0.07 % - is a mixture of: water, noble gases (neon, helium, krypton, xenon) and other trace components.

Also present in air in minute quantities in some areas such as industrial areas are pollutant gases. These are: hydrogen sulphide, sulphur dioxide and carbon monoxide.

- NB:**
- (i) *The percentages are by any volume of air.*
 - (ii) *The percentage composition of the various gases may vary from one place to another.*
 - (iii) *Noble gases are inert in nature i.e. they do not react at ordinary temperatures and pressure.*
 - (iv) *The above percentages are not to be memorized except for:*
 - *nitrogen,*
 - *oxygen and*
 - *Carbon dioxide.*
 - (v) *Nitrogen and Oxygen are the main components of air.*

(b) Importance (uses) of the gases in air to organisms

The most important attribute of air is its *life-sustaining property*. Human and animal life would not be possible without oxygen in the atmosphere.

Animals use oxygen during respiration to break glucose chemically in their body to carbon dioxide and energy. The energy released helps them to perform various activities.

Word equation:

Glucose + Oxygen → Carbon dioxide + Energy.

While on the other hand, green plants use carbon dioxide from the atmosphere to manufacture their own food during photosynthesis, releasing oxygen as a byproduct.

Carbon dioxide + Water $\xrightarrow{\text{Energy}}$ Carbohydrate + Oxygen

Note that:

- (i) *Respiration reduces the amount of oxygen in the atmosphere.*
- (ii) *Photosynthesis reduces the amount of carbon dioxide from the atmosphere and at the same time increases the amount of oxygen in the atmosphere.*

In addition to providing life-sustaining properties, the various atmospheric gases can be isolated from air (using fractional distillation of liquid air) and be used in industrial and scientific purposes or applications.

Question: *Differentiate between Respiration and Photosynthesis.*

(c) Uses of Noble Gases

- (i) Liquefied noble gases under pressure, particularly xenon, are used as solvents in infrared spectroscopy. They are useful for this because they are transparent to *infrared radiation* and therefore do not obscure the spectra of the dissolved substances.
- (ii) Neon lamps are used for: - art, advertising, and airplane beacons.

5.10 Separation of the Components of Air

The components (gases) in air are separated by a method called *fractional distillation* of liquid air. This process purifies and liquefies air at very cold temperatures. The liquid air is then boiled to isolate the gases (*a process called fractional distillation*). The method takes the advantage of the different boiling points of the various gases in air. The different components are called *fractions*. Each fraction boils off at different boiling point.

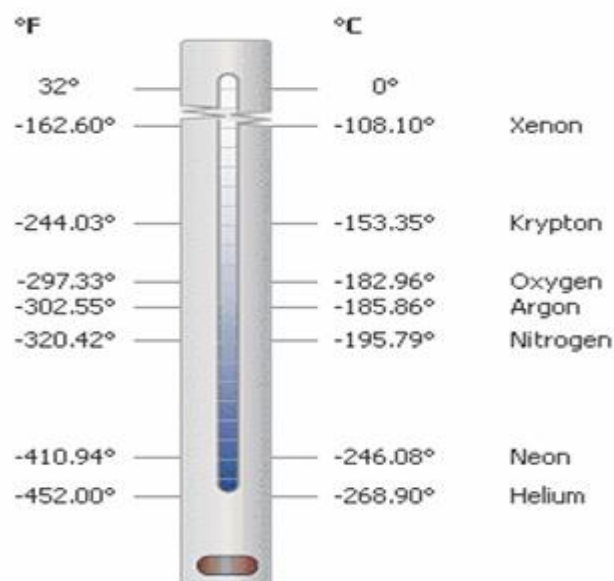


Figure 5

Liquid nitrogen boils at $-195.79\text{ }^{\circ}\text{C}$, argon at $-185.86\text{ }^{\circ}\text{C}$, and oxygen at $-182.96\text{ }^{\circ}\text{C}$. As the boiling temperature is increased, nitrogen vaporizes from the liquid air first, followed by argon, and then oxygen.

(a) Steps followed in fractional distillation of liquid air

- Air is first filtered to remove any dust particles. (If not removed, they act as impurities hence affecting the freezing/boiling points of the components of air)..
- The dust free air is successively compressed and cooled to -80°C . Water vapour and carbon dioxide freeze at their freezing points $0\text{ }^{\circ}\text{C}$ and -78°C respectively. *The removal is done to prevent them from blocking the pipes in the rest of the system.
- The remaining liquid air is then fractionally distilled.
- The different fractions boil off at their respective boiling points.

The Table 5 below shows the different boiling points of the gases.

Gas	Chemical formula/Symbol	Boiling point ($^{\circ}\text{C}$)
Helium	He	-269
Neon	Ne	-246
Nitrogen*	N_2	-196
Argon	Ar	-186
Oxygen*	O_2	-183
Krypton	Kr	-152
Xenon	Xe	-108

Table 5

NB: Considering the main gases Nitrogen and Oxygen in the table above, Nitrogen boils off before oxygen.

5.2 Combustion.

Combustion, is the process of rapid or slow oxidation or burning of a substance with simultaneous evolution of heat and, usually, light. It is the chemical combination of substances with atmospheric oxygen to produce *oxide* and *other products*. The term combustion, however, also embraces oxidation in the broad chemical sense, and the oxidizing agent may be different substances. Some substances burn when heated in air (i.e. give out flame) while others do not burn. Those which burn are said to undergo rapid combustion. For example burning of Sodium, Potassium, Magnesium and fuels (petrol, paraffin, diesel etc). While those which do not burn are said to undergo slow combustion e.g. rusting of iron.

Since during combustion, a substance reacts with the atmospheric oxygen, it can be used in an experiment to find if any air is used up during the process of combustion.

E.g. Burning of : - phosphorous or
 - candle.

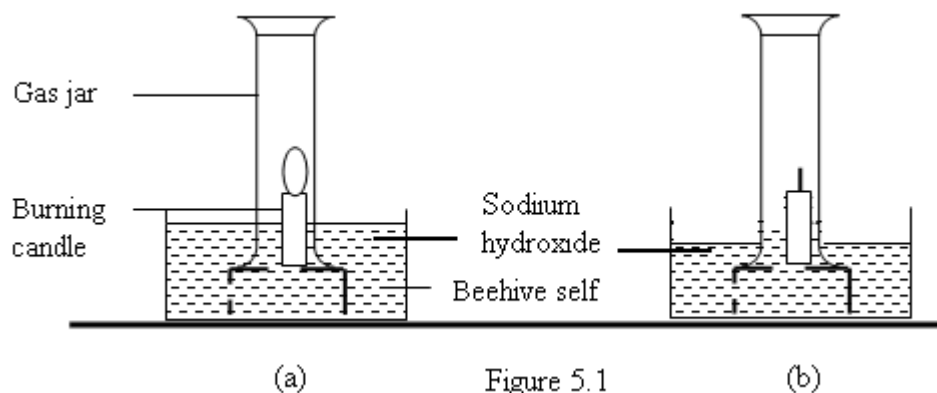
Experiment 5.1 To find if any air is used up when a candle burns.

Apparatus

A burning candle, Sodium hydroxide, a trough, a gas jar and a beehive shelf.

Procedure

- Stick a burning candle on a beehive shelf.
- Place the beehive shelf in a trough.
- Fill the trough with dilute sodium hydroxide solution to cover part of the candle.
- Invert a gas jar and quickly lower it over the burning candle.
- Allow it to stand on the beehive shelf as shown in figure 5.1 below.



Observation

- The candle continued to burn for a few seconds and the flame went off.
- The level of the sodium hydroxide solution in the trough lowered and that in the gas jar rose up.

Explanation

- In this experiment "*active air*" supports the burning of the candle.
- When the active air gets exhausted (i.e. used up), the flame goes out.
- The atmospheric pressure acting on the surface of the sodium hydroxide pushes the solution to fill the space left by the "*active air*".

Conclusion: Some air is used when candle burns in air.

NB: The role of the sodium hydroxide is to absorb carbon dioxide, which is one of the products of a burning candle. Carbon dioxide extinguishes burning fire, therefore, if not removed as soon as possible, it will extinguish the burning candle before all the active air is used up.

(a) Products of a Burning Candle

Experiment 5.2 To determine the products of a burning candle

Apparatus/Requirements

Burning candle, a U-tube, beaker, cold water, 2 delivery tubes, a test tube, a suction pump.

Procedure

- Set up the apparatus as shown in figure 5.3 below.
- Light the candle and at the same time, switch the pump on. (The pump helps to pull the gaseous products of the burning candle and keeps it burning)
- Leave the apparatus to stand for some time.

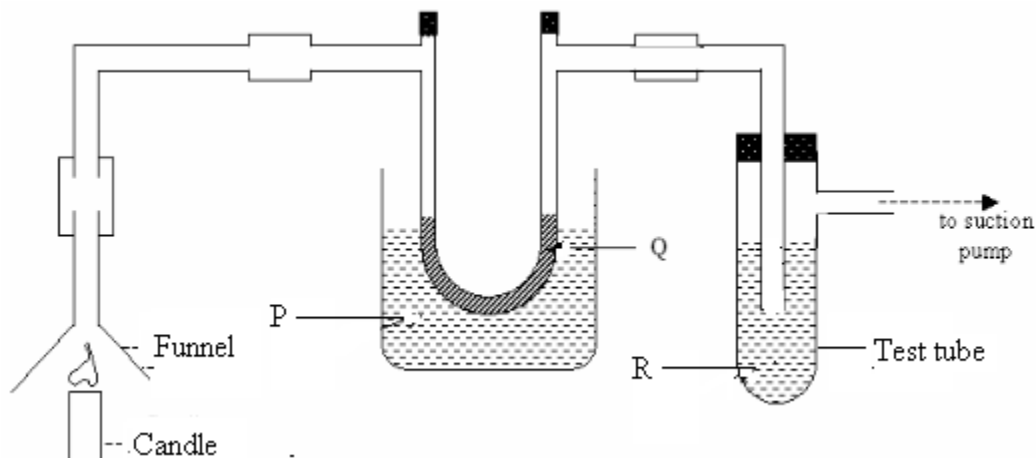


Figure 5.3 The set up of the apparatus for a burning candle.

P	=	Cold water	- Helps to condense water vapour.
Q	=	Liquid formed during condensation.	
R	=	Calcium hydroxide, $\text{Ca}(\text{OH})_2$, solution (lime water).	

Observations

- A colourless liquid collects in the U-tube.
- When the liquid is tested with anhydrous Copper (II) Sulphate, the white anhydrous copper (II) sulphate turned blue. This test confirms the presence of water.
- The colourless solution of limewater turned milky. This confirms the presence of carbon dioxide.

Conclusion:

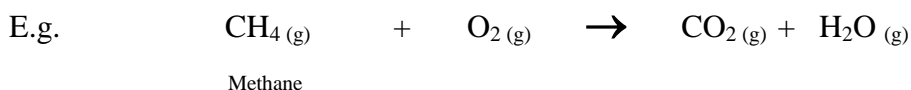
The above observations show that water (H₂O) and carbon dioxide (CO₂) are formed when a candle burns in air.

The results of the experiment also showed that the active part of air is oxygen, since the two products (H₂O & CO₂) contain oxygen atoms. Yet the candle wax is a *hydrocarbon* i.e. only contains carbon and hydrogen.

Hydrocarbon

A hydrocarbon is a compound that contains carbon and hydrogen only.

Therefore, Hydrocarbon + Oxygen → Carbon dioxide + Water



(b) Determination of the percentage of Oxygen in air

The percentage of oxygen in air can be determined by using:-

- (i) Copper metal or
- (ii) Rusting of iron.

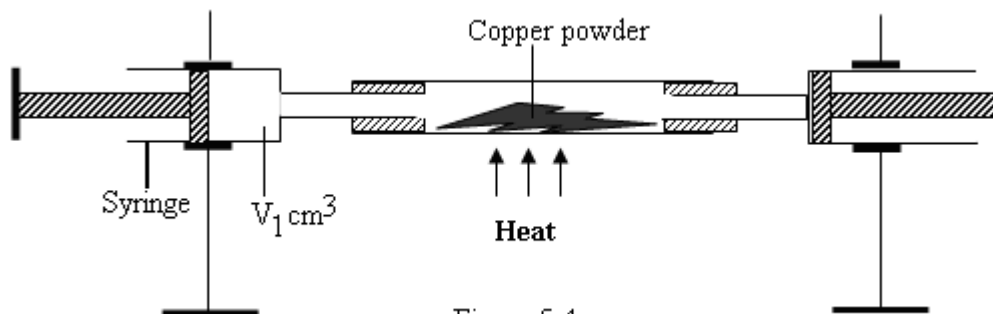
Experiment 5.3 To determine the percentage of oxygen in air using copper metal

Apparatus:

A hard Pyrex glass tube, copper turnings, 2 rubber stoppers, 2 syringes, weighing machine and source of heat

Procedure:

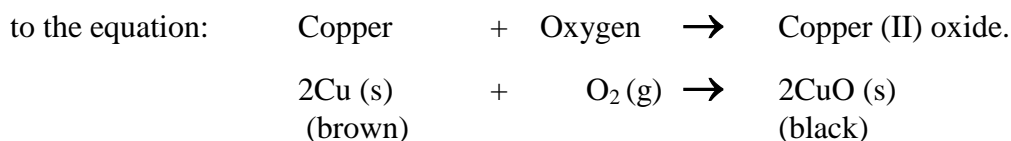
- Place copper turning in a combustion tube.
- Connect two syringes at the ends of the combustion tube. One tube containing a known volume, V₁, of air and another having zero volume of air as shown in figure 5.4 below.



- Heat the copper strongly until it is red hot.
- Pass air from the syringe into the tube forward and backward.
- After a few minutes, allow the apparatus to cool and then weigh.
- Repeat the procedure of heating, cooling and weighing until a constant mass is achieved.

Observations

- The piston of the syringe with zero volume of air moved backward to a constant volume, V_2 .
- The brown solid copper turned black, due to the formation of copper (II) oxide according



Calculation:

Vol. of oxygen used to oxidize copper = *Initial volume of air – Residual volume of air*

$$= V_1 - V_2$$

$$\begin{aligned} \text{Percentage of oxygen in air} &= \left(\frac{\text{Vol. of oxygen used}}{\text{Initial vol. of air}} \right) \times 100 \\ &= \left(\frac{V_1 - V_2}{V_1} \right) \times 100 \end{aligned}$$

If the experiment is accurately performed, the value obtained is approximately 21%.

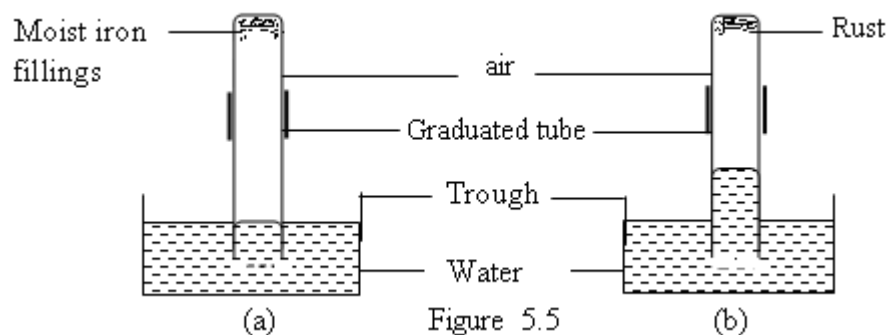
Experiment 5.4 To determine the percentage of oxygen in air using rusting method

Apparatus

Trough, measuring cylinder, water, retort stand, steel wool, rubber tubing

Procedure:

- Place a wet steel wool in the bottom of a measuring cylinder.
- Invert the measuring cylinder clamp it up and place it in the water in a trough.
- Adjust the water levels inside and outside the cylinder to the same level using a rubber tubing.
- Read and record the initial volume of air, V_1 , as shown in figure 5.5 (a) below.
- Leave the setup to stand for about a week and note the level of the water in the cylinder.
- Read and record the final volume of air, V_2 .



Observations:

- The water level in the tube rose to a constant level while the level in the trough dropped as shown in figure 5.5 (b) above.
- The steel wool turned from grey to brown.

Calculation:

$$\text{Vol. of oxygen used to in rusting} = \text{Initial volume of air} - \text{Residual volume of air}$$

$$= V_1 - V_2$$

$$\text{Percentage of oxygen in air} = \left(\frac{\text{Vol. of oxygen used}}{\text{Initial vol. of air}} \right) \times 100$$

$$= \left(\frac{V_1 - V_2}{V_1} \right) \times 100$$

If the experiment is accurately performed, the value obtained is approximately 21%.

Assumptions:

- In this experiment it is assumed that the steel wool is in excess to use up all the oxygen in the cylinder.
- The volume occupied by the steel wool is negligible.

Example

1. In an experiment to determine the percentage of active air using a wet steel wool placed in a measuring cylinder inverted in water, the initial volume of air was 240 cm^3 . After 7 days, the water level in the cylinder rose to a constant value, 192 cm^3 . Calculate,
- (a)
 - (i) The volume of air used.
 - (ii) The percentage of air used.
 - (b)
 - (i) Name the main gas left in the cylinder.
 - (ii) State one use of the gas you have named in b (i) above.

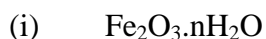
Solution:

(a) (i) Initial volume of air $= 240\text{ cm}^3$
Final volume of air $= 192\text{ cm}^3$
Volume of oxygen used in rusting $= \text{Initial vol. of air} - \text{Final vol. of air}$
 $= 240 - 192$
 $= \mathbf{48\text{ cm}^3}$

(ii) Percentage of oxygen in air $= \frac{\text{Volume of oxygen used}}{\text{Initial Volume of air}} \times 100$
 $= \frac{48}{240} \times 100$
 $= \mathbf{20\%}$

- (b) (i) Nitrogen.
(ii) It is an important plant nutrient.

2. A clean sample of steel wool was placed in a test tube containing some water and the test tube was inverted in a trough of water. After three days the volume of air in the test tube changed from 20 cm^3 to 16 cm^3 and a brown layer formed on the steel wool.
- (i) Write the formula of the brown solid formed.
 - (ii) Calculate the percentage decrease in the volume of air in the test tube.

Solution:

(ii) Initial volume of air = 20 cm^3

Final volume of air = 16 cm^3

Volume of oxygen used in rusting = Initial vol. of air - Final vol. of air

= $20 - 16$

= **4 cm^3**

(ii) Percentage of oxygen in air = $\frac{\text{Volume of oxygen used}}{\text{Initial Volume of air}} \times 100$

= $\frac{4}{20} \times 100$

= **20%**

5.3 Atmospheric Pollution

Air Pollution is the addition of harmful substances called pollutants into the atmosphere resulting in damage to the environment, human health, and quality of life.

5.31 Major Sources of Pollutants

Sources of most air pollutants come from the following activities:

- (i) Generation of electricity,
- (ii) Burning fossil fuels (natural gas, coal) and oil, to power industrial processes of manufacturing products and automobiles
- (iii) Forest fires that emit solid particles into the atmosphere.
- (iv) Volcanic eruptions
- (v) Smog, a mixture of smoke (particulates) and fog.

5.32 Types of Atmospheric Pollutants

There are two types/categories of atmospheric pollutants, namely:-

- (a) Gaseous pollutants (poisonous gases) and
- (b) Solid particles.

Both types of pollutants have adverse harmful effect on the environment, human health and quality of life as discussed below.

(a) Poisonous Gases

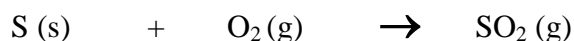
These arise from combustion of fuels, reaction of atmospheric nitrogen and oxygen, and the decomposition of insecticides, perfume sprays and other substances.

Examples of gaseous pollutants include:-

(i) Sulphur (IV) oxide and Hydrogen sulphide

They are formed when the element sulphur is burned in air and hydrogen.

Sulphur + Oxygen \rightarrow Sulphur (IV) oxide



Sulphur + Hydrogen \rightarrow Hydrogen sulphide



Harmful Effects of Sulphur dioxide and hydrogen sulphide

\Rightarrow Toxic effect

These gases (SO_2 and H_2S) are toxic and cause respiratory diseases which result into breathing problems and promotes cancer.

\Rightarrow Acidic effect

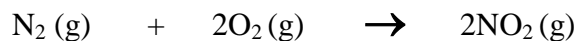
They also dissolve in water to form “acid rain”. The acid rain damages roofs, buildings, crops and other plants.

It also kills the aquatic life e.g. fish in rivers, lakes and other water bodies.

(ii) Nitrogen dioxide (NO_2)

- Nitrogen dioxide is formed during lightning from a reaction between atmospheric nitrogen and oxygen according to the equation:

Nitrogen + Oxygen \rightarrow Nitrogen dioxide



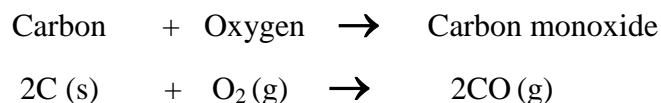
- It also comes from the internal combustion in engines where temperatures are very high.

Harmful Effects

- Like SO₂ and H₂S, the gas is toxic and promotes respiratory diseases.
- Since it is acidic oxide, it also forms “acid rain” which kills the aquatic life.

(iii) Carbon monoxide (CO)

Carbon monoxide is formed when fuels such as petrol, diesel mostly from *automobile* and industries/factories, charcoal, kerosene etc burn in insufficient supply of air according to the equation:



Harmful Effects

Carbon monoxide is a very poisonous gas. If inhaled, it combines with hemoglobin, hence cutting off the supply of oxygen to the body cells. Thus the normal body functions are affected. Too much of it kills.

(iv) Chlorine (Cl₂)

Chlorine is formed when insecticides and perfume sprays are decomposed in the atmosphere.

Harmful Effects

Chlorine destroys ozone layer in the upper part of the atmosphere. The ozone layer is important because it absorbs harmful radiations from the sun.

When this layer gets removed the harmful radiations reach the surface of the earth and cause skin cancer, eye problems and also damages crops and plants in general.

(b) Solid Particles

Large quantities of solid particles suspended in air are also pollutants.

Examples of solid particle pollutants include:

- Lead particles - produced when leaded petrol which contains lead burns in vehicle engines.
- Carbon soot - from incompletely burned fuels.

- Dust
 - from - Cement, Sugar and Flour factories
 - Quarrying and
 - Forest fires.
- Smog, mixture of smoke and fog.
- Ultra fine dust particles, dislodged by soil erosion.

Harmful Effects

- (i) Lead is a dangerous air pollutant. When it accumulates in the body it damages the brain.
- (ii) These particles can enter the lungs and are retained there, hence increase the risk of respiratory problems such as lung cancer and other diseases, when they enter the lungs and are retained in it.
- (iii) Smog, a mixture of smoke and fog, irritates the eyes, throat, and lungs and also damages plants.
- (iv) Ultra fine dust particles, dislodged by soil erosion when water and weather loosen layers of soil, increase airborne particulate levels.

5.33 Measures taken to reduce atmospheric pollution.

There are several measures put in place to reduce atmospheric pollution. Some of which include:

- (i) Promoting use of lead-free petrol in vehicles e.g. Shell petrol station out lets selling “Extra Unleaded Petrol”.
- (ii) Promoting the use of cars whose exhaust pipes are fitted with catalytic converter. The converter converts the harmful gases into harmless ones.
- (iii) Use of better and more efficient filter systems in factories and industries.
- (iv) Better processing of fuels to make them free from sulphur compounds.
- (v) Removal of old vehicles from the transport system.
- (vi) Equipping the fire brigade with efficient and sophisticated equipments, to fight fire out breaks.
- (vii) Planting trees and grass to reduce soil erosion.

5.4 Rusting

Rusting is a complex chemical reaction in which the iron combines with both oxygen and water to form hydrated iron (III) oxide (rust).

(a) Chemical name of rust

The chemical name of rust is *hydrated iron (III) oxide*.

Hydrated - means contains water molecules.

Rust is a reddish-brown solid which is relatively weak and brittle.

(b) Chemical formula of rust

The chemical formula of rust is: **$\text{Fe}_2\text{O}_3 \cdot n\text{H}_2\text{O}$**

NB: - The *n* in the formula of rust represents the number of molecules of water.
- The dot in the formula shows that the water molecules are weakly attached to the iron (III) oxide.

(c) Conditions necessary for rusting to occur

The conditions necessary for rusting to occur are:-

- Iron,
- Oxygen and
- Water (moisture).

Using experiments, the above conditions necessary for iron to rust can be investigated

Experiment 5.4 To investigate the conditions necessary for iron to rust

Apparatus/Requirements

3 test tubes, oil, water, anhydrous Calcium chloride, cotton wool, source of heat, clean nails

Procedure

- Label the test tubes 1, 2 and 3.
- Place clean nails in each test tube.
- In test tube 1, add ordinary water.
- In test tube 2, add freshly boiled water followed by a layer of cooking oil.
- In test tube 3, push cotton wool half-way and add crystals of calcium chloride to remove water vapour from the air.

- Set up the apparatus as shown in figure 5.6 below.
- Leave the set up to stand for two or three days and then observe the nails in each test tube.

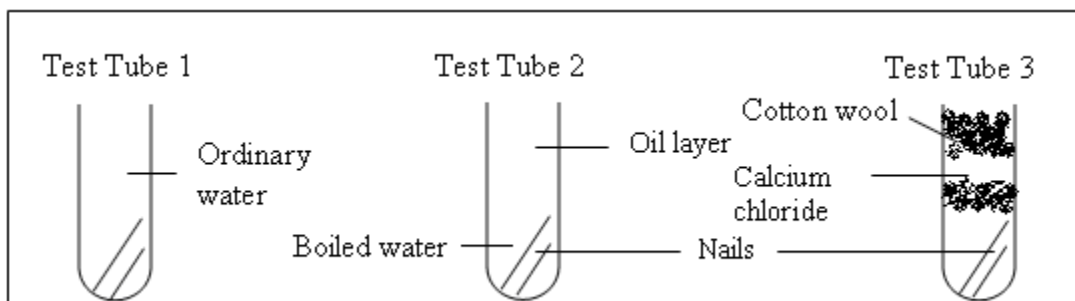


Figure 5.6

Observation

The iron nails in test tube 1 (where both moisture and air were present, were found covered with a brown solid substance while there was no observable change in the nails in test tubes 2 and 3 (i.e. the nails remained clean).

Conclusion

Oxygen and water are both necessary for the rusting of iron.
No rusting of iron will occur in dry air or in air-free water.

(d) Formation of rust.

During rusting, iron reacts with oxygen to form iron (III) oxide according to the equation:

Word Equation: Iron + Oxygen → Iron (III) oxide

Chemical equation: $4\text{Fe (s)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{Fe}_2\text{O}_3\text{(s)}$

Then the oxide formed reacts with water to form hydrated iron (III) oxide according to the equation:

Word Equation: Iron (III) oxide + Water → Hydrated Iron(III) oxide

Chemical equation: $\text{Fe}_2\text{O}_3\text{(s)} + n\text{H}_2\text{O(l)} \rightarrow \text{Fe}_2\text{O}_3 \cdot n\text{H}_2\text{O(s)}$
(rust)

(e) Effect of common salt on rusting

The common salt (Sodium Chloride) speeds up the rate of rusting.

NB: Carbon dioxide gas is also believed to speed up the process of rusting.

5.41 Disadvantages of Rusting

Rusting is a serious economic problem. Large sums of money are spent each year to replace damaged structures made of iron and steel. This is because rust forms flakes which peels off exposing more iron or steel to rusting. This weakens structures such as bridges, churches rafters and many others which eventually collapse. It is therefore, important to prevent materials made of iron from rusting.

(a) Methods of preventing rusting

The only way of preventing iron from rusting is by keeping water and air away from coming into contact with iron. This can be done by using four major methods. These include:

1. Covering it with an impermeable surface coating

This is done by:

- (i) *Painting* - Commonly used in ships, bridges, cars, fences and iron sheets.
- (ii) *Greasing* }
- (iii) *Oiling* } - are commonly used to protect moving parts in engines, pulleys
- (iv) *Tarring* - Coating of iron with tar (oily, dark-colored bituminous products). Tar is used to coat iron materials usually underground e.g. sewage pipes, water pipes, bottom parts of bridges, ships etc.
- (iv) *Plastic coating* - is largely used to protect equipments such as refrigerators, pliers e.t.c.

2. Coating it with an active metal

The most common examples of this method are:

(i) *Galvanizing*

This is coating of iron with zinc. The process involves dipping iron in molten zinc. In the presence of corrosive solutions, an electric potential is set up between the iron and the zinc, causing the zinc to dissolve. The Zinc reacts with air forming a thin coating of zinc oxide which protects iron from rusting.

(ii) *Tin plating*

This involves dipping iron into molten tin to form a layer of tin on top of the iron. Tin plating is used in food cans because it is un-reactive and non-toxic.

(iii) *Electroplating*

Electroplating is the electrochemical process of depositing a thin layer of metal on, another metal. Using this method, iron is electroplated to prevent it from rusting. Electroplating also makes iron articles shinny and attractive.

(v) *Sacrificial protection*

This is the type of protection where zinc or magnesium block are attached on to iron material at intervals. Since these metals are more reactive than iron, they corrode in preference to iron. It is done mostly in underground pipes and ships.

(vi) *Enameling*

In industry, enamel (known as vitrified or porcelain enamel) is applied to cast iron or sheet of steel that has previously been stamped into proper shape. Enamel is very hard and does not allow water or air to pass through it to the iron. It has no action on food. Examples of equipments that are coated with enamel are: Cookers, refrigerators etc.

3. Alloying the iron

For example calculated amounts of chromium or chromium and nickel are alloyed with the iron to form a stainless steel. Steel is chemically resistant to corrosion. It is used in cutlery and other types of kitchen ware.

4. Using drying agents

Silica gel - Some equipment made of iron are packed and supplied with small bags of silica gel. The silica gel absorbs any moisture in the package thus preventing rusting.

Self-Check 5

- The product formed when a candle burns in excess air is
A. carbon monoxide. B. methane.
C. water vapour. D. soot.
- The substance that can **not** cause air pollution from the following list is
A. hydrogen sulphide. B. carbon dioxide.
C. water vapour. D. Sulphur dioxide.
- Which one of the following processes adds sulphur dioxide into the atmosphere?
A. Burning coal and oil.
B. Decaying organic matter.
C. Heating limestone in a kiln.
D. Fractional distillation of liquid air.
- Which one of the following pairs of gases does **not** cause atmospheric pollution?
A. Sulphur dioxide and carbon dioxide.
B. Helium and hydrogen.
C. Nitrogen dioxide and sulphur trioxide.
D. Carbon monoxide and carbon dioxide
- Which one of the following is the chemical formula of rust?
A. $\text{FeO} \cdot n\text{H}_2\text{O}$ B. $\text{Fe}_2\text{O}_3 \cdot n\text{H}_2\text{O}$
C. $\text{Fe}_2\text{O}_3 \cdot n\text{H}_2\text{O}$ D. $\text{Fe}_3\text{O}_2 \cdot n\text{H}_2\text{O}$
- Which one of the following gases is obtained by fractional distillation?
A. Oxygen. B. Ammonia.
C. Sulphur dioxide. D. Hydrogen.
- Dry air was passed through sodium hydroxide solution and then over heated copper metal. The residual gas consists of
A. oxygen. B. carbon dioxide.
C. water vapour. D. nitrogen.
- The boiling points of some gases are given below.

Gas	Boiling point /°C.
Helium	-269
Nitrogen	-196
Oxygen	-183
Xenon	-108

When a liquid mixture of above gases is fractionally distilled, which one of them will vaporise first?

A. Xenon.

B. Helium.

C. Oxygen.

D. Nitrogen.

9. Which one of the following pairs of gases are removed first before air is liquefied?

A. Nitrogen and oxygen.

B. Water vapour and carbon dioxide.

C. Carbon dioxide and oxygen.

D. Water vapour and nitrogen.

10. Which one of the following gases in the atmosphere makes rain water acidic?

A. Sulphur dioxide.

B. Carbon monoxide.

C. Ammonia.

D. Nitrogen.

SECTION B

11. (a) Define the term *atmosphere*.

(b) State the approximate percentage of oxygen and nitrogen in air.

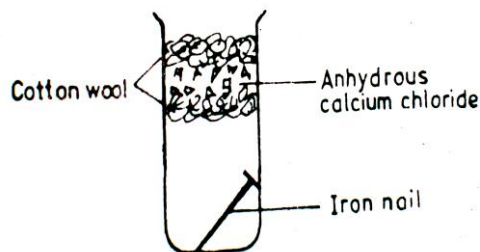
(c) Name the process by which oxygen is

(i) used up from the atmosphere.

(ii) replaced in the atmosphere.

12. (a) State the conditions necessary for rusting to occur.

(b) During an investigation to show the conditions under which an iron nail may rust, an experiment was set up as shown in figure 3 below:



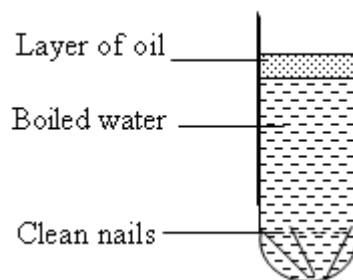
State the conditions which was being eliminated.

(c) State **one** disadvantage of rusting.

(d) (i) What is galvanised iron?

(ii) State **one** use of galvanised iron.

13. (a) An experiment was set up as shown in the diagram below to investigate conditions under which iron rusts.



- (i) What is the purpose of the layer of oil in the experiment?
- (ii) Why was boiled water used?
- (iii) State what was observed after some days.
- (iv) Give one method which can be used to prevent rusting.
- (b) With the aid of diagram describe an experiment you would carry out to show that rusting requires both oxygen and water in order to occur.
- (c) Describe **four** ways of preventing rusting.
14. (a) How would you show experimentally that air contains:
- (i) Carbon dioxide?
- (ii) Water vapour?
- (iii) Oxygen?
- (b) Describe an accurate experiment using either copper turnings or iron fillings through which you would determine the percentage by volume of oxygen in air.
- (c) State and explain two common uses of oxygen.
15. (a) (i) State two properties which show that air is a mixture.
- (ii) Name two gases, other than oxygen, that are constituent of air and give their approximate percentages in air.
- (iii) Write the name and formula of the product of the reaction between sulphur and oxygen.

CHAPTER SIX

PREPARATION, DRYING AND COLLECTION OF GASES IN THE LABORATORY

Learning objectives

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

1. (a) Name (i) *the methods of collection of gases.*
(b) State: (i) *the factors which determine the method of collections.*
(ii) *the appropriate drying agents for the common gases.*
 2. *Recall the precautions taken in drawing diagrams for preparation of gases in the laboratory.*
 3. *Define the following terms and in each case give examples:*
 - (i) *Hygroscopic substance*
 - (ii) *Deliquescent substance*
 - (iii) *Efflorescent substance*
-

6.10 Preparation of Gases

Generally, a gas is prepared in some form of a generator (reacting vessel).

This can be: - a round bottomed flask if heating is required.

- a flat bottomed flask or
- a conical flask, } - if no heating is required.

One reactant normally a liquid or aqueous is added to the other placed in the flask by using either: - a dropping funnel or

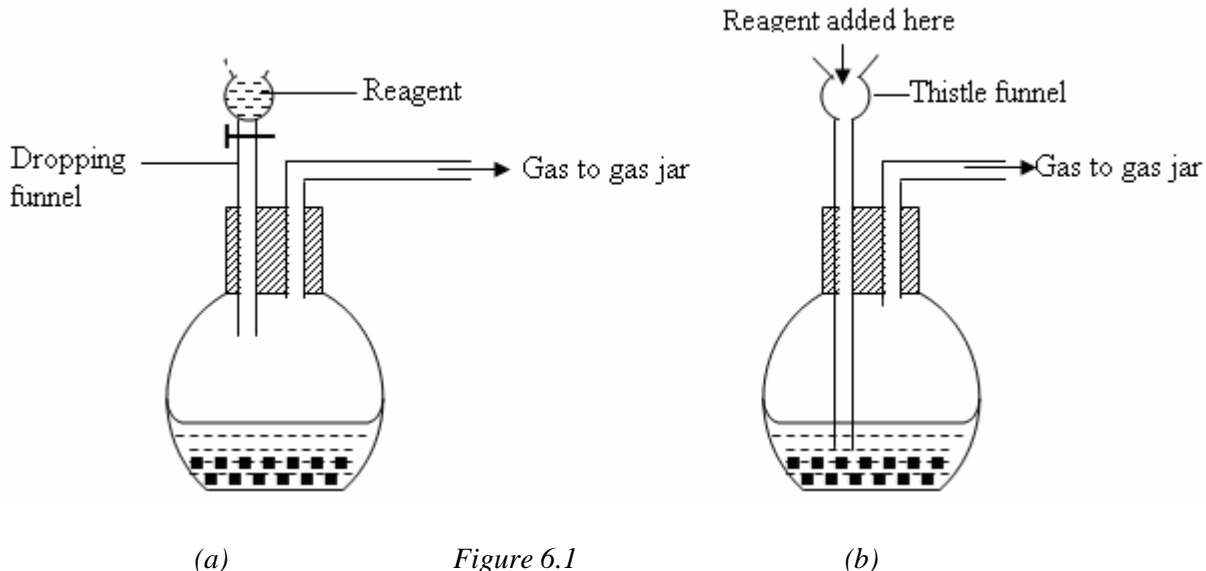
- a thistle funnel.

The gaseous product(s) leave(s) the generator through a delivery tube to the collecting gas jar via a drying agent, if the gas is required dry.

NB:

- ❖ *If a dropping funnel is used, to drop the reactant, the tap must be closed to prevent the gas from escaping the flask as shown in figure 6.1 (a).*
- ❖ *If a thistle funnel is used, its lower end must be dipped below the surface of the reactants. This prevents the gas from escaping from the flask.*
- ❖ *The delivery tube must always be well above the level of the reactants as shown in figure 6.1 (b).*

Diagrams showing the different Types of Generators



6.20 Drying of Gases

If a gas contains water vapour, it is not dry. If required dry, it is dried by passing it through a drying agent placed in a bottle. The delivery tube used to deliver the gas from the generator to the drying agent must be dipped below its surface. And the delivery tube which delivers the gas to the gas jar must be well above the drying agent as shown in the diagram below.

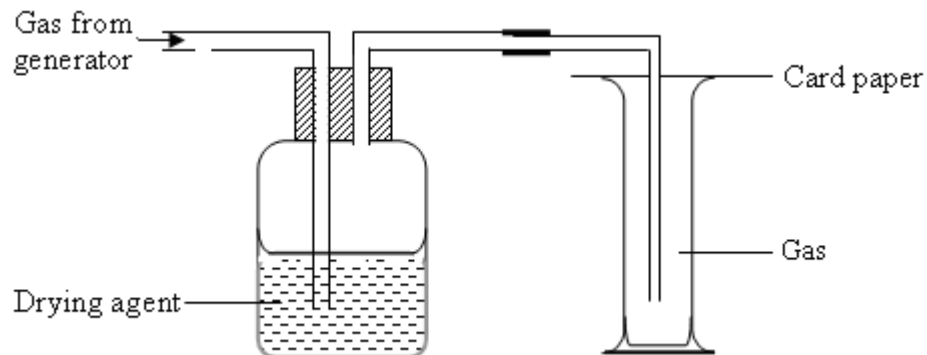


Figure 6.2

For solid drying agents like anhydrous calcium chloride (CaCl_2), a U-tube is used and the setup is as shown in the diagram below.

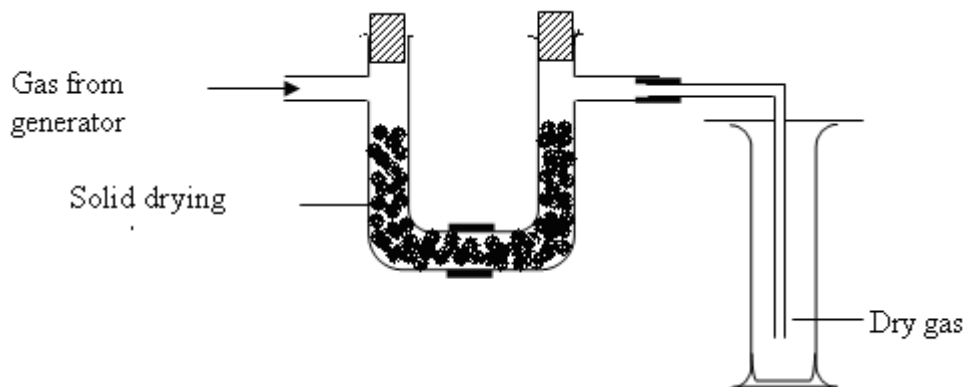


Figure 6.3

(a) Drying agents

A drying agent is a substance (compound) which has high affinity for water. Drying agents are used to remove water vapour from other compounds or elements (gases).

Examples of drying agents are:

- Hygroscopic and
- Deliquescent substances.

Definitions:

- (i) Hygroscopic substance: - *is a substance which absorbs water from the atmosphere.*
- (ii) Deliquescent substance: - *is a solid substance which absorbs water vapour from the atmosphere to form a solution.*

Deliquescence:- *is the process of absorbing water from the atmosphere by a solid to form a solution.*

Table 6.1 shows examples of hygroscopic, deliquescent and efflorescent substances.

Deliquescent substance		Hygroscopic substance	
Name of Compound	Chemical Formula	Name of Compound	Chemical Formula
Potassium hydroxide	KOH	Calcium oxide	CaO
Sodium hydroxide	NaOH	Anhydrous Copper (II) sulphate	CuSO ₄
Sodium nitrate	NaNO ₃	Copper (II) oxide	CuO
Calcium Chloride	CaCl ₂	Con. Sulphuric acid	H ₂ SO ₄
Iron (III) chloride	FeCl ₃	Ethanol	C ₂ H ₅ OH
Phosphorous (V) oxide	P ₂ O ₅	All deliquescent substances	

Table 6.1

The table 6.2 below shows the common drying agents used for drying appropriate gases.

Drying agent	Formula	Appropriate gas
Conc. Sulphuric acid	H ₂ SO ₄	All gases except ammonia
Phosphorous (V) oxide	P ₂ O ₅	
Silicon (IV) oxide (Silcagel)	SiO ₂	All gases
Calcium oxide	CaO	Ammonia
Anhydrous or Fused Calcium Chloride	CaCl ₂	Hydrogen sulphide

Table 6.2

NB:

- ❖ *A gas can be dried by bubbling/passing it slowly through a bottle containing concentrated sulphuric acid or through a U-tube containing a solid drying agent such as anhydrous calcium chloride or silica gel.*
- ❖ *The drying agent chosen should be one which does **not** react with the gas to be dried.*
- ❖ *Never dry a gas and collect over water.*

6.30 General notes on preparation of gases.

In preparing gases, the following points should be stated or shown clearly.

- ❖ Reactants.
- ❖ The conditions necessary.
- ❖ Drying agents used, if the gas is required dry.
- ❖ Method of gas collection.
- ❖ Diagram of the apparatus and
- ❖ The reaction equation.

- (a) **Reactants:** - State which reactant is placed in the flask and which one is added from a thistle or dropping funnel.

(b) **The conditions:**

Most reactants require optimum conditions for the reactions to take place. These conditions include:

- Heat/Warming.
- Concentration of the reactants.
- Physical state of the reactants. E.g. powder or granules, etc.

- Pressure (mostly for gaseous reactants).
- Catalyst
A catalyst is substance which increases the rate of reaction and remains unchanged at the end of the reaction.

(c) Drying agents:

The appropriate drying agent used should be stated if the gas is required dry.

(d) Methods of gas collection

The method of the gas collection must be shown clearly. The method by which a gas is collected depends on its *density* and *solubility in water*. There are majorly four methods of gas collection, namely:

- *Collection over water(Downward displacement of Water)*
- *Collection by downward delivery of gas (or upward displacement of air)*
- *Collection by upward delivery of gas (or downward displacement of air) and*
- *Collection in syringe.*

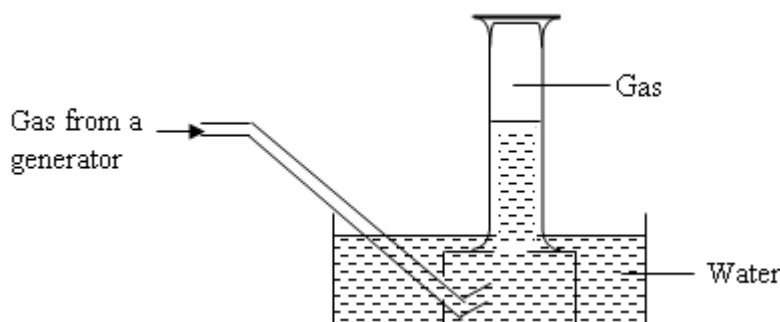
(i) Collection over water(Downward displacement of Water)

This method applies only to gases that are insoluble or slightly soluble in water. Examples of gases that can be collected over water are:

- Carbon monoxide, Carbon dioxide and Oxygen.

The method does not apply to highly soluble gases such as:

- Ammonia and
- Hydrogen chloride.



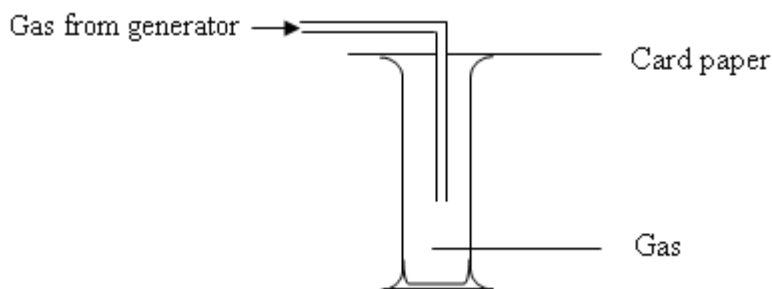
(ii) Collection by downward delivery of gas (upward displacement of air)

This method applies only to gases that are denser than air.

Examples of gases that are denser than air are:

- Carbon dioxide (CO_2),
- sulphur dioxide (SO_2),
- Chlorine (Cl_2) and
- Carbon monoxide (CO).

This method is called downward delivery because the gas is delivered downwards and the air which originally occupies the gas jar is displaced upwards.

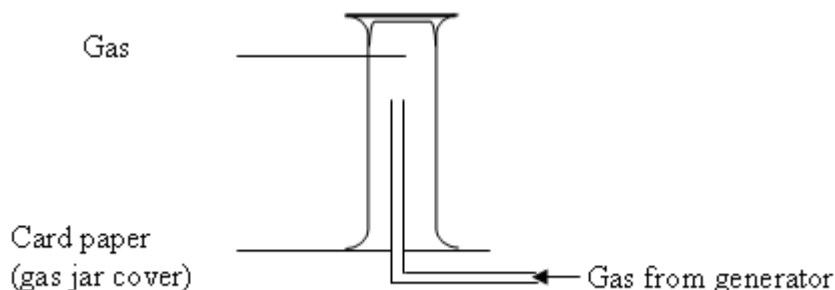


(ii) Collection by upward delivery of gas (downward displacement of air)

This method applies only to gases that are lighter than air.

E.g. Hydrogen (H_2) and
Ammonia (NH_3)

This method is preferred for the lighter gases because when the gas is delivered in to the upper part of the gas jar, it makes it much easier for it to displace the heavier air down wards by the help of pull of gravity.



(iii) Collection in syringe.

This method is applicable to any gas irrespective of its density and solubility. However, it is suitable for collecting gases of approximately the same density as air.



Self-Check 6.0

- Which one of the following gases can be collected over water?
A. Hydrogen chloride.
C. Ammonia.
B. Carbon monoxide.
D. Sulphur dioxide.
- The gas that **cannot** be dried using concentrated sulphuric acid is
A. sulphur dioxide.
C. ammonia.
B. hydrogen sulphide.
D. carbon monoxide.
- Which of the following gas/gases can be collected by upward delivery of the gas method?
(i) Hydrogen
(ii) Sulphur dioxide
(iii) Ammonia
(iv) Chlorine
A. (i), (ii) and (iii) only
C. (i) and (iii) only
B. (i) and (ii) only
D. (i) and (iv) only
- Which one of the following drying agents is suitable for drying ammonia gas?
A. Calcium oxide
C. Calcium Chloride
B. Conc. Sulphuric acid
D. Phosphorous (V) oxide
- The following are drying agents except:
A. Calcium oxide
C. Calcium Chloride
B. Conc. Sulphuric acid
D. Copper (II) sulphate

SECTION B

6.
 - (a) Define the following terms in each case give an example.
 - (i) Hygroscopic substance.
 - (ii) Deliquescent substance.
 - (iii) Deliquescence.
 - (b) List the common drying agents used in the laboratory to dry gases.
7.
 - (a) Name two gases that can be collected by the following methods.
 - (i) Upward delivery method.
 - (ii) Downward delivery method.
 - (iv) Over water.
 - (b) Give reasons why the gases you have chosen are collected by the above methods.

CHAPTER SEVEN

OXYGEN (O₂)

Learning objectives

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

1. (a) Describe the laboratory preparation of oxygen from:
 - (i) Hydrogen peroxide solution and
 - (ii) Potassium Chlorate.
 - (b) Name the catalyst in the preparation of oxygen.
 - (c) Describe manufacture (Industrial preparation) of oxygen.
 2. State the physical and chemical Properties of oxygen.
 3. List the elements according to their order in the Reactivity or activity series.
 4. Define and give examples of oxides
 5. State the uses of oxygen.
-

7.0 Introduction

Oxygen, chemical symbol O, and chemical formula O₂, is a gaseous element. It belongs to group six in the Periodic Table.

Occurrence

Oxygen occurs in nature both in free (or uncombined) and combined state. It occurs in the Free State in atmosphere occupying 21% by volume of air. In combined state it occurs in many compounds such as:

- Water,
- Acids (mineral and organic acids)
- Bases, and
- Most rocks and minerals.

It plays important role in the lives of living organisms. As such it is a constituent of all living tissues.

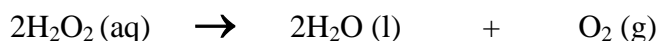
7.1 Laboratory Preparation of Oxygen

In the laboratory, oxygen is commonly prepared by using:

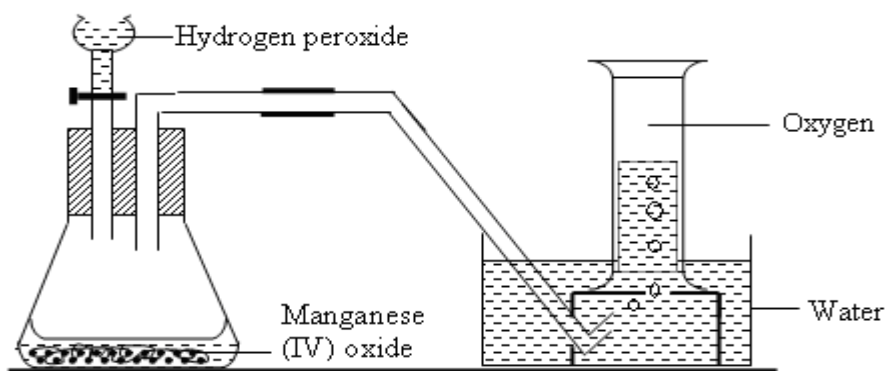
- (i) Hydrogen peroxide (H_2O_2) and
- (ii) Potassium chlorate (KClO_3).

(a) Preparation of Oxygen from Hydrogen peroxide solution:

Oxygen is prepared in the laboratory by the decomposition of hydrogen peroxide solution in the presence of a catalyst such as manganese (IV) oxide or copper (II) oxide. And the gas is collected over water.

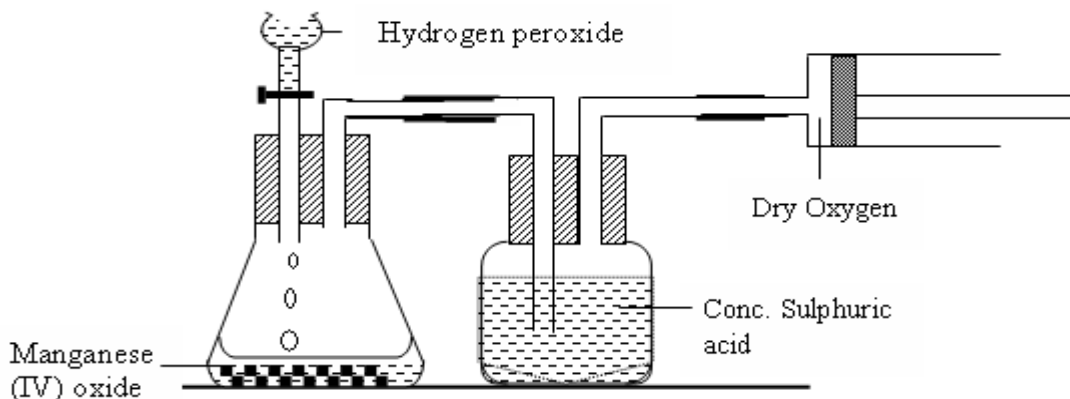


The Diagram Showing Laboratory Preparation of Oxygen



If the gas is required dry, it is passed through concentrated sulphuric acid and is then collected in syringe as its density is approximately the same as that of air below.

The Diagram Showing Laboratory Preparation of Dry Oxygen

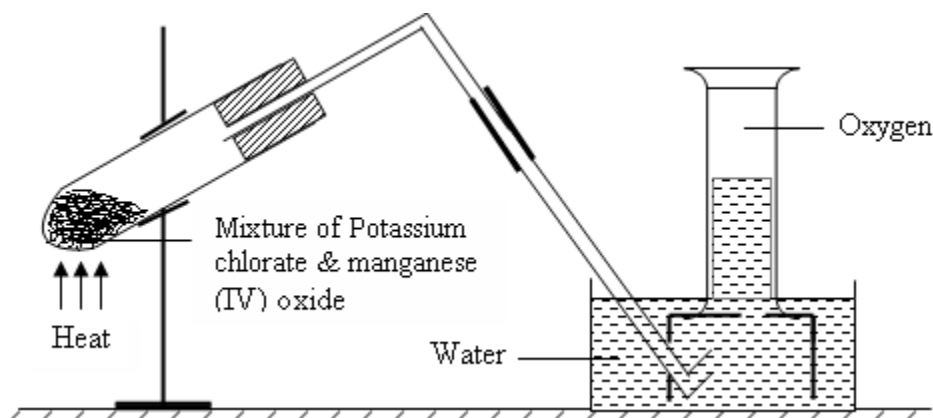


(b) Preparation of Oxygen from Potassium Chlorate:

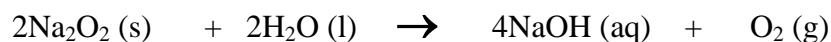
Oxygen is prepared in the laboratory by heating a mixture of potassium chlorate and manganese (IV) oxide (catalyst) in a boiling tube and the gas is collected over water.



The Diagram Showing Laboratory Preparation of Oxygen



NB: Oxygen can also be prepared by adding cold water to solid sodium peroxide according to the equation:



Chemical Test of oxygen

Oxygen is tested by using a glowing splint. Plunge a glowing splint into a gas jar of oxygen.

Observation: It rekindles a glowing splint.

7.2 Properties of Oxygen

(a) Physical Properties

- (i) It is a colourless gas.
- (ii) It is odourless gas.
- (iii) It is a tasteless gas.
- (iv) It is slightly soluble in water.
- (v) Its density is approximately the same as that of air.

(b) Chemical Properties

All elements (except noble gases and some un-reactive metals such as gold) combine with oxygen to form compounds called *oxides*. The reaction is called *Oxidation reaction*.

Oxidation is the addition of oxygen to a substance.

The readiness of elements to combine with oxygen is shown in a series called *reactivity or activity series*.

7.3 Reactivity or activity series

A reactivity series is an arrangement of elements beginning with the most reactive element and ending with the least reactive element.

Potassium	K	Most reactive
Sodium	Na	
Calcium	Ca	
Magnesium	Mg	
Aluminium	Al	
(Carbon)	(C)	
Zinc	Zn	
Iron	Fe	
Lead	Pb	
(Hydrogen)	(H)	
Copper	Cu	
Mercury	Hg	
Silver	Ag	
Tin	Sn	
Gold	Au	Least reactive

Increasing reactivity ↑

NB: Potassium and sodium are very reactive as such they are kept under oil or paraffin, thus preventing them from reacting with the oxygen in the atmosphere.

MEMORY AID

You may use the following statements to recall the reactivity series.

1. For metals only.

PoSo CaMAI ZILC MSG

2. The full list can be remembered by using the statement:

Popular Scientists Can Make A Clear Zoo In Low Humid Country-side More Satisfactory Than Geographers.

7.31 Reactions of Oxygen with elements

Oxygen supports the burning of:

- Metals and
- Non-metals.

(i) Reactions with Metals

Metals from potassium to copper (K, Na, Ca, Mg, Al, Zn, Fe, Pb, and Cu) react readily when heated in air to form *metallic oxides*. The reactivity decreases from potassium to copper.

Mercury, Silver and gold are not oxidized easily.

The general word equation is:



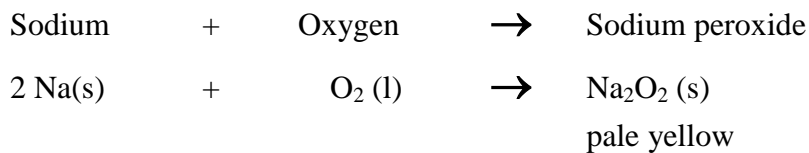
⇒ **Sodium:**

The combustion of sodium is divided into two, namely;

- Complete combustion
- Incomplete combustion

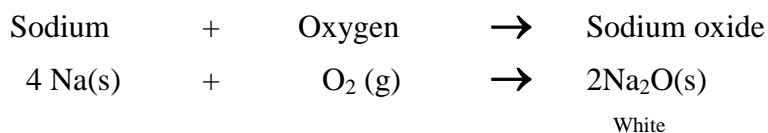
(i) Complete combustion

Sodium burns more brightly in oxygen than in air. It burns with a yellow flame forming a pale yellow solid called sodium peroxide.



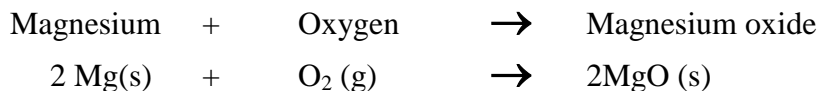
(ii) Incomplete combustion

In a limited supply of oxygen, sodium burns to form a white solid called sodium oxide.



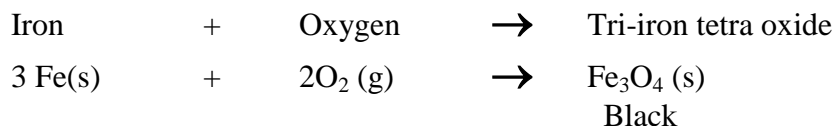
⇒ **Magnesium:**

Magnesium ribbon burns brightly in oxygen to form a white solid called magnesium oxide.



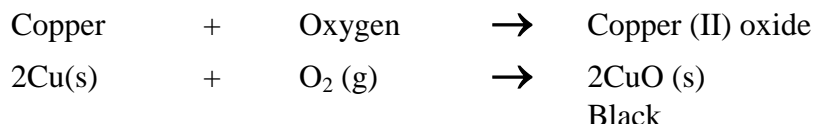
⇒ **Iron:**

When iron wool is strongly heated in air or oxygen, it burns with brick red flame forming a black residue called *tri-iron tetra oxide*.



⇒ **Copper**

When copper is heated in air or oxygen, it burns with blue-green flame forming a black surface coating of copper (II) oxide.



(ii) **Reaction with non-metals**

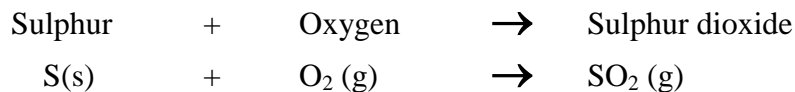
Non-metals react with oxygen to form *non metallic oxides*. Some of these oxides dissolve in water to form acidic solution. Such oxides are called *acidic oxides* or *acid anhydrides*.

Examples include: CO₂, SO₂, NO₂.



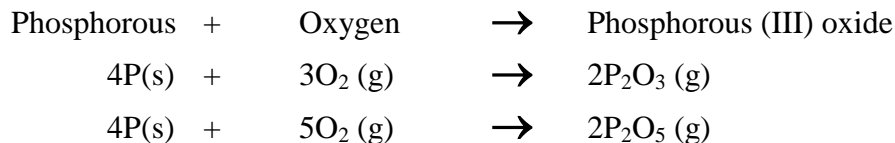
⇒ **Sulphur**

Sulphur burns with purple blue flame forming white fumes of sulphur dioxide.



⇒ **Phosphorous**

Phosphorous burns with a very bright yellow flame forming white smoke which is a mixture of *phosphorous pent-oxide* and *phosphorous tri-oxide*.



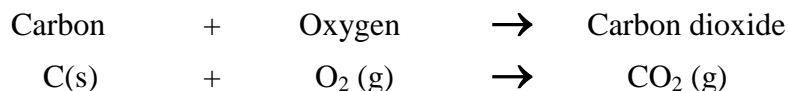
⇒ **Carbon:**

The combustion of carbon is divided into two, namely;

- Complete combustion
- Incomplete combustion

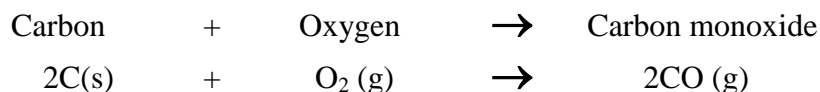
(i) **Complete Combustion**

Carbon burns in excess air or oxygen to form carbon dioxide.



(ii) Incomplete Combustion

In incomplete supply of oxygen i.e. in limited air, carbon burns to form carbon monoxide according to the equation:



7.32 Manufacture (Industrial Preparation) of Oxygen Gas

Oxygen is manufactured by a process called *fractional distillation of liquid air*.

- First, the air is filtered to remove unwanted components such as water vapor, dust and other particles. (They must be removed before the air is liquefied or they will freeze in the column where distillation occurs).
- The remaining air is compressed and passed through beds of adsorption beads to remove carbon dioxide and the remaining water vapour.
- The air expands and cools until it liquefies at about -190°C .
- The liquid air is fed into the top of a distillation column. At the same time, the gaseous portion of the air stream (the part that is still compressed) is fed into the bottom of the column filled with perforated trays.
- The gaseous air rises up through the column, bubbling through the liquid that trickles down the column through the trays. As the liquid trickles down, it heats up and reaches the boiling point of nitrogen. The nitrogen boils off first near the top of the column at a boiling point of -196°C and is followed by argon at a boiling point of -186°C .
- Oxygen having a higher boiling point than that of argon or nitrogen, remains in liquid form until it reaches the bottom of the column, where the temperature is highest, and then boils off at a boiling point of -183°C .

7.4 Oxides

Oxides are compounds that consist of an element (metal or non-metal) and oxygen only. They are formed when elements combine with oxygen.

Word equation: *Element* + *Oxygen* → *Oxide*

Classification of oxides

Oxides are classified into seven groups, namely;

- (i) Acidic Oxides
- (ii) Basic oxides
- (iii) Neutral Oxides
- (iv) Amphoteric Oxide
- (v) Higher Oxides
- (vi) Peroxide
- (vii) Mixed Oxides

(a) Acidic Oxides

Acidic oxides are oxides of nonmetals which dissolve in water to form acidic solutions. E.g. CO_2 , SO_2 , SO_3 , NO_2 .

(b) Basic oxides

A basic oxide is an oxide of metal which reacts with acids to form salt and water only. E.g. CaO , CuO , MgO , ZnO etc.

(c) Neutral Oxides

A neutral oxide is an oxide which does not react either with an acid or alkali to form a salt. Examples of a neutral oxides are CO and NO .

(d) Amphoteric Oxide

An amphoteric oxide is an oxide which behaves both as an acid and base.
E.g. ZnO , Al_2O_3 and PbO .

(e) Higher Oxides

A higher oxide is an oxide which contains more oxygen atom than the expected.
E.g. Lead (IV) oxide (PbO_2). But the common (expected) oxide of lead is lead (II) oxide (PbO).

(f) Peroxide

A peroxide is a type of higher oxide containing two oxygen atoms.
E.g. Hydrogen peroxide (H_2O_2), Sodium peroxide (Na_2O_2).

(g) Mixed Oxides

These are oxides of two metals. Examples of mixed oxide include:-

$\text{Zn(FeO}_2\text{)O}_2$, Fe_3O_4 etc.

Uses of Oxygen

1. Respiration

Oxygen is used to aid respiration where the natural supply is insufficient. As seen in the following:-

- (i) In hospitals: It is administered to patients with breathing difficulty to aid respiration
- (ii) Mountain climbing.
- (iii) Deep sea diving.
- (iv) In high altitude flying.

3. Liquid oxygen is used to burn fuel in rocket engines.

4. Liquid oxygen is used in explosives.

5. Oxygen is used in combustion of substances.

Self-Check 7.0

1. The rate of decomposition of hydrogen peroxide can be increased by the addition of
- | | |
|---------------------------------|---------------------------------|
| A. copper (II) sulphate. | B. manganese (IV) oxide. |
| C. vanadium (V) oxide. | D. finely divided iron |

2. The oxide that dissolves **most readily** in water is
A. sodium oxide. B. calcium oxide.
C. lead (II) oxide. D. copper (II) oxide.
3. Which one of the following is observed when magnesium burns in air?
A. A bright flame. B. A white powder.
C. Brown specks. D. Molten beads.
4. Which one of the following oxides is a mixed oxide?
A. Al_2O_3 . B. ZnO .
C. Fe_3O_4 . D. Na_2O .
5. A reaction does **not** take place when one of these mixtures is heated. Which is it?
A. Zinc and aluminium oxide. B. Zinc and copper oxide.
C. Magnesium and Zinc oxide. D. Zinc and lead (II) oxide.
6. Sodium burns in excess air forming
A. sodium oxide. B. sodium peroxide.
C. sodium hydroxide. D. sodium hydride.
7. Which one of the following oxides is **not** amphoteric?
A. ZnO . B. Al_2O_3 .
C. Fe_2O_3 . D. PbO .
8. Which one of the following oxides is amphoteric?
A. Calcium oxide. B. Aluminium oxide.
C. Sodium oxide. D. Copper oxide.
9. Which one of the following oxides is soluble in water?
A. Lithium oxide. B. Iron (III) oxide.
C. Aluminium oxide. D. Copper (II) oxide.
10. Which one of the following is **not** an acidic oxide?
A. Carbon dioxide. B. Carbon monoxide.
C. Sulphur dioxide. D. Phosphorus (V) oxide.
11. Which one of the following colours is observed when sodium is burnt in air?
A. Blue. B. Yellow.
C. Green. D. Purple.
12. Which one of the following oxides will form a metal when heated with aluminium?
A. $\text{K}_2\text{O(s)}$. B. MgO(s) .
C. $\text{Na}_2\text{O(s)}$. D. PbO(s) .
13. Which one of the following metals can displace hydrogen from dilute hydrochloric acid?
A. Copper. B. Magnesium.
C. Lead. D. Silver.

14. Beginning with the most reactive, the order of reactivity of the following metals with water is

A. Sodium → Magnesium → Lead → Copper.
B. Magnesium → Sodium → Copper → Lead.
C. Copper → Lead → Magnesium → Sodium.
D. Lead → Copper → Sodium → Magnesium.

15. Which one of the following metals can displace zinc from zinc sulphate solution?

A. Cu. B. Pb.
C. Ca. D. Ag.

16. The cation that can **not** be reduced by aluminium metal is

A. Zn^{2+} (aq) B. Fe^{2+} (aq)
C. Mg^{2+} (aq) D. Cu^{2+} (aq)

17. Which one of the following metals does **not** displace iron from a solution of iron (III) nitrate?

A. Zinc. B. Copper.
C. Magnesium. D. Aluminium.

18. A steel tank is protected from rusting when it is connected to a metal, M as shown.



Which one of the following metals is **not** likely to be M?

A. Aluminium. B. Zinc.
C. Magnesium. D. Copper.

19. Which one of the following metals reacts with water at ordinary temperature to give hydrogen?

A. Aluminium. B. Calcium.
C. Zinc. D. Iron.

20. An oxide of metal Q can be reduced by metal K but not by metal Z. The order of the reactivity of these metals is

A. Z, K, Q. B. K, Z, Q.
C. Q, K, Z. D. K, Q, Z.

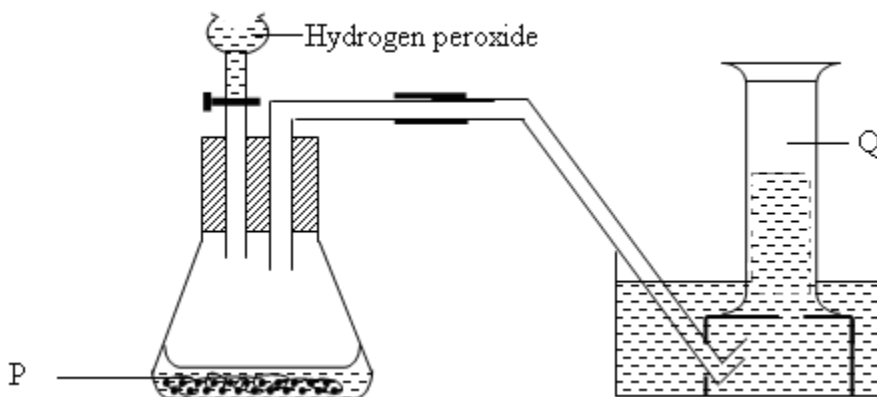
SECTION B

21. When a test tube containing cold hydrogen peroxide was brought near a non-luminous flame of Bunsen burner, a colourless gas was evolved.

(a) (i) Name the gas and give its formula.

- (ii) State the test for the gas.
- (iii) Write the equation for the reaction leading to the formation of the gas.
- (b) Name one substance that can be used to speed up the rate of reaction of formation of the gas.

22. Study the diagram in the figure below.



- (a) Identify P and Q.
 - (b) Give the equation for the reaction leading to the formation of gas Q.
 - (c) Describe what is observed when burning:
 - (i) Sulphur is plunged into a jar filled with gas Q.
 - (ii) Magnesium is plunged into a jar filled with gas Q.
 - (d)
 - (i) Give the equations of the reactions in c (i) and (ii).
 - (ii) The products in (c) were separately dissolved in water and the resulting solutions were tested with blue litmus paper. State which solution turned the blue litmus paper red? Explain your observation.
 - (e)
 - (i) State one method by which gas Q is obtained commercially
 - (ii) Give any **two** uses of gas Q.
23. (a) Classify each of the following oxides as acidic, amphoteric, basic or neutral.
- (i) Calcium oxide
 - (ii) Aluminium oxide
 - (iii) Zinc oxide

- (iv) Copper (II) oxide
 - (v) Carbon monoxide
 - (vi) Sulphur (IV) oxide
- 24.**
- (a)
 - (i) Draw a labelled diagram to show how a sample of oxygen can be prepared in laboratory from potassium chlorate.
 - (ii) Write the equation for the reaction that takes place.
 - (b) State and explain what happens when each of the following substances are lowered in a gas jar of oxygen and water added to the products.
 - (i) Burning sodium,
 - (ii) Ignited magnesium,
 - (iii) Hot iron.
 - (c) Name **one** natural process by which oxygen can be obtained.
- 25.** Oxygen can be prepared in the laboratory using hydrogen peroxide and a substance X.
- (a) Name substance X.
 - (b) Write equation leading to the formation of oxygen.
 - (c) State:
 - (i) The role of X.
 - (ii) The conditions for the reaction.
 - (d) Name **one** process that increases the amount of oxygen in the atmosphere.

CHAPTER EIGHT

WATER AND HYDROGEN

Learning objectives

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

1. (a) State: (i) Physical and chemical properties of water.
(ii) Sources of water.
(b) Describe the process of water purification.
 2. (a) Explain water pollution.
(b) State the harmful effects of water pollution to living organisms.
(c) State the uses of water.
 3. (a) Define the following terms:
- Unsaturated solution, saturated solution, super-saturated solution and suspension:
(b) Differentiate between solution and suspension.
-

8.0 Water (H₂O)

Water is an oxide of hydrogen gas. Its chemical formula is H₂O. The chemical name of water is *Hydrogen oxide*. Water is a compound because it consists of two elements (hydrogen and oxygen) chemically combined.

(a) Physical States of water

Water exists in three states, namely:

- Liquid,
- Solid (ice), and
- Gas (Water vapour)

(b) Sources of water

The sources of water include the following:-

- Rain water,
- Spring water (well water),
- River water,
- Lake water,
- Sea water,
- Ocean water and
- Tap water.

Water from these sources is never pure. This is because it contains dissolved substances and suspended solid particles. However, impure water can be purified (made pure) by a process called *purification*.

(c) Purification of water

A sample of fairly pure water can be obtained from the above sources by a process called *distillation*.

The Process of Distillation

- Impure water is heated in a distillation flask to the boiling point 100 °C.
- The steam that comes off is condensed in a *lie big condenser*.
- The solid impurities are left in the distillation flask. (See Chapter 3 section 3.5)

Properties of Water

(a) Physical Properties.

- (i) It freezes at a temperature of 0 °C (273 K).
- (ii) It boils at a temperature of 100 °C (373 K) when pressure is 760 mmHg or 1 atmosphere.
- (iii) Its maximum density is 1 g/cm³ at a temperature of 4 °C.
- (iv) It is neutral to litmus.
- (V) Water is an excellent solvent.

(b) Chemical Properties

(i) Chemical test for water

Water is tested by using the following compounds.

- Anhydrous Copper (II) sulphate and
- Cobalt chloride.

With Anhydrous Copper (II) sulphate:

- Water turns white crystals of anhydrous copper (II) sulphate to blue.

With Cobalt Chloride:

- Water turns blue cobalt chloride to pink.

(ii) Action of water on metals.

K	}	<i>Attack water</i>
Na		

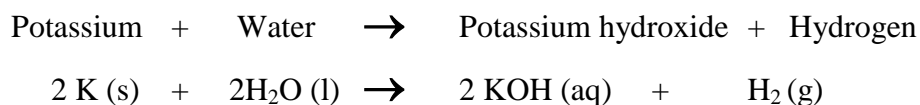
Ca	}	<i>Attack steam</i>
Mg		
Al^*		
Zn		
Fe		

Pb	}	<i>Do not attack water or steam</i>
Cu		
Hg		
Ag		
Au		

**Al: Does not attack water or steam because of the oxide layer formed which is insoluble and prevents contact between water and the metal.*

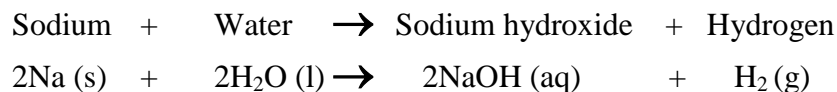
❖ Action of Potassium on water:

- Potassium burns violently with a purple flame as it darts on water forming potassium oxide and hydrogen gas.
- The oxide reacts immediately with the water to form potassium hydroxide solution and hydrogen gas.



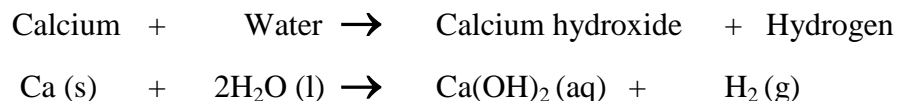
❖ Action of Sodium on water

- Sodium reacts violently with water, but does not burn unless it is restricted in movement.
- It melts into silvery ball which runs on the water with a hissing sound.
- Effervescence occurs liberating a colourless gas.
- The final products are sodium hydroxide and hydrogen gas.



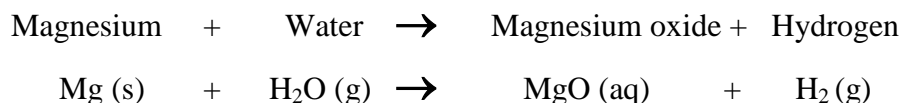
❖ **Action of Calcium on water**

- Calcium sinks and reacts steadily with effervescence giving off a colourless gas (hydrogen).
- The calcium gradually disappears and a white milky suspension is formed. This cloudiness is due to the formation of un-dissolved calcium hydroxide which is white.



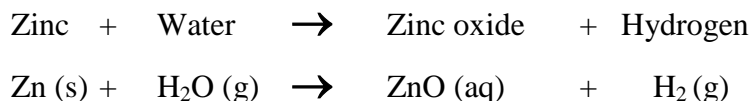
❖ **Action of Magnesium on water**

Magnesium reacts very slowly with cold water. Heated magnesium burns in steam with a bright flame forming white ash.



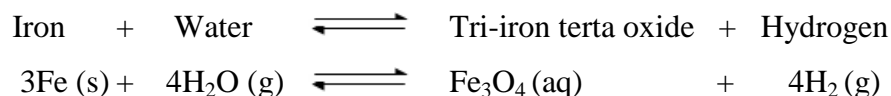
❖ **Action of Zinc on water**

Zinc does not attack hot or cold water. It reacts only when heated strongly to redness in a current of steam. Zinc oxide and hydrogen are formed.



❖ **Action of Iron on water**

Iron does not attack water (rusting takes place only when air is present as well) but it is readily attacked by excess of steam at red hot to form tri-iron tetra oxide. The reaction is reversible.



- NB:**
- With cold water iron rusts as long as oxygen is present.
 - Lead dissolves in water containing oxygen.
 - Copper and metal below hydrogen in the activity series have no action on water or steam.

8.1 Water Pollution:

Pollution is the process of making something pure impure by adding harmful or unpleasant substances called pollutants.

Pollutant in rivers, seas and lakes include:-

- Sewage (wastes from towns and industries),
- Nitrates (washed by rain from fertilizers used in agriculture),
- Insecticides,
- Detergents,
- Mercury and
- Lead compounds.

Harmful Effects of Water Pollution:

- (i) *Sewage:* Contaminates water leading to water born diseases such as typhoid, bilharzias etc.
- (iii) *Mercury* from industrial wastes accumulates in fish and may cause brain damage to people who eat such fish.
- (iii) *Lead compounds* from car exhaust pipes are washed into rivers, lakes and seas. Lead is dangerous poison. In that it replaces calcium ions in the teeth and bones in the body.
- (iv) *Nitrates* promote the growth of green algae which use up dissolved oxygen in water thus depriving aquatic animals of oxygen. This leads to the death of these animals.
- (v) *Insecticides* do not decompose readily and accumulate in water.
- (vi) *Detergent solution* used by human beings is carried into the water bodies. They are non-biodegradable and accumulate in water.

Uses of Water

(a) Domestic uses:

In homes, water is used for:

- (i) Drinking,
- (ii) Cooking,
- (iii) Washing,
- (iv) Bathing.

(b) Industrial uses

- (i) In the industries, water is used for cooling hot machine parts.
- (ii) It is also used in food processing for mixing ingredients.
- (iii) It is used as a solvent.

(c) Agriculture

- (i) Water is used in irrigation to grow crops.
- (ii) It is used by both plants and animals for growth.

(d) Sciences

- (i) In sciences, water is used as a solvent to make solution
- (ii) Running water is used for generating Hydro-Electric Power.

8.2 Water and Solutions

Water is a universal solvent. That is it dissolves many substances. Substances that dissolve in water are said to be *soluble* and are called *solutes*. Those which, do not dissolve are said to be *insoluble*.

When a solute dissolves in a solvent it forms a homogenous mixture called *solution*.

Definitions:

A solution is a uniform (homogenous) mixture of two or more substances.

A solute is a substance that dissolves in a solvent.

A solvent is a substance which dissolves a solute.

NB: *A solution in water is called an aqueous solution. Aqueous means dissolved in water.*

(a) Types of Solutions:

There are three types of solutions, namely:-

- Unsaturated solution.
- Saturated solution and
- Supersaturated solution.

(i) Unsaturated solution.

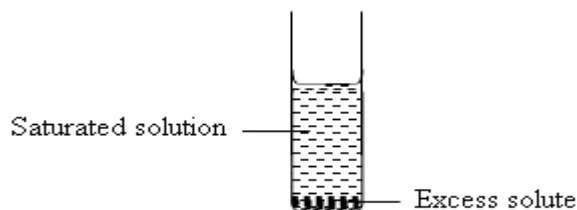
An unsaturated solution is one which can dissolve more solute at the same temperature.

E.g. a spatula endful of sodium chloride (common salt) added into a beaker of water dissolves quickly and one more can still be dissolved if added.

(ii) Saturated solution:

A saturated solution of a solute at a particular temperature is one which can dissolve no more solute at that temperature.

This is known when some solute fails to dissolve and instead settles at the bottom of the container as shown in the diagram below.



(iii) Super-saturated solution:

A super-saturated solution is one that contains more of the solute than it could hold at that temperature if crystals of the solute were present.

Super-saturated solutions are in unstable condition and they can only be obtained from a few pure compounds.

E.g. Hydrated salts of sodium.

- Hydrated sodium sulphate ($\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$).
- Hydrated Sodium Thiosulphate ($\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$).

A super-saturated solution can be made from hydrated sodium sulphate ($\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$), by cooling the solution slowly.

While from Hydrated Sodium Thiosulphate ($\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$), by cooling the solution quickly.

Care should be taken to:-

- exclude all dust which provides centres for crystallization.
- avoid shaking or any form of disturbance of the solution.

(b) Suspension:

A suspension is a liquid containing small particles of solid spread throughout it and the particles settle on standing.

(c) Differences between Solution and Suspension:

A suspension differs from a solution in three ways:

- (i) It contains solid particles which can be seen.
- (ii) Its solid particles settle on standing.
- (iii) Filtration separates it into a filtrate and residue.

HYDROGEN (H₂)

Learning objectives

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

- Describe the laboratory preparation of hydrogen.*
 - Name the catalyst used in the preparation of hydrogen.*
 - Describe test for hydrogen.*
- State the physical and chemical Properties of hydrogen.*
- State the uses of hydrogen.*

8.3 Introduction

Hydrogen, the lightest and the smallest atom of any chemical element. It exists as a gas at room temperature. It was named by French chemist *Antoine Lavoisier* from the Greek words for “water former.” A hydrogen atom contains one proton, and one electron revolving around the nucleus as a result it is the first element in the periodic table of the elements and is represented by the symbol H.

Occurrence:

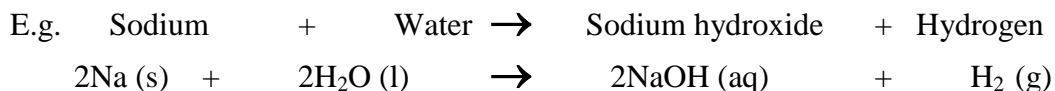
Pure hydrogen gas rarely occurs in nature, although volcanoes and some oil wells release small amounts of hydrogen gas. It occurs in a combined state in compounds such as:

- Organic compounds (Carbon compounds) such as:
 - Fats, proteins, carbohydrates and Alcohols
- Hydrocarbon - Alkanes, alkenes and alkynes
(Studied in Organic Chemistry).
- Inorganic compounds (Compounds that do not include carbon), such as:
 - Water (H₂O), Ammonia (NH₃), Acids and Metallic hydroxides.

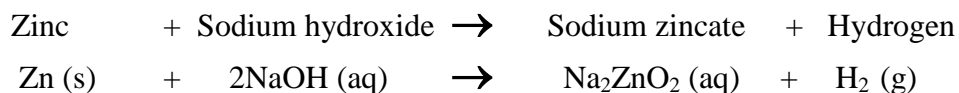
8.31 Preparation of Hydrogen

Hydrogen can be prepared from the following sources.

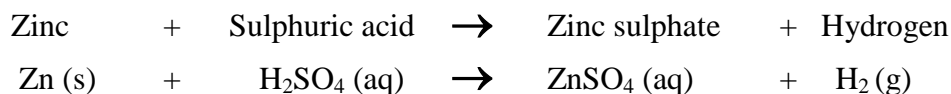
❖ *Reactions of some metals with water.*



❖ *Reactions of some metals with alkali.*



❖ *Reactions of dilute acids with some metals above hydrogen in Electro Chemical Series.*

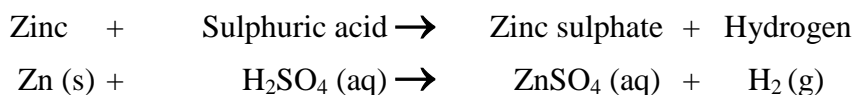
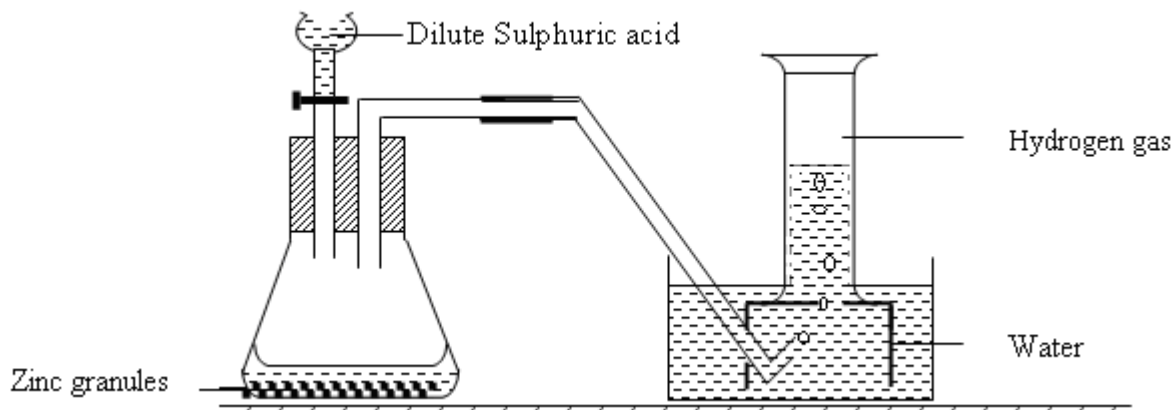


However, some of the above reactions are not convenient for laboratory preparation of hydrogen except the action of dilute acids on metals.

Laboratory Preparation of Hydrogen:

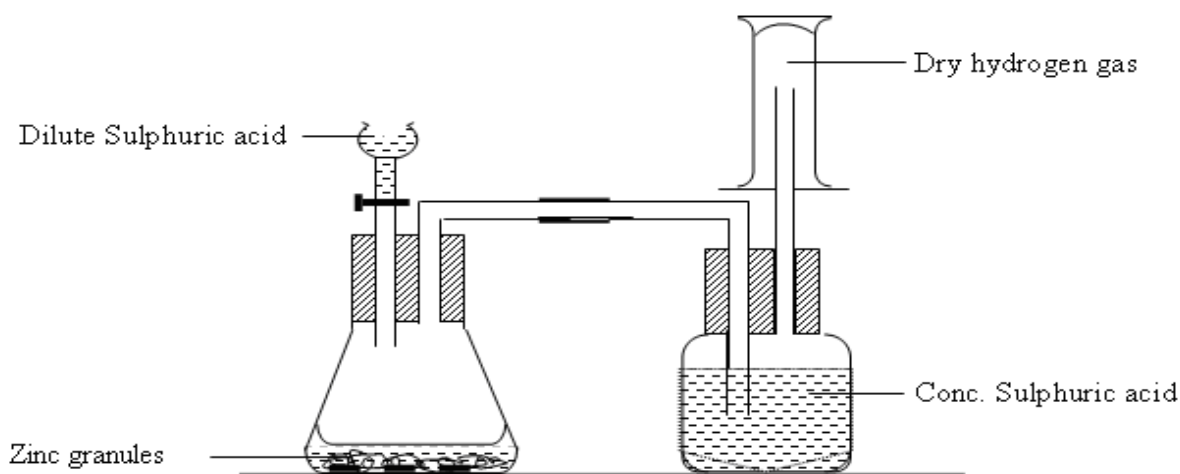
In the laboratory, Hydrogen is usually prepared by dropping dilute sulphuric acid on zinc granules in a flat bottom flask. The reaction can be speeded up by adding copper (II) sulphate as a catalyst and the gas is collected over water as shown in the diagram below.

Diagram showing Laboratory Preparation of Hydrogen



If the gas is required dry, it is dried by passing it through concentrated sulphuric acid and then collected by upward delivery method as the gas is lighter than air.

Diagram showing the laboratory preparation of hydrogen



Test for Hydrogen:

When a burning splint is plunged into a container containing a mixture of hydrogen and air, it burns with a pop sound.

Properties of Hydrogen:

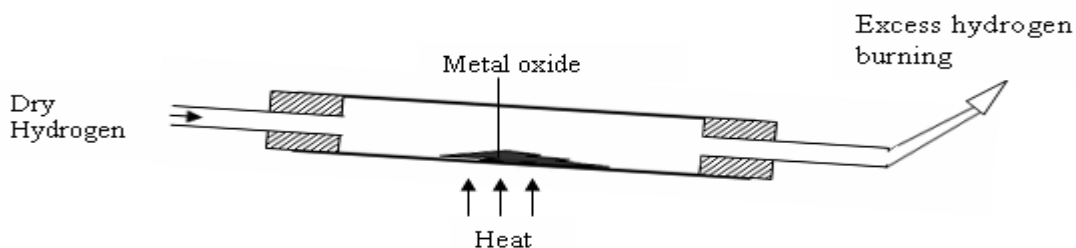
(a) Physical Properties

- (i) It is a colourless gas
- (ii) It is odourless gas.
- (iii) It is tasteless.
- (iv) It is the lightest gas.
- (v) It is insoluble in water.

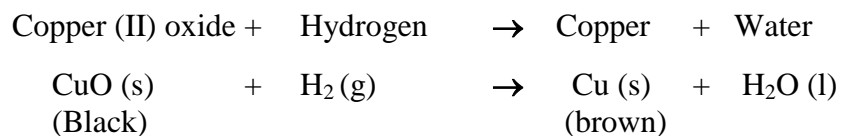
(b) Chemical Properties

(i) Reactions with some heated oxides.

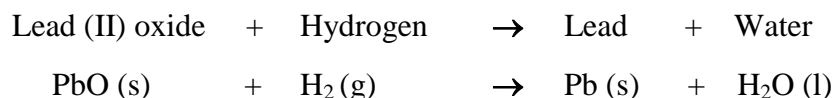
When hydrogen is passed over heated oxides of copper (II) oxide (CuO), lead (II) oxide (PbO) and tri-iron tetra oxide (Fe_3O_4), in a hard glass tube, the oxides are reduced to the metals and the hydrogen is oxidized to a colourless liquid which turns white anhydrous copper (II) sulphate blue.



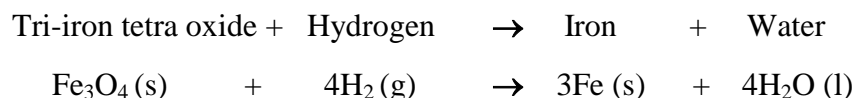
➤ With Copper (II) oxide, the black oxide turns to brown.



❖ With Lead (II) oxide:



❖ With tri-iron tetra oxide:



NB: - The above reactions are called redox reaction. Redox reaction is a type of reaction where both **Reduction** and **Oxidation** occur simultaneously.

Oxidation: - is the addition of oxygen to a substance.

Reduction: - is the removal of oxygen from a substance.

In the above reactions the oxides are called *oxidizing agents* and the hydrogen is called a *reducing agent*.

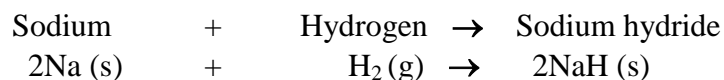
Oxidizing agent: - is a substance that adds oxygen to a substance.

Reducing agent: - is a substance that removes oxygen from a substance.

Hydrogen does not reduce metals above iron in the reactivity series.

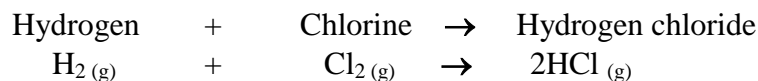
(ii) Reactions with some metals.

Hydrogen combines with some metals forming hydrides. For example, it reacts with heated sodium forming sodium hydride.



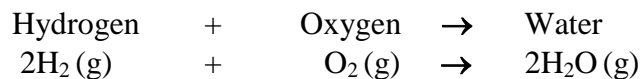
(iii) Reactions with some non-metals.

Hydrogen reacts with non-metals forming chlorides. For example, it reacts with chlorine forming misty fumes of hydrogen chloride.



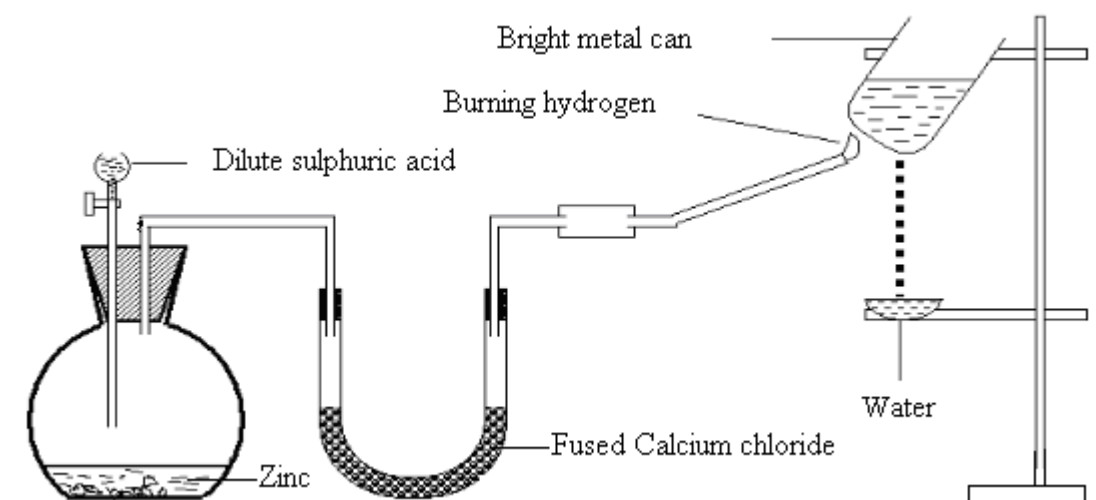
(iv) Combustion

Hydrogen gas does not support burning. It burns quietly with a pale blue flame forming steam as the only product. When the steam is cooled, it condenses to liquid water.



The steam can be condensed into liquid water as shown in the diagram below.

The diagram showing burning of Hydrogen to form water



Uses of Hydrogen

Hydrogen is used:

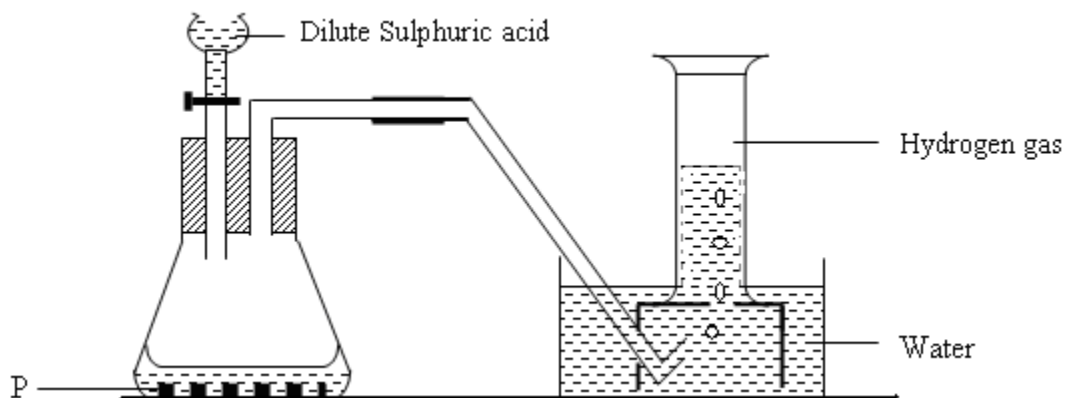
- (i) In the manufacture of ammonia by Haber process.
- (ii) To harden vegetable oils into solid fats such as margarine, 'kimbo' and blue band.
- (iii) As fuel (liquid hydrogen) in rockets.
- (iv) To produce oxy-hydrogen flame for welding.
- (v) In filling balloons as it is lightest gas.
- (vi) In converting coal into petrol.

Self-Check 8

- The formula of the oxide formed when steam is passed over heated iron is
A. FeO. B. Fe₂O₃.
C. Fe(OH)₃. D. Fe₃O₄.
- Hydrogen is used on a large-scale to:
A. manufacture detergents. B. harden vegetable oils.
C. reduce metal oxides in extraction of metals. D. make mining explosives.
- Which one of the following properties of hydrogen is applied during its collection by upward delivery?
A. It is slightly soluble in water.
B. It is a gas with low boiling point.
C. It is less dense than air.
D. It forms an explosive mixture with oxygen.
- Which one of the following substances can be obtained by a process of hydrogenation of oil?
A. Petroleum. B. Margarine.
C. Soap. D. Rubber.
- When hydrogen is passed over heated copper (II) oxide, a brown solid is formed. This shows that
A. copper is above hydrogen in the activity series.
B. copper is an oxidising agent.
C. hydrogen is a reducing agent.
D. hydrogen is inert towards metals.
- Which one of the following substances will react with magnesium to give hydrogen?
A. Dilute hydrochloric acid. B. Aqueous ammonia.
C. Potassium hydroxide. D. Ethanol.
- In the preparation of hydrogen from zinc and dilute hydrochloric acid, the rate of reaction is increased by adding
A. nickel B. copper sulphate
C. platinum D. manganese dioxide
- Which one of the following reactions can be used to prepare hydrogen in the laboratory?
A. Reacting calcium with dilute sulphuric acid.
B. Reacting sodium with water.
C. Reacting zinc with dilute hydrochloric acid.
D. Reacting magnesium with steam.
- Which one of the following elements can reduce aluminium oxide when heated together strongly?
A. Calcium. B. Lead.
C. Iron. D. Copper.
- A reaction does not take place when one of these mixtures is heated. Which is it?
A. Zinc and aluminium oxide. B. Zinc and copper oxide.
C. Magnesium and Zinc oxide. D. Zinc and lead (II) oxide.

SECTION B

11. The diagram below shows a method used in the preparation of hydrogen in the laboratory



- (a)
 - (i) What is metal P?
 - (ii) Write an equation of the reaction taking place in the flask
 - (iii) Why is it possible to collect hydrogen over water?
 - (iv) Describe one chemical test that can be used to identify hydrogen gas.
 - (b)
 - (i) Name one impurity in the gas.
 - (ii) How can the impurity in (i) be removed?
 - (iii) The reaction above can be made faster by addition of a catalyst. Name the catalyst.
 - (c)
 - (i) If the hydrogen is passed over heated copper (II) oxide for a long time, what would you observe?
 - (ii) What type of the reaction has hydrogen undergone?
 - (iii) Give the equation of the reaction between the oxide and hydrogen.
12. (a) (i) State the conditions under which magnesium can react with water.
 (ii) Write equation for the reaction that takes place in (a)(i).
- (b) The product from (a) was shaken with water. State what was observed.
- (c) Dilute hydrochloric acid was added to the product in (b). Write an ionic equation for the reaction that took place.
13. Magnesium reacts with steam to give a solid P and a gas Q.
- (a) Identify:
 - (i) Solid P.
 - (ii) Gas Q.
 - (b) State how Q could be tested.
 - (c) Write equation for the reaction between P and hydrochloric acid.
14. Figure 2 shows a set up of apparatus to investigate the reaction between metals and steam.

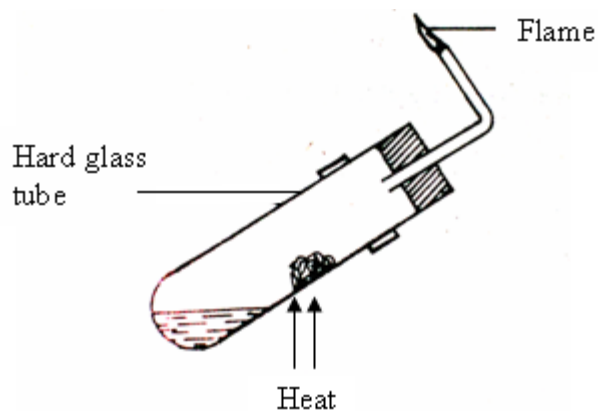


Figure 2

- (a) Suggest a suitable metal that could be used in the experiment.
 - (b) (i) What would be observed in the glass tube?
(ii) Write the equation for the reaction in the glass tube.
 - (c) (i) Suggest the gas that is being burnt at the end of the tube.
(ii) Write the equation for the combustion of the gas in (c)(i).
15. (a) Define the following terms.
- (i) Solution
 - (ii) Saturated solution
 - (iii) Super saturated solution
 - (iv) Suspension
- (b) (i) What is the difference between a saturated solution and an unsaturated solution?
(ii) How can you prove that a solution of sodium chloride in water is saturated?
- (c) State: (i) **three** differences between a solution and a suspension.
(ii) the differences between a saturated solution and a super saturated solution.

CHAPTER NINE

ACIDS AND BASES

Learning objectives

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

1. (a) Define the term acid.
(b) State the types of acids and give examples for each.
 2. (a) Define the term base.
(b) Give examples of bases that are soluble in water.
 3. (a) Define the term indicator.
(b) State the effect of acids and bases on indicators.
(c) Identify solutions as acids or bases using the pH values.
 4. State: (i) the properties of acids and bases.
(ii) Uses of bases.
-

Acids and Bases are two classes of chemical compounds that display generally opposite characteristics. These characteristics are discussed below.

9.0 Acids

Definition: An acid is a compound which, when dissolved in water, produces hydrogen ions, $H^+(aq)$, as the only positive ion.

9.1 Types of acids

There are two types of acids namely:

- (i) Organic acids.
- (ii) Mineral acids (inorganic)

(a) Organic Acids

These are acids that occur naturally in living things. Examples of organic acids are shown in table 9.1.

Acid	Where it is found
Citric acid	Citrus fruits like lemon and oranges
Tartaric acid	Grapes
Acetic (ethanoic) acid	Vinegar
Formic (methanoic) acid	Ants' and bees stings and stinging beetles
Lactic acid	Sour milk

Table 9.1 Organic acids

(b) Mineral acids

These are acids made in the laboratory from substances derived from minerals. The most common mineral acids are shown in table 6.2 below.

Acid	Chemical formula
Hydrochloric acid	HCl
Nitric acid	HNO ₃
Sulphuric acid	H ₂ SO ₄
Phosphoric acid	H ₃ PO ₄

Table 9.2 Mineral acids

- NB:**
- The mineral acids are strong acids. (See chapter 12 section 12.1).
 - They burn the skin and other materials. For this reason they are said to be corrosive and therefore should be handled with a lot of care.

9.2 Bases

Definition: A base is a compound which contains oxide (O^{2-}) or hydroxide (OH^-) ions and reacts with an acid to form a salt and water only.

Bases are oxides and hydroxides of metals. They neutralize acids (i.e. they cancel the effect of acids when the two are mixed).

(a) Solubility of Bases

Most bases are insoluble in water but a few are soluble. The soluble bases form or comprise a special class called *alkalis*.

Alkalis An alkali is a compound which, when dissolved in water, forms hydroxide ions (OH^-) as the only negatively charged ions.

NB: Like the strong acids, strong alkalis are very corrosive and should be handled with a lot of care.

The table 9.3 below shows the common bases and their solubility.

Base	Chemical Formula	Solubility
Potassium hydroxide	KOH	Highly soluble
Sodium hydroxide	NaOH	
Calcium hydroxide	Ca(OH) ₂	Moderately soluble
Magnesium hydroxide	Mg(OH) ₂	Slightly soluble
Copper (II) oxide	CuO	All are insoluble
Copper (II) hydroxide	Cu(OH) ₂	
Iron (II) oxide	FeO	
Iron (II) hydroxide	Fe(OH) ₂	
Iron (III) oxide	Fe ₂ O ₃	
Iron (III) hydroxide	Fe(OH) ₃	

Table 9.3 Common bases and their solubility

The summary of solubility and the effect of heat on nitrates

Metal	Solubility	Effect of heat
K Na Ca	The hydroxides of these metals are soluble in water and are alkalis	Hydroxides of K and Na do not decompose by heat.
Mg Al Zn Fe Pb Cu	These metals form hydroxides which are insoluble in water.	Hydroxides of these metals decompose into oxide and water.
Hg Ag Au	Hydroxides of these metals do not exist.	

9.3 Indicators

These are substances that change colour in acids and bases. They are therefore used to detect the presence of acids and bases.

Types of Indicators

There are mainly three types of indicators namely:

- (i) Litmus
- (ii) Plant extracts
- (iii) Commercial Indicators

(a) Litmus

This is the most common type of indicators. It turns blue in a basic solution and red in acidic solution. Litmus can be used either as a solution or more commonly it is soaked on a paper which is then dried forming litmus paper. There are two colours of litmus papers. These are red and blue.

(b) Plant Extract Indicators

These are indicators extracted from plant parts such as flowers, leaves and other parts. The extracts show distinct colours in acidic and basic solutions.

Experiment 9.1 To Extract Indicators from Plant parts

Apparatus/Chemicals

Mortar and pestle, Ethanol, red cabbage leaves and hibiscus

Procedure

- (a) Cut flowers of hibiscus into pieces and grind them using mortar and pestle.
- (b) Add a little of ethanol or propanol and continue grinding until you obtain a fair amount of extract as deep as possible.
- (c) Decant about 3 cm³ of the extract in to a test tube.
- (d) Repeat the procedure (a) to (c) using red cabbage leaves.

Observations

Yellowish/Green and Reddish colouration are obtained from the flowers of hibiscus and the red cabbage leaves respectively.

The fresh extracts are indicators and can be used to test acidic and basic solutions of substances.

Experiment 9.2 To classify substances in to acids and bases using a plant extract indicator

Apparatus

9 test tubes, test tube rack, droppers, cabbage leave extract and solutions of the substances listed in table 9.4 below.

Procedure

Add 3–4 drops of freshly prepared extract from red cabbage leaves (as described in experiment 9.1) to 2 cm³ of solutions of the substances shown in the table 9.4 below in different test tubes.

Record any colour changes.

Observations/Results

The colour changes upon addition of the red cabbage extract to the various solutions are shown in table 9.4.

Test Solution	Colour observed on addition of cabbage extract
Hydrochloric acid	Red
Nitric acid	
Sulphuric acid	
Vinegar	
Lemon juice	
Potassium hydroxide	Blue
Sodium hydroxide	
Calcium hydroxide	
Ammonia	

Table 9.4 Colour of common substances in cabbage extract

Disadvantages of Plant Extract indicators

- (i) Most plant extracts are unstable indicators and change colour on long exposure. For this reason, they are only reliable as acid-base indicators only when freshly prepared.
- (ii) They can not be used to show strength of acids and bases.

(c) **Commercial Indicators**

These are commercially prepared acid-base indicators. They are stable and last for a long time without changing colour.

Examples of commercial indicators are:

- Litmus
- Phenolphthalein,
- Methyl orange and
- Bromothymol blue.
- Universal indicator

Colour of Indicators in acids and bases

The various colours shown in table 9.5 are obtained when the indicators listed below are added to acidic and basic solutions.

The table below shows the colours of the indicators in the laboratory in different media.

Indicator	Neutral	Colour in	
		Acid	Alkali
Litmus Paper: - Red litmus paper - Blue litmus paper	Red Blue	Red Red	Blue Blue
Phenolphthalein	Colourless	Colourless	Pink/Red
Methyl Orange	Orange	Orange	Yellow
Methyl Yellow	Yellow	Pink	Yellow
Screened Methyl Orange	Pale Blue	Red	Green
Litmus	Purple	Red	Blue
Bromothymol blue	Blue	Yellow	Blue

Table 9.5 Colours of indicators in different media.

(ii) **Universal Indicator**

A universal indicator is a mixture of several indicators used to test acidity (strength of acid) and alkalinity (strength of alkalis). It is sold either in form of a liquid or strips of paper in which the universal indicator has been soaked and then dried. When the indicator is added to different solutions, it shows a range of colours depending on the pH of each of the solution.

NB: *The indicator is sold together with a colour chart which shows its various colours and the corresponding pH range.*

9.4 pH Scale

The pH scale is the scale used to measure the acidity and alkalinity of substances. It ranges from 0 – 14. Substances that are neither acidic nor alkaline (basic) are said to be neutral and have pH of 7. Examples of neutral substances are:

- Pure water and
- Solutions of normal salts. (See Chapter 12 section 12.3).

Solutions with pH less than 7 are said to be acidic. The lower the pH value, the more acidic the solution is.

Solutions with pH greater than 7 are alkaline. The higher the pH value, the more alkaline the solution is.

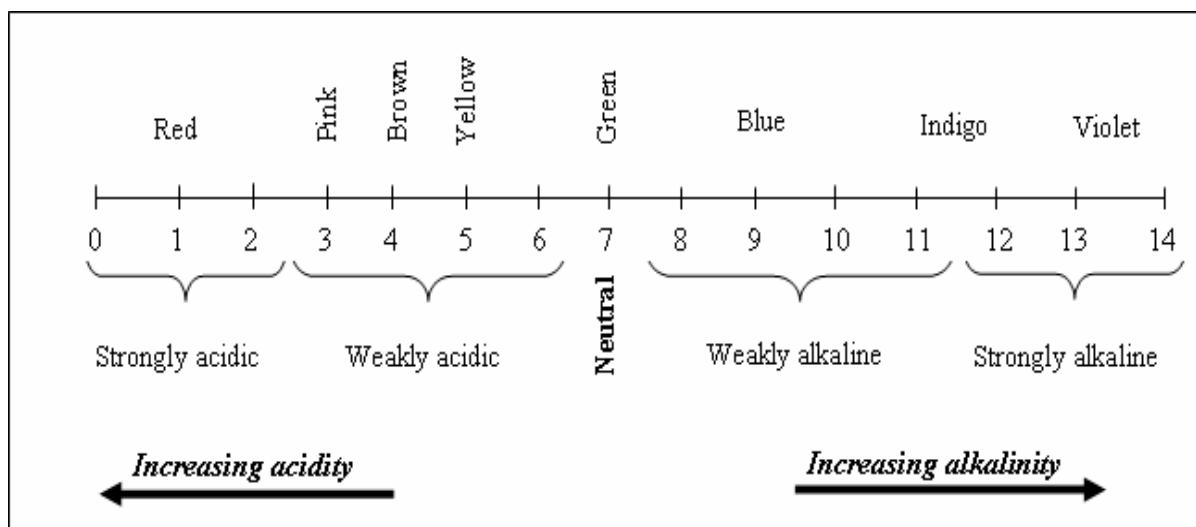


Figure 9.6 Colour code of the universal indicator

Experiment 6.2

To determine the pH of common solutions in the laboratory using Universal Indicator

Apparatus/Chemicals

Test tubes, test tube rack, dropper, Universal indicator/Colour chart, solutions of the substances listed in able 6.6 below.

Procedure

- Place 7 test tubes in a test tube rack.
- Put 1 cm³ of each of the solution in each test tubes.
- Add 3 – 4 drops of the universal indicator into each of the test tubes.

- Note the color changes and match the colour given by each solution with the code of the universal indicator.

Observations

The observations made are shown in table 9.7 on page 128.

Test Solution	pH
Mineral acids: - Hydrochloric acid - Nitric acid - Sulphuric acid - Phosphoric acid	1
Neutral Substances: - Pure water - Solution of normal salts	7
Weak base: - Ammonia	9
Strong basess: - Sodium hydroxide - Potassium hydroxide	14

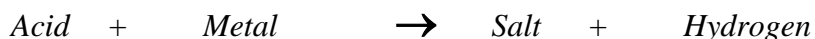
Table 9.7 Common solutions in the laboratory and their pH values

9.6 Properties of acids and Bases

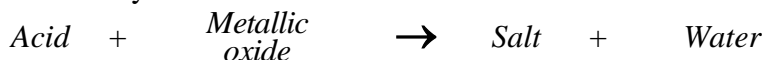
(a) Properties of acids

Acids have the following properties.

- They have a sour taste.
- They have a pH of less than 7.
- Acids turn blue litmus solution red.
- All acids contain chemically combined hydrogen which can be replaced by a metal above hydrogen in the activity series.



- Acids react with bases (oxides and hydroxides of metal) forming salt and water only.



- (vi) Acids react with metallic carbonates and hydrogen carbonate forming carbon dioxide, salt and water only.



(b) Properties of Bases

Bases have the following properties.

- (i) Most bases have bitter taste.
- (ii) They feel soapy.
- (iii) All bases neutralize acids forming a salt and water only.
- (iv) Most bases are insoluble in water but a few are soluble.
- (v) Solutions of soluble bases have pH of greater than 7.
- (vi) Strong alkalis have a powerful corrosive effect on the skin.

(c) Uses of Bases

Bases are used in various fields for different purposes. Uses of some of the common bases are given below.

- (i) *Sodium hydroxide*: It is used:
 - in the manufacture of soaps, paper and many other chemicals.
 - as an oven cleaner
 - for unblocking pipes.
- (ii) *Calcium hydroxide*: is used to neutralize acidic soils in agriculture.
- (iii) *Magnesium hydroxide*: is used to make stomach anti-acids.

Self-Check 9

1. Which one of the following substances will dissolve in water to give a solution that would turn red litmus paper blue?
A. Sodium chloride. B. Sodium hydroxide.
C. Sodium sulphate. D. Sodium nitrate.
2. Which one of the following oxides is soluble in water?
A. Lithium oxide. B. Iron (III) oxide.
C. Aluminium oxide D. Copper (II) oxide.
3. The oxide that dissolves most readily in water is
A. sodium oxide. B. calcium oxide.
C. lead (II) oxide. D. copper (II) oxide.
4. Which one of the following gases dissolves in water to form an acid solution?
A. Nitrogen. B. Ethene.
C. Sulphur dioxide. D. Carbon monoxide.
5. An acid accidentally spilled on a pupil's clothes is best neutralised by a solution of pH
A. 1 B. 4
C. 5 D. 8
6. Which one of the following aqueous solutions will produce hydrogen gas with a magnesium ribbon? That of
A. pH=14 B. pH=7
C. pH=11 D. pH=2
7. Most metals react with dilute mineral acids to form
A. hydrogen gas only.
B. the salt of the metal and water.
C. the salt of the metal only.
D. the salt of the metal and hydrogen gas.
8. Which one of the following is/are property/properties of bases?
(i) They feel soapy.
(ii) All bases neutralize acids forming a salt and water only.
(iii) Solutions of soluble bases have pH of greater than 7.
A. all B. (i) and (ii) only
C. (i) and (iii) only D. (i) only

CHAPTER TEN

STRUCTURE OF THE ATOM AND THE PERIODIC TABLE

Learning objectives

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

1. (a) *Describe the structure of the atom.*
(b) *State the properties of the subatomic particles.*
 2. (a) *Write the electronic configurations of elements.*
(b) *Differentiate between the electronic structure and the atomic structure.*
 3. (a) *Define the following terms:*
 - (i) *Atomic number*
 - (ii) *Mass number (Atomic mass or nucleon number)*
 - (iii) *Isotope*
 - (iv) *Relative Atomic Mass (RAM)*
(b) *Solve problems involving Relative Atomic Mass of isotopes from relative abundance of the isotopes of an element.*
 4. (a) *Recall the characteristics of the Periodic Table.*
(b) *State the physical and chemical properties of:*
 - (i) *Alkali metals.*
 - (ii) *Alkaline Earth metals.*
 - (iii) *Halogens.*
 5. *Explain:*
 - (i) *How atomic size vary across the period (left–right) and down the group.*
 - (ii) *Why noble gases are inert.*
-

(a) The atom

An atom is the smallest (electrically neutral) indivisible particles of an element that can take part in chemical reaction.

(b) The structure of the atom

The atom is made up of smaller particles collectively called *subatomic* particles.

10.1 Subatomic particles

There are three subatomic particles namely:-

- Protons
- Electrons and
- Neutrons

The subatomic particles differ from one another in their properties (e.g. mass and charge) and location.

Properties and location of subatomic particles

(a) Properties

Protons

- (i) They are present in the nucleus of every atom.
- (ii) They have positive charge
- (iii) They have a unit mass i.e. have a mass of 1.

Electrons

- (i) They have negative charge.
- (ii) They have a mass of $\frac{1}{1840}$.
- (iii) They play a decisive role in determining the chemical properties of an element especially those in the outermost shell of an atom.

Neutrons

- (i) They are particles with no charge i.e. are neutral.
- (ii) They have a mass of 1.

(b) Location of the particles

The protons and the neutrons are located in the nucleus (the central region of the atom). They are called nucleons. While the electrons are located outside the nucleus in regions called *shells* or *energy levels*. They revolve around the nucleus the same way that the planets revolve around the sun.

Table 10.1 shows the properties and the location of the subatomic particles

Name of particle	Symbols	Mass	Charge	Location
Proton	p	1	+1	In the nucleus
Neutron	n	1	0	In the nucleus
Electron	e	$\frac{1}{1840}$	-1	Out side the nucleus

Table 10.1

N.B

- ❖ *The atoms of elements are electrically neutral. This is because the number of protons in the nucleus is equal to the number of electrons outside the nucleus. Since their charges are opposite, they cancel; hence the atom has no charge.*
- ❖ *The number of subatomic particles varies from one element to another.*
- ❖ *The masses of subatomic particles are measured by a unit called atomic mass unit (amu).*
- ❖ *1 amu is equivalent to the mass of a hydrogen atom, which is the smallest atom.*

General structure of the Atom

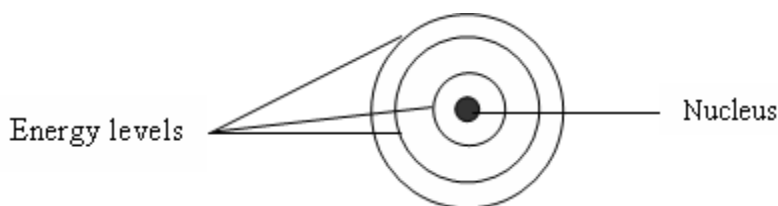


Figure 10.1

N.B:

- ❖ *The shells are oval in shape but for clarity we use perfect circles.*
- ❖ *The nucleus occupies a very small area compared to the overall size of the atom.*
- ❖ *Most of the atom is Empty space with the electrons only occupying a small space around the nucleus.*

10.2 Energy levels and the arrangement of electrons around the nucleus

(a) Energy Levels

The electrons in an atom are not located at the same place. They revolve in orbits that are located at definite distances from the nucleus. Each of the orbits is called energy level. And each energy level is associated with a certain amount of energy.

The energy levels are represented by circles that are numbered from 1, 2, 3, 4, 5, 6 and 7 from the nucleus. The first energy level is the one closest to the nucleus and energy level 7 is the furthest from the nucleus.

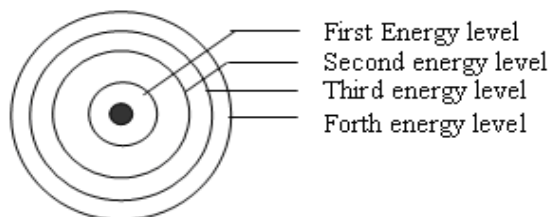


Figure 10.2 Organization of energy levels in an atom.

Points to note

- (i) The energy in which an electron is located depends on the amount of energy it possesses.
- (ii) Electrons with low energy are located in the energy levels that are close to the nucleus.
- (iii) Electrons with high energy levels are located in the energy levels further away from the nucleus.
- (iv) Electrons possessing the same amount of energy occupy the same energy levels.

(b) Arrangement of electrons in the energy levels

The arrangement of electrons in various energy levels (shells) in an atom is called **electronic configuration**. Electrons always occupy the lowest available energy level, i.e. the unfilled energy level is filled up, and additional electrons occupy the next higher energy level.

The maximum number of electrons that an energy level can hold at any one time is fixed. It is given by the expression:

$$2n^2 \quad \text{Where } n = \text{the number of energy levels.}$$

It is, however, important to note that for the first 20 elements, the third shell or energy level can only hold a maximum of 8 electrons.

The electronic configuration of an element can be represented simply by numbers separated by commas (,) or dots (.) or ratio symbol (:).

The first number denotes the number of electrons in the first energy level; the second number denotes the number of electrons in the second energy level and so on.

For example, the electronic configuration of the most reactive element (the first element in the activity series), Potassium may be written as:

2:8:8:1

2,8,8,1

2.8.8.1

The first type is the most commonly used in most Chemistry texts books.

Table 10.2 shows the electronic configuration of the first 20 elements in the periodic table.

No.	Element	Symbols	Total number of Electrons	Number of electrons in every level.			
				1	2	3	4
1	Hydrogen	H	1	1			
2	Helium	He	2	2			
3	Lithium	Li	3	2	1		
4	Beryllium	Be	4	2	2		
5	Boron	B	5	2	3		
6	Carbon	C	6	2	4		
7	Nitrogen	N	7	2	5		
8	Oxygen	O	8	2	6		
9	Fluorine	F	9	2	7		
10	Neon	Ne	10	2	8		
11	Sodium	Na	11	2	8	1	
12	Magnesium	Mg	12	2	8	2	
13	Aluminium	Al	13	2	8	3	
14	Silicon	Si	14	2	8	4	
15	Phosphorous	P	15	2	8	5	
16	Sulphur	S	16	2	8	6	
17	Chlorine	Cl	17	2	8	7	
18	Argon	Ar	18	2	8	8	
19	Potassium	K	19	2	8	8	1
20	Calcium	Ca	20	2	8	8	2

Table 10.2

10.3 Characteristics of the Atom

Definitions of Atomic number and Mass number of an element

Atoms of any one element are the same but differ from of other elements. This difference is due to the differing number of electrons, protons and neutrons they contain.

(a) Atomic number

Atomic number of an element is the number of protons in the nucleus of an atom.

It is denoted by Z . A neutral atom has an equal number of protons and electrons. Hence the atomic number also refers to the number of electrons in an atom.

(b) Mass number (Atomic mass or nucleon number)

The mass number of an element is the total number of protons and neutrons in the nucleus of an atom.

It is denoted by A , and is calculated from the formula:

$$\begin{aligned}\text{Mass number} &= \text{Number of protons} + \text{Number of neutrons} \\ &= \text{Atomic number} + \text{Number of neutrons}\end{aligned}$$

$$A = Z + n$$

Where n = number of neutrons.

Examples:

1. An atom has mass number 39 and atomic number 20. Calculate the number of neutrons in the nucleus of the atom.

Solution $A = 39,$ $Z = 20,$ $n = ?$

$$\begin{aligned}A &= Z + n \\ 39 &= 20 + n \\ n &= 39 - 20 \\ \mathbf{n} &= \mathbf{19}\end{aligned}$$

2. A neutral atom has mass number 40 and the number of electrons is 20.

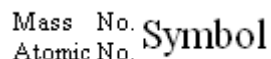
- (a) State the atomic number.
(b) Calculate the number of neutrons.

Solution

- (a) Atomic number, $Z = 20$
(Because $Z = \text{Number of electrons in a neutral atom}$)
- (b) $A = 40,$ $Z = 20,$ $n = ?$
- $$\begin{aligned}A &= Z + n \\ n &= A - Z \\ &= 40 - 20 \\ \mathbf{n} &= \mathbf{20}\end{aligned}$$

10.31 Symbolic representation of an atom of an element

An atom of an element can be represented by its symbol together with its atomic number and mass number. In this case the atomic number (Z), is written as a subscript and a mass number as a superscript on the left of the symbol of the element.



Generally, using symbols of mass number and atomic number, we can represent the atom as:



Where: X = symbol of an element,
 A = mass number and
 Z = atomic number

For example Sodium atom has mass number 23 and atomic number 11. It can be represented as: ${}_{11}^{23}\text{Na}$

The symbolic representation of an atom of an element provides information about the atomic structure of the atom. For example from the example of Phosphorus, ${}_{15}^{31}\text{P}$ provides information about the atomic structure of Phosphorus atom.

Mass number, A, of Phosphorus, = 31 (superscript)
Atomic number, Z, of Phosphorus = 15 (subscript)
Number of electrons = atomic number (15)
Number of neutron = A - Z
 = 31-15

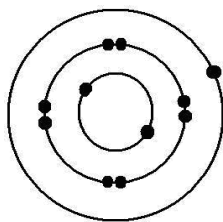
∴ Number of neutron = 16

The information provided can be used to draw the *atomic structure* of an atom.

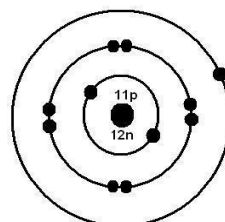
Atomic structure of an atom

The atomic structure of an atom shows the number of protons and neutrons in the nucleus of an atom and the electrons in their energy levels.

Electronic and Atomic structure of Sodium atom ($^{23}_{11}\text{Na}$ and 2:8:1)



a) **Electronic structure**
Shows only electrons in the shells



b) **Atomic structure**
Shows - The nucleus & its contents.
- Electrons in the shells.

10.4 Isotopes

Definition:

Isotopes of an element are atoms which have the same atomic number but different mass number due to the number of neutrons in the nucleus of the atom.

Thus isotopes of an element:

- (i) Have the same number of protons and electrons.
- (ii) Have different number of neutrons.

Examples of elements that have isotopes are:

- Hydrogen, Carbon, Chlorine, e.t.c.

Carbon exists in three isotopes designated *carbon-12*, *carbon-13* and *carbon-14*.

The numbers 12, 13 and 14 refer to the mass numbers of atoms of each isotope.

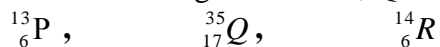
Table 10.3 shows isotopes of the common elements.

Element	Isotopes	Name of isotopes	A	Z	n
Hydrogen	^1_1H	Hydrogen	1	1	0
	^2_1H	Deuterium	2	1	1
	^3_1H	Tritium	3	1	2
Carbon	$^{12}_6\text{C}$	Carbon-12	12	6	6
	$^{13}_6\text{C}$	Carbon-13	13	6	7
	$^{14}_6\text{C}$	Carbon-14	14	6	8
Chlorine	$^{37}_{17}\text{Cl}$	Chlorine-37	37	17	20
	$^{35}_{17}\text{Cl}$	Chlorine -35	35	17	18

Table 10.3

Examples

1. Which of the following elements P, Q and R are isotopes?



Solution: P and R are isotopes of the element.

Reason: They have the same atomic number but different mass number.

2. The table below shows the details of the subatomic particles of the elements W, X, Y and Z. Study it carefully and use it to answer the questions that follow.

Atom	No. of protons	No. of electrons	No. of neutrons
W	11	12
X	17	17	18
Y	18	18	22
Z	17

- (a) Complete the table.
 (b) If the mass number of atom Z is 37, calculate the neutron number of Z.
 (c) (i) State the elements that are isotopes?
 (ii) Give reason for your answer.

Solution

(a)

Atom	No. of protons	No. of electrons	No. of neutrons
W	<u>11</u>	11	12
X	17	17	18
Y	18	18	22
Z	17	<u>17</u>	<u>20</u>

- (b) $A = 37$, $Z = 17$, $n = ?$

$$A = Z + n$$

$$37 = 17 + n$$

$$n = 37 - 17$$

$$\mathbf{n = 20}$$

- (c) (i) X and Z
 (ii) X and Z have the same atomic number but different number of neutrons.

10.41 (a) Relative Atomic mass (RAM)

Relative Atomic mass of an element is the average mass of its atoms relative to $\frac{1}{12}$ the mass of an atom of carbon-12.

$$\text{Relative atomic mass} = \frac{\text{Average mass of atom of an element}}{\frac{1}{12} \text{ of the mass of 1 atom of carbon - 12}}$$

Points to note:

- ❖ *The mass of atom is very small. It is therefore difficult to measure the mass of a single atom.*
- ❖ *Hydrogen atom, being the lightest, was originally chosen as the standard atom and given an atomic mass of 1.*
- ❖ *The masses of other atoms were measured relative to that of hydrogen, and are called RAM.*
- ❖ *The RAM is measured using an instrument called mass spectrometer.*
- ❖ *Today, relative atomic mass of elements is based on the carbon-12 scale instead of hydrogen.*
- ❖ *On this scale, a carbon atom with 6 protons and 6 neutrons weighs exactly 12 atomic mass unit.*

(b) Calculating the Relative Atomic Mass of isotopes from relative abundance of the isotopes of an element

The Relative Atomic Mass of an element may be calculated from given percentages or ratio of the relative abundance of the isotopes.

Calculation of Relative Atomic Mass from given ratio of isotropic abundance

Example 1

A sample of chlorine is a mixture of two isotopes, ^{35}Cl and ^{37}Cl in the ratio 3:1. Calculate the RAM of chlorine atom.

$$\begin{aligned} \text{Solution:} \quad \text{Ratio} &= ^{35}\text{Cl} : ^{37}\text{Cl} \\ &= 3 : 1 \\ \\ \text{Total} &= 3 + 1 \\ &= 4 \end{aligned}$$

$$\begin{aligned}
 \text{Mass contributed by } ^{35}\text{Cl} &= \frac{\text{Ratio of isotopes}}{\text{Total ratio}} \\
 &= \frac{3}{4} \times 35 \\
 &= \mathbf{26.25}
 \end{aligned}$$

$$\begin{aligned}
 \text{Mass contributed by } ^{37}\text{Cl} &= \frac{1}{4} \times 37 \\
 &= \mathbf{9.25}
 \end{aligned}$$

$$\begin{aligned}
 \text{RAM of chlorine} &= 26.25 + 9.25 \\
 &= \mathbf{35.5}
 \end{aligned}$$

NB: - RAM of Chlorine is either 35 or 37. It is the average mass of a mixture of atoms ^{35}Cl and ^{37}Cl
 - 35.5 is not the mass number of Chlorine.

Calculating Relative Atomic Mass from given percentages of isotropic abundance

Example 2

In a sample of chlorine gas, the relative abundance of chlorine-35 atoms is 75% and that of chlorine-37 is 25%. Calculate the relative atomic mass of chlorine.

Solution:

$$\begin{aligned}
 \text{Percentage of } ^{35}\text{Cl} &= 75 \\
 \text{Percentage of } ^{37}\text{Cl} &= 25 \\
 \text{Total percentage} &= 75 + 25 \\
 &= 100\%
 \end{aligned}$$

$$\begin{aligned}
 \text{Mass contributed by } ^{35}\text{Cl} &= \frac{\text{Percentage of } ^{35}\text{Cl}}{\text{Total percentage}} \times \text{Mass number } ^{35}\text{Cl} \\
 &= \frac{75}{100} \times 35 \\
 &= \mathbf{26.25}
 \end{aligned}$$

$$\begin{aligned}
 \text{Mass contributed by } ^{37}\text{Cl} &= \frac{25}{100} \times 37 \\
 &= \mathbf{9.25}
 \end{aligned}$$

$$\begin{aligned}
 \text{RAM of chlorine} &= 26.25 + 9.25 \\
 &= \mathbf{35.5}
 \end{aligned}$$

Example 3

Neon gas exists naturally as three isotopes, neon-20, neon-21 and neon-22. In a sample of neon gas, the relative abundance of the three isotopes is neon-20, (90.92%), neon-21, (0.26%, and neon-22, (8.82%). Calculate the relative atomic mass of neon.

Solution:	Percentage of neon-20	= 90.92
	Percentage of neon-21	= 0.26
	Percentage of neon-22	= 8.82
	Total percentage	= 100%

$$\begin{aligned}\text{Mass contributed by neon-20} &= \frac{90.92}{100} \times 20 \\ &= \mathbf{18.184}\end{aligned}$$

$$\begin{aligned}\text{Mass contributed by neon-21} &= \frac{0.26}{100} \times 21 \\ &= 0.055\end{aligned}$$

$$\begin{aligned}\text{Mass contributed by neon-22} &= \frac{8.82}{100} \times 22 \\ &= \mathbf{1.940}\end{aligned}$$

$$\begin{aligned}\text{RAM of neon} &= 18.184 + 0.055 + 1.940 \\ &= \mathbf{20.18}\end{aligned}$$

Self-Check 10.1

1. Element M belongs to group III of the Periodic Table. The most likely formula of its oxide is
A. M_2O .
B. MO_3 .
C. M_3O_2 .
D. M_2O_3 .
2. Which one of the elements with the following atomic numbers reacts most vigorously with water?
A. 2:8:1
B. 2:8:2
C. 2:8:8:1
D. 2:8:8:2
3. Isotopes of an element have got
A. same number of protons and neutrons.
B. same number of electrons and neutrons.
C. different number of electrons and protons.
D. same number of electrons and protons.
4. Which one of the elements whose atomic numbers are given below gives an acidic oxide?
A. 2:8:1
B. 2:8:3
C. 2:8:6
D. 2:8:8:2
5. In which group of the periodic table is the element ${}^{24}_{12}M$ found?
A. I.
B. II.
C. III.
D. IV.
6. Which one of the following is **wrong** about an element with atomic number 7?
A. is a non-metal.
B. has 5 electrons in its outer valence shell.
C. forms acidic oxides.
D. reacts by loss of electrons.
7. A compound has the formula $M_3(PO_4)_2$. To which group of the Periodic Table does the element M belong?
A. II.
B. III.
C. IV.
D. V.
8. The element M belongs to group IV in the periodic table. The formula of the oxide of M is
A. M_2O_3 .
B. M_2O_5 .
C. M_2O .
D. MO_2 .
9. Which one of the metal atoms whose electronic structures given below forms a nitrate of the type $M(NO_3)_2$?
A. 2:8:0
B. 2:8:1
C. 2:8:2
D. 2:8:3

10. Which one of the following atomic numbers will indicate the least reactive element?
 A. 2:8:1
 B. 2:8:2
 C. 2:8:7
 D. 2:8:8
11. Which one of the following is **wrong** about calcium ion?
 A. 20 protons.
 B. 18 electrons.
 C. 20 neutrons.
 D. 40 neutrons.

The table below shows the number of protons, electrons and neutrons of atom R, S, T, U and V. Use the information in the table to answer questions 12 to 14.

Atom	No. of protons	No. of electrons	No. of neutrons
R	11	11	12
S	17	17	18
T	18	18	22
U	19	19	20
V	17	17	20

12. Which one of the above atoms are isotopes?
 A. R and S.
 B. S and T.
 C. U and V.
 D. S and V.
13. Which atoms form positive ions of the same charge?
 A. R and S.
 B. R and U.
 C. U and V.
 D. S and V.
14. Which of the above atoms belong to an inert element?
 A. R.
 B. S.
 C. T.
 D. U.
15. If atom R belongs to element R and atom V belonged to element V, what would be the formula of the compound formed between element R and V?
 A. RV.
 B. R₂V.
 C. RV₂.
 D. R₂V₃.
16. When elements X and Y are heated together they form a compound with the formula X₃Y₂. Elements X and Y have the following electronic structures respectively
 A. 2:8:1 and 2:5
 B. 2:8:2 and 2:4
 C. 2:8:1 and 2:6
 D. 2:8:2 and 2:5
17. Chlorine atom has electronic configuration of 2, 8, 7. The electronic configuration of a chlorine ion (Cl⁻) is
 A. 2:8:7
 B. 2:8:8
 C. 2:8:6
 D. 2:8:5

18. The atomic number of aluminium ion (Al^{3+}) is 13. The electronic configuration of the aluminium (Al^{3+}) is
A. 2:8 **B.** 2:8:3
C. 2:8:6 **D.** 2:8:8
19. The number of neutrons in the atom of an element Y represented by $^{213}_{84}\text{Y}$ is
A. 84 **B.** 129
C. 213 **D.** 297
20. What is not true about the atoms $^{12}_6\text{X}$ and $^{14}_6\text{Y}$?
A. They have the same number of protons.
B. They have the same number of electrons.
C. They are atoms of the same element.
D. They have the same number of neutron.
21. The number of particles in the nuclei of atoms Q, R, S and T are shown in the table below.
- | Atom | Number of Particles | |
|------|---------------------|----------|
| | Protons + Neutrons | Neutrons |
| Q | 40 | 20 |
| R | 40 | 22 |
| S | 45 | 24 |
| T | 45 | 25 |
- Which of the atoms are isotopes?
A. Q and R. **B.** Q and T.
C. R and S. **D.** S and T.
22. Element M forms the ion M^{3+} . The atomic number of M is
A. 12 **B.** 13
C. 15 **D.** 19
23. The electronic configurations of elements W, X, Y and Z are: W = 2:8:2; X = 2:8:6; Y = 2:8:7; Z = 2:8:3.
 Which of the following pairs are non-metals?
A. X and Y. **B.** Y and Z
C. W and Z. **D.** W and X.
24. The ion formed by the element X of atomic number 13 is
A. X^{3+} . **B.** X^{2+} .
C. X^{2-} . **D.** X^{3-} .
25. The atomic number of an element is
A. the number of electrons and protons. **B.** the number of protons and neutrons.
C. the number of neutrons. **D.** the number of protons.

SECTION B

26. An atom of element X contains 15 electrons and 16 neutrons.
- (a)
 - (i) State the mass number of X.
 - (ii) Write the electronic structure of X.
 - (b)
 - (i) Write the formula of a chloride of X.
 - (ii) State the type of bond that exists in the chloride of X.
 - (c) Suggest how an aqueous solution of the oxide of X would affect litmus paper.
27. The table of protons, neutrons and electrons in particles W, X, Y and Z are shown in the table below.

Particle	No. of Protons	No. of Electrons	No. of Neutrons
W	6	6	6
X	9	10	10
Y	12	10	10
Z	19	19	20

- (a) Which one of the particles is
 - (i) a cation?
 - (ii) an anion?
 - (b) Write the electronic configuration of Z.
 - (c)
 - (i) State the valency of W.
 - (ii) Give a reason for your answer.
28. The atomic number of element Q is 13.
- (a) Write the electronic configuration of an atom of Q.
 - (b) To which group in the Periodic Table does Q belong?
 - (c) State whether Q would conduct electricity or not.
 - (d)
 - (i) Write the formula of the oxide of Q.
 - (ii) State the type of bonding in the oxide of Q.

- 29.** The number of electrons' protons and neutrons in atoms A, B, C and D are shown in the table below.

Atom	Electrons	Protons	Neutrons
A	8	8	8
B	16	16	16
C	13	13	14
D	x	3	4

- (a) Determine: (i) the value of x.
(ii) the approximate relative atomic mass of C.
- (b) Write the electronic configurations of the following atoms and ions:
(i) A
(ii) A^{2-}
(iii) C
(iv) C^{3+}
- (c) State two atoms that are of elements in the same group of the Periodic Table.
- 30.** Elements X and Y with atomic numbers 12 and 8 respectively react to form a compound W.
- (a) Write the electronic configuration of
(i) X.
(ii) Y.
- (b) State whether W is:
(i) a gas, liquid or solid at room temperature.
(ii) Covalent or ionic.

THE PERIODIC TABLE

A modern Periodic Table consists of horizontal rows called *periods* and vertical columns called *groups*.

10.5 Characteristics of the Periodic Table

In the Periodic Table, there are eight groups and seven periods. Elements are arranged in order of atomic numbers. (I.e. the number of protons in the nucleus).

(a) Periods

Period I - contains only hydrogen and Helium. Hydrogen is placed in both groups I and VII. This is because some of its properties resemble those of the group I (alkali metals) and group VII (halogens).

Period II and III (*short periods*) - contain 8 elements each.

Periods IV and V (*long periods*) - each contains 18 elements. Eight of these correspond to the eight elements in the short periods. The other 10 are called *transition elements*, which show some properties to the elements before them and also some properties similar to those after them.

Period VI and VII - contain 32 and 17 elements respectively.

(b) Groups

Elements of the same group have the same number of electrons in the outer shell and they have similar properties.

Group I

There are 6 metals in this group. They include Lithium (Li), Sodium (Na), Rubidium (Rb), Cesium (Cs) and Francium (Fr). **Group 1** metals are called *alkali metals*.

Group II

Group II contains 6 metals namely, Beryllium (Be), magnesium (Mg), strontium (Sr), barium (Ba), and radium (Ra).

The group II metals are called *alkaline earth metals*.

Group VII (Halogens family)

The elements in this group are nonmetals. They include fluorine, (F), chlorine (Cl), bromine (Br), iodine (I) and astatine (At).

Group VIII (or group “O”); noble gases;

The elements in this group are all gaseous elements. These are helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe) and radon (Rn).

Part of the Periodic Table showing the first twenty elements

Period	I	II											III	IV	V	VI	VII	VIII
1	H																	He
2	Li	Be	Transition elements										B	C	N	O	F	Ne
3	Na	Mg											Al	Si	P	S	Cl	Ar
4	K	Ca																
5																		
6																		
7																		

Periodicity of properties of the elements

(a) Physical properties

The physical properties of elements in the same group vary gradually and regularly.

Group 1 (Alkali metals)

All alkali metals have the following properties.

- They have
- (i) Low
 - density
 - melting point
 - boiling point and
 - (ii) High
 - conductivity of heat and electricity.

Group VII (Halogens)

The elements in this group include: Fluorine (F), Chlorine (Cl), Bromine (Br) and iodine (I).

The physical properties such as melting point, boiling point and density of group VII members rise gradually.

(b) Chemical properties

The chemical properties vary gradually for both different and same group of elements. The chemical properties of elements are determined by the atomic size (atomic radius).

(i) Atomic size across the period (left–right)

The atomic radii of elements decrease across the period (left-right). In each Period each atom has one more electron and one more proton than the preceding atom (i.e. atom before it). The extra electron has little effect on the size of the atom because it is in the same outer orbital. However, the extra proton increases the charge on the nucleus and attracts all the electrons closer to the nucleus thus decreasing the size of the atom.

(ii) Atomic size down the group

For the elements in the same group, the atomic size (radius) increases down the group. In each group, each atom has one more Period (complete filled shell or energy level). The extra shell has shielding effect on the charge on the nucleus and the power of the nucleus to attract the electrons in the outer shell is decreased. This makes it easier for the atom to lose the valency electron(s), hence the atomic size and reactivity of the metals increase down the group.

The Periodic Table is accounted for by the regular arrangement of electrons in atoms of the elements. The Periodicity of electronic configuration leads to Periodicity of chemical properties, which intern depends on the configuration of the outer most orbital. The following examples make it clear.

Alkali metals		Halogens		Noble gases	
Symbol	Electronic Configuration	Symbol	Electronic Configuration	Symbol	Electronic Configuration
				He	2
Li	2 : 1	F	2 : 7	Ne	2 : 8
Na	2 : 8 : 1	Cl	2 : 8 : 7	Ar	2 : 8 : 8
K	2 : 8 : 8 : 1	Br	2 : 8 : 18 : 7	Kr	2 : 8 : 18 : 8

The halogens (Group VII elements)

The halogens are non-metals that occupy the seventh group of the Periodic Table. They are so called because they react with metals to form salts. The elements in this group include the following:

Flourine (F), Chlorine (Cl), Bromine (Br), Iodine (I) and Astatine (At).

Atomic structure of the halogens

The atomic structures of the first two halogens are shown in the table below.

Element	Symbol	Electronic Configuration	Protons	Neutrons
Fluorine	F	2 : 7	9	10
Chlorine	Cl	2 : 8 : 7	17	18

All the halogens have seven electrons in their outmost energy level. Because of this, the halogens have closely related physical and chemical properties.

NB: *The structures of bromine and iodine are more complex and are not discussed here.*

Properties of halogens

(a) Physical properties

Some of the physical properties of halogens are shown in the table below.

Element	Melting point (°C)	Boiling point (°C)	Density (g/cm ³)	Colour	Physical state at room temperature
Fluorine	-220	-188	0.00169	Pale yellow	Gas
Chlorine	-102	-35	0.00321	Yellow- green	Gas
Bromine	-7	59	2.93	Red-brown	Liquid
Iodine	114	184	4.93	Purple- black	Solid

- (i) The melting points and the boiling points increase down the group. This is because, as the size of the molecules increase, the forces of attraction between them get stronger. As a result, more energy is required to separate the molecules.
- (ii) The densities also increase down the group.
- (iii) The colours of the elements deepen as one moves down the group from pale yellow for fluorine to purple-black for iodine.
- (iv) There is also gradually shift in the physical state at room temperature. Both fluorine and chlorine are gases, bromine is a volatile liquid and iodine is a solid that easily sublimates to give a purple vapour.
- (v) All the halogens have strong choking smell.
- (vi) The halogens are non-conductors of electricity.

(b) Chemical Properties of Halogens

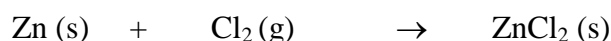
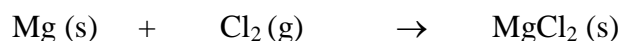
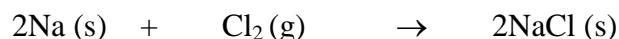
Halogens have one electron less than noble gas structures. Each atom tends to attract an electron to complete its outer octet and therefore halogens are very reactive. The reactivity (ability to attract electrons) decreases down the group. That is, the smaller the atom, the more reactive it is and vice versa. The most reactive halogen is fluorine and the least reactive is iodine.

Reaction with metals

The halogens react with metals to form metallic halides. E.g. Fluorine and chlorine react with metals to form metallic fluoride and chloride respectively.

Reaction of Chlorine with Sodium, magnesium and zinc

Sodium, magnesium and zinc react with chlorine to form sodium chloride, magnesium and chloride.



Noble Gases (Group 0)

The noble gas configurations are very stable and the electrons are not used in chemical reactions.

10.51 Formation of ions

Atoms are electrically neutral. This is because the number of electrons is equal to the number of protons. However, atom can gain or lose electrons forming ions.

When an atom gains electron or electrons the number of negative charges increases above the positive ones resulting in a negative ion.

When an atom loses electron or electrons, the number of negative charges decreases below the positive ones resulting in a positive ion.

Except for helium, all other elements in group eight have eight electrons in their outermost energy level, i.e. octet of electrons. The major characteristics of the elements is that they hardly react with other elements. In other words, they are chemically stable. This is because the possession of eight electrons in the outermost energy level confers stability to an element. Helium contains two electrons in its first and only energy level, i.e. duplet of electrons. This structure functions like the octet of electrons and is chemically stable.

Atoms of other elements lose or gain electrons when they react with other atoms. In losing or gaining electrons, such atoms always aim to attain a stable electronic configuration similar to that of the nearest noble gas. The reactivity of an element is determined by how readily it loses or gains electrons to form an ion. Highly reactive elements lose or gain electrons very readily.

Formation of positive ions (cations)

The formation of positive ions is seen with metallic elements and hydrogen. For example, a sodium atom has one electron in its outermost energy level (2: 8:1). If it loses this electron, the energy level also disappears. The next energy level becomes the outermost (2:8). This way, the stable electronic configuration of eight electrons in the outermost energy level is attained. This is similar to the electronic configuration of neon (2: 8). The resulting particles is no longer called a sodium atom but a sodium ion. It has 10 electrons and 11 protons. Thus it has an excess one positive charge and is therefore positively charged. Its symbol is Na^+ . Such a positively charged ion is called cation.

Formation of negative ions (Anions)

The formation of negative ions is seen with non-metals. For example a chlorine atom has seven electrons in the outermost energy level (2:8:7). It can gain one electron to attain the stable electronic configuration of eight electrons in the outermost energy level. This way, it attains the electronic configuration of argon (2:8:8). It has an excess of one negative charge and is therefore negatively charged. Its symbol is Cl^- . Such a negatively charged ion is called an anion.

10.52 Chemical Families

Alkali Metals (Group I Elements)

These include Lithium, sodium, potassium, Rubidium and Cesium. An alkali metal has one electron in its outer orbital and it readily loses this electron to form an ion of the form, X^+ , which has a stable noble gas configuration. E.g. for the first three elements we have:

Li^+ ion (2),
 Na^+ ion (2:8), and
 K^+ ion (2:8:8).

The reactivity of the alkali metals increases down the group. The larger the atom the more reactive the element is and vice versa. Therefore for the first three, potassium is the most reactive alkali metal followed by sodium. (*See reactivity series on page 96*)

Properties of alkali metals

1. *Potassium and sodium*

(a) **Physical properties**

They are soft with a silvery luster when freshly cut

They have low

- tensile strength.
- melting point.
- density.

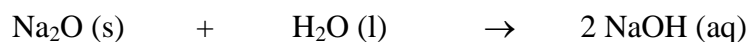
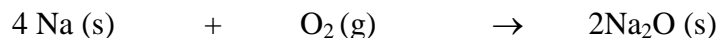
(b) **Chemical properties.**

(i) **Reaction with air**

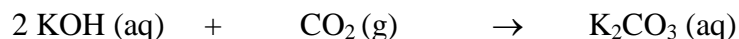
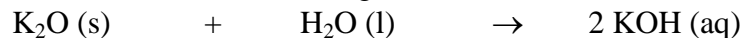
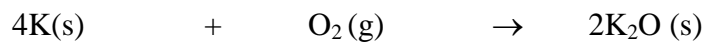
The metals tarnish rapidly because a film of oxide forms. The oxide absorbs water vapour from the atmosphere and dissolves increases it forming hydroxide solution, which slowly absorbs carbon dioxide from air. The final product is sodium carbonate or potassium carbonate solution.

Equations for the Reactions

\Rightarrow ***With Sodium metal***



\Rightarrow ***With Potassium metal***



Due to the high reactivity with air, the metals are stored under liquid paraffin or kerosene.

Sodium burns in air or oxygen with a bright flame, forming mainly yellow, sodium peroxide (Na_2O_2). In insufficient supply of oxygen, sodium oxide Na_2O is formed. Potassium burns with a lilac flame forming deep yellow potassium peroxide (K_2O_2).

(i) Reaction with water

The metals react vigorously with water as discussed on page 108. The heated metal combines readily with Cl_2 and H_2 . Lithium is an alkali metal with the electronic configuration 2:1. Its properties differ substantially from those of sodium and potassium.

2. Lithium

Chemical reaction with:

(a) Air

When a tiny piece of lithium (about 1 mm^3) is heated on a piece of broken porcelain held by tongs, it burns with a scarlet flame (dazzling like Mg), forming, Li_2O .

(b) Water

Lithium reacts slowly with water forming hydrogen and hydroxide

1.53 The alkaline earth metals (Group II metals)

The alkaline earth metals occupy the second group in the periodic table. They are beryllium, magnesium, calcium, strontium, barium and radium. Only first three are the will be dealt with as they belong to the first twenty elements in the periodic table.

Compounds of the alkaline earth metals are found in abundance in rocks of the earth. Their oxides are basic and dissolve in water to form alkaline solutions. It is from these two properties that the alkaline earth metals get their name.

Atomic structures of the alkaline earth metals

The atomic structures of the alkaline earth metals are shown in table below.

Element	Symbol	Electronic configuration	Atomic No. (Number of protons)	Number of neutrons
Beryllium	Be	2 : 2	4	5
Magnesium	Mg	2 : 8 : 2	12	12
Calcium	Ca	2 : 8 : 8 : 2	20	20

Each of the alkaline earth metals has two electrons in its outermost energy level. As a result of this, they have closely related physical and chemical properties.

Atomic radius and ionic radius of the alkaline earth metals

As is the case with the alkali metals, both the atomic radius and the ionic radius of the alkaline earth metals increases down the group. This is because, the number of energy levels increases successively by one as one moves down the group. For each element, the atomic radius is larger than the ionic radius. This is because, when the metals ionize, they lose the outermost energy level. Thus in each case, the neutral atom has one energy level more than its corresponding ion.

Physical properties of alkaline earth metals

- (i) *Hardness*
The alkaline earth metals are relatively hard metals but much softer than iron. Of the first three, beryllium is the hardest followed by magnesium with calcium being the softest.
- (ii) *Conductivity of heat and electricity*
They are good conductors of both heat and electricity.
- (iii) *Melting and boiling points.*
They have relatively higher melting points and boiling points. As compared to alkali metals.
- (iv) *Density*
Alkaline earth metals have higher densities than the densities of alkali metals.
- (v) *Malleability*
They are malleable. I.e. they can be hammered into a sheet.
- (vi) *Ductility*
They are ductile. I.e. they can be drawn into wires.

Chemical properties of alkaline earth metals

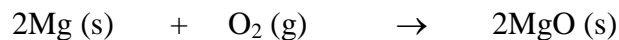
Alkaline earth metals are very reactive metals, although not as the alkali metals. Because of this reactivity, the alkaline earth metals are never found as free elements in nature. An alkaline earth metal has two electrons in its outermost energy level and it readily loses these electrons to form an ion of the form, X^{2+} , which has a stable noble gas configuration. E.g. for the first three elements we have:



(a) Reaction with air

(i) Magnesium

Magnesium ribbon burns with brilliant white flame leaving a white ash containing magnesium oxide and a little magnesium nitride.

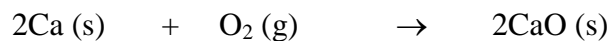


Magnesium oxide is a basic oxide. It dissolves in water to form an alkaline solution Mg(OH)_2 , which turns red litmus paper blue.



(ii) Calcium

Calcium burns with a red flame leaving a white solid called calcium oxide.



The reaction appears to be much slower than that for magnesium.

The solid dissolves sparingly in water to form calcium hydroxide solution, which turns red litmus paper blue.



(b) Reaction with water

See Chemical properties of water, Chapter 8 page 109.

Self-Check 10.2

1. Which one of the following hydroxides dissolves in water to form an alkaline solution?
A. Potassium hydroxide. **B.** Zinc hydroxide.
C. Lead (II) hydroxide. **D.** Copper (II) hydroxide.

2. A compound has the formula $M_2(PO_4)_3$. To which group of the Periodic Table does the element M belong?
A. II. **B.** III. **C.** IV. **D.** V.

3. Which one of the following oxides is soluble in water?
A. Lithium oxide. **B.** Iron (III) oxide.
C. Aluminium oxide. **D.** Copper (II) oxide.

SECTION B

4. An element M has electronic structure 2:8:8:2.
(a) State the group to which this element belongs in the Periodic Table.
(b) Element M was put in warm water.
 (i) State what was observed.
 (ii) Write equation for the reaction that took place.
(c) Name **one** use of element M.

5. Part of the Periodic Table indicating the positions of elements W, X and Z is shown below.

	I	II	III	IV	V	VI	VII	VIII
1								
2	W		X				Z	
3								
4								

- (a)** (i) Write the formula of the oxide of W.
 (ii) The oxide of W was dissolved in water.
 State whether the resultant solution is acidic, neutral or alkaline. Explain your answer.

- (b)** Write the formula of the compound formed between X and Z.

- (c)** Which one of the atoms W, X and Z has the largest atomic radius?

6.

(a)	Element	Atomic radius/nm	Ionic radius/nm
	A	0.133	0.078
	B	0.090	0.120
	C	0.157	0.098

- (i) Two elements are metals. Which are they?
- (ii) Which is the non-metallic element? How do know that?
- (iii) The two metallic elements belong to then same group of the periodic table. Which is more reactive? Which one will have greater atomic number?
- (b) The table below gives some properties relating to elements in period 3 of the periodic table and its oxides.

Element	Na	Mg	Al	Si	P
Formula of oxide	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₂ O ₅
Atomic radius	1.57	1.36	1.25	1.70	1.10

- (i) The atomic radii of the elements in the periodic table decrease from left to right. Explain this.
- (ii) What connection is there between the properties of the oxides and the metallic and non metallic nature of the elements in this period?
- (iii) Which oxides are:
- Acidic
 - Basic
 - Amphoteric?
- (c) Both sulphur and chlorine follow phosphorous in this period. Write the formulae of their oxides.
- (d) Explain in terms of nuclear charge why the K⁺ ion is smaller than the Cl⁻ ion, though they have the same number of electrons.

CHAPTER ELEVEN

CHEMICAL BONDING

Learning objectives

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

1.
 - (a) *State and define the types of bonding.*
 - (b) *State the properties of electrovalent (ionic) and covalent compounds.*
 - (c) *Explain the formation of covalent bonds*
E.g. in - Hydrogen molecule,
- Oxygen molecule,
- Water molecule and
- Carbon dioxide molecule.
 2.
 - (a) *Describe: (i) Simple Molecular structures and Giant atomic structures.*
(ii) Metallic bonding.
 - (b) *State the properties of metallic structures.*
-

11.0 Why do elements Combine?

Elements react/combine in order to attain stability of the noble gas structure of stable octet or duplet) of electrons. On this general principle, elements combine in two main forms of combination or bonding.

Types of Bonding

There are two main types of bonding, namely:

- Electrovalent or Ionic Bond and
- Covalent Bond.

(a) Electrovalent (Ionic) Bonding

An electrovalent (ionic) bond is one due to the transfer of one or more electrons from one atom or group to another atom or group.

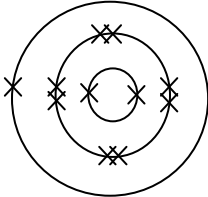

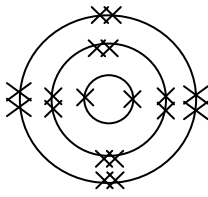
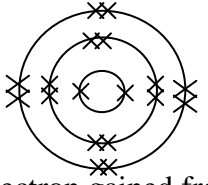
In this type of bonding, an atom of a metallic element and hydrogen or a group loses, from its outermost shell, an electron or a number of electrons equal to its valency. The electron(s) pass over to the outer energy levels of non-metallic atoms with which the metal is combining. By this means, an electron duplet or octet is left behind in the metal and at

the same time created in the non-metal. The atoms of both elements now have acquired the outer electron structure of a noble gas. The metallic particles have a positive charge from the excess proton(s) left in the nucleus, while the non-metal particles are negatively charged from the added electron(s). The two ions attract one another and rearrange themselves to form a solid compounds.

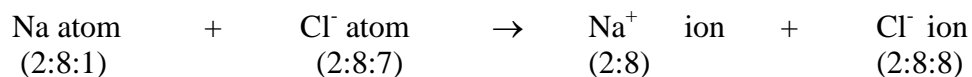
(b) Formation of Ionic Compounds

Examples

1. Sodium Chloride

Atom/Ion	Number of proton	Number of electrons	Electronic Structure
Sodium atom Before combination	+11	-11 (2:8:1)	
Sodium ion After combination	+11	-10 (2:8)	 One electron (valency electron) pass to chlorine leaving Sodium ion Na^+
Chlorine atom Before combination	+17	-17 (2:8:7)	
Chlorine ion After combination	+17	-18 (2:8:8)	 One electron gained from sodium atom increasing the number of electrons to 18 thus forming Chloride ion Cl^-

The two ions have opposite charges and therefore attract each other to form a crystalline solid.



NB:

- ❖ *Sodium atom is bigger than sodium ion. The loss of one electron from the outermost shell led to loss of the shell hence reduction in the atomic radius.*
- ❖ *Chlorine atom and the ion have the same size. The one electron gained joined the seven already in the outermost shell. The one electron has no effect on the atomic radius.*
- ❖ *Both ions (Na^+ and Cl^-) have noble structure.*

2. Magnesium chloride

In the magnesium ion the two excess nuclear protons produce a double positive charge; in each chloride ion, the excess electron produces a single negative charge.

Before combination	Magnesium atom (Mg)		Two chlorine atoms (2Cl)	
	Protons	Electrons	Protons	Electrons
	12	2:8:2	12	2:8:7
	↓		↑	
	These valency electrons pass to the chlorine atoms			
After combination	Magnesium ion (Mg ²⁺)		Two chloride ions (2Cl ⁻)	
	Protons	Electrons	Protons	Electrons
	12	2:8	17	2:8:8
			17	2:8:8

3. Calcium Oxide

In the calcium ion the two excess nuclear protons produce a double positive charge; in oxide two excess electrons produce a double negative charge.

Before combination	Calcium atom (Ca)		Oxygen atom (O)	
	Protons +20	Electrons 2:8:8:2 ↓	Protons 8	Electrons 2:6 ↑
These valency electrons pass to the chlorine atoms				
After combination	Calcium ion (Ca²⁺)		Oxygen ions (O²⁻)	
	Protons +20	Electrons 2:8	Protons +8	Electrons 2:8

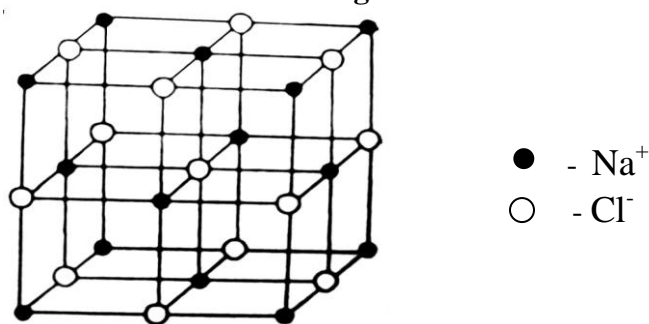
11.1 Giant Ionic Structure

In their solid forms, ionic compounds exist in characteristic three-dimensional crystalline structures called ionic crystal lattice. The lattice is held together tightly by the attractions between the positively and negatively charged ions. An example of giant ionic structure is sodium chloride.

The structure of Sodium chloride

Sodium chloride crystallises as a face centred cube, made up of positively sodium (Na⁺) ions and negatively charged chloride (Cl⁻) ions. In an end face of the cube, a sodium ion occupies the centre and the four corners of the face are also occupied by sodium ions with four chloride ions spaced at equidistant from the central sodium ion. In the next face, the positions of the sodium and chloride ions are reversed. Thus, each sodium ion is surrounded by 6 chloride ions. Similarly each chloride ion is surrounded by *six* sodium ions. The arrangement of the ions is repeated many times resulting in an ionic structure called *giant ionic structure*.

The diagram below shows the arrangement of ions in sodium chloride crystal lattice



Crystals

A crystal is a solid that consists of particles (atoms, molecules or ions) arranged in an orderly and repetitive manner, resulting in a definite external form.

A crystal has a definite geometrical shape with flat sides and sharp edges.

Non-crystalline solids

Some solids do not form crystals and are said to be non-crystalline or amorphous ('of no definite shape'). Their particles are arranged in a random manner.

Properties of Electrovalent (ionic) Compounds

Ionic compounds have the following properties.

1. They are hard crystalline solids and do not vaporize easily.
2. They consist of oppositely charged ions.
3. They do not conduct electricity in solid state. Since the ions are held together by strong forces of attraction. But when in molten or aqueous state, they conduct electricity. This is because the ions are free to move about hence, they are able to carry charge from one point to another.
4. They are soluble in polar solvents e.g. water but are not soluble in non-polar organic solvents like ethanol, benzene, ether e.t.c.
5. They have high melting and boiling points. E.g. the melting point of sodium chloride is 801°C and the boiling point is 1465°C . This is due to the strong attraction between the oppositely charged ions which creates a very strong bond.

11.2 The covalent Bond

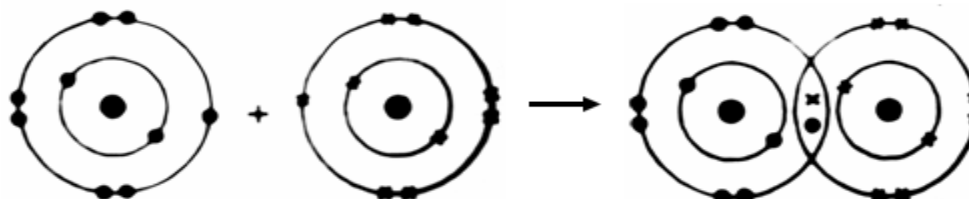
Covalent bonds are formed when atoms of non-metals combine chemically with one another. Like in the ionic bond, covalently bonded atoms also attain the stable gas electronic configuration. In this case, however, the configuration is attained through sharing of electrons. This is illustrated by the following examples.

(a) Single covalent Bonds

Formation of fluorine molecules

Each fluorine atom contains seven electrons in its outer most energy level. It requires one extra electron to form the stable octet. Fluorine molecule is a diatomic molecule i.e. contains two atoms of Fluorine chemically combined. During the formation of a Fluorine molecule, two fluorine atoms come close together. Their outer energy levels overlap. Each atom contributes one electron to form an electron pair which is shared by both atoms. The result is that each atom has eight electron in its outer most energy level. The atoms are held together because the nucleus of each atom attracts the shared electron pair towards it self. This attraction constitutes the **covalent bond**.

Unlike ionic compounds which exist as two or more separate ions, covalent bonding results in the formation of only one particle. The new particle is called a **molecule**. In the case of Fluorine, it is called a Fluorine molecule and has the formula F_2 .

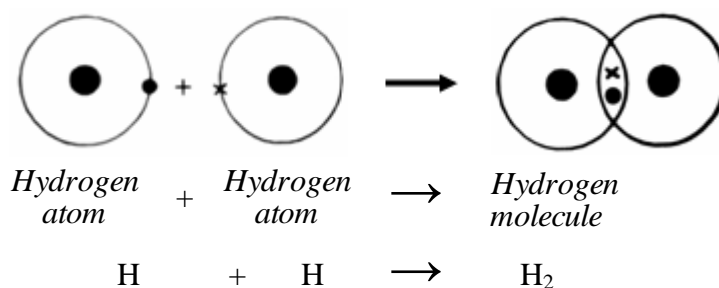


Fluorine atom + Fluorine atom \rightarrow Fluorine molecule.



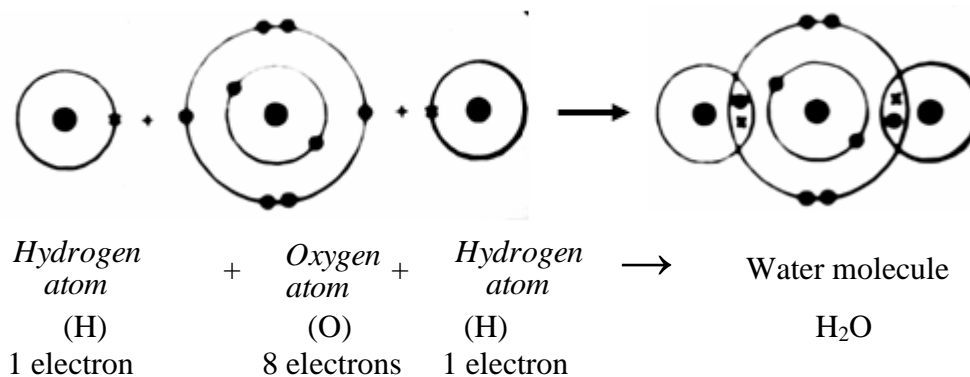
Formation of hydrogen molecule

A hydrogen molecule is a diatomic molecule. It is formed by combination of two atoms. A Hydrogen atom has only one energy level and only one electron. When the two atoms approach one another, their energy levels overlap. Each atom contributes one electron to form a pair of electrons which is shared between them. This way, each atom attains a stable electronic structure similar to that of helium. A covalent bond is thus established between the two atoms resulting in a Hydrogen molecule. (H_2).



Formation of water molecule

A water molecule consists of two Hydrogen atoms and one Oxygen atom bonded covalently. Each hydrogen atom contributes one electron to be shared with the oxygen atom. This way, each Hydrogen atom attains the stable electron duplet and the oxygen atom attains the stable electron octet. Thus a water molecule contains two covalent bonds each bonding the oxygen atom to two separate hydrogen atoms.



N.B *Covalent bonds are often shown using a single line between the symbols of the bonding atoms. Using this method, the structures of some molecules are shown below.*

Molecule	Structure of molecule.
Hydrogen	H – H
Fluorine	F – F
Chlorine	Cl – Cl
Water	$\begin{array}{c} \text{O} \\ / \quad \backslash \\ \text{H} \quad \text{H} \end{array}$

Ammonia	$\begin{array}{c} \text{H}-\text{N}-\text{H} \\ \\ \text{H} \end{array}$
Methane	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array}$

11.3 Multiple covalent Bonds

In some covalent compounds, the combined atoms are formed by two or more covalent bonds. The compounds are said to contain **multiple covalent bonds**.

(a) Formation of double bond

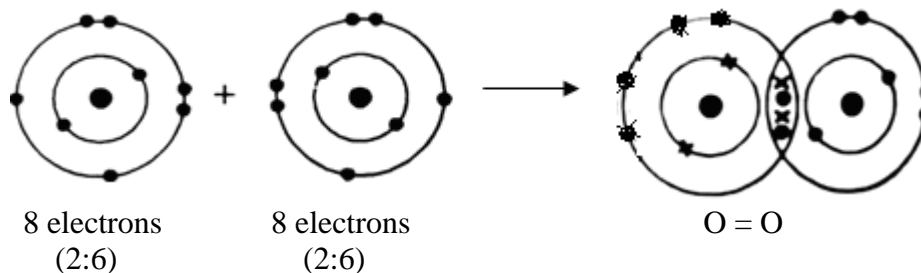
Examples of compounds with double bonds include;

- Oxygen (O_2)
- Carbon dioxide.

Formation of oxygen molecule

Oxygen molecule is a diatomic molecule. Each oxygen atom contains six electrons in its outer most energy level. It therefore requires two electrons to complete the octet. When two oxygen atoms combine, their outer most shells overlap, and each atom contributes a pair of electrons thus, sharing the two pairs of electrons. This way, each atom acquires a stable structure like that of neon. The bond formed is called **a double bond**.

Oxygen atom + Oxygen atom → Oxygen molecule



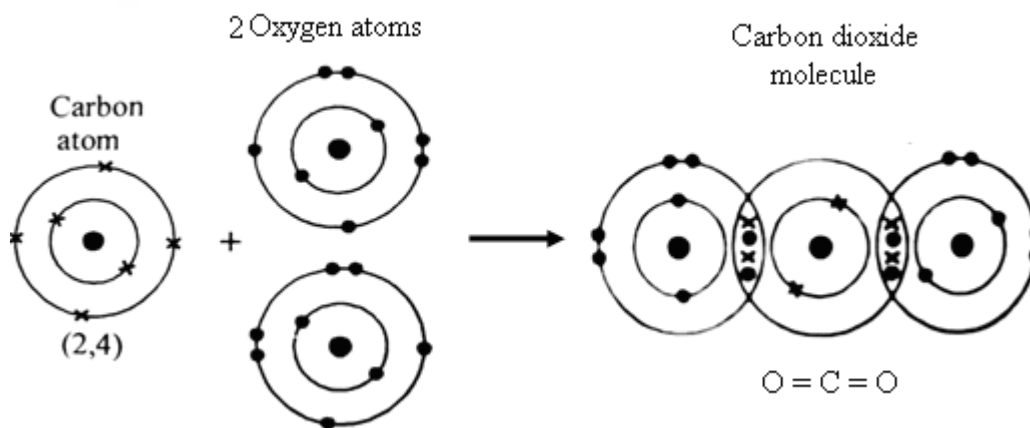
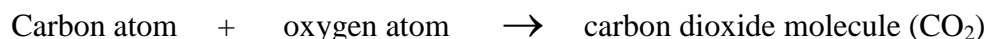
NB: The double bond is often represented as two lines between the symbols of the combining atoms.

(b) Formation of “double” double bond e.g. in CO_2

Example of a compound with double “double covalent bond” is carbon dioxide.

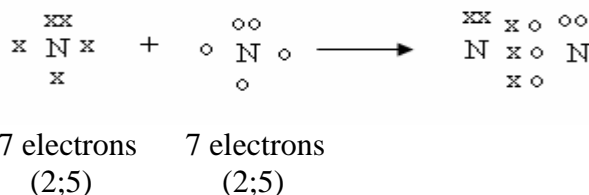
Formation of carbon dioxide molecule (CO₂)

In the carbon dioxide molecule (CO₂), one carbon atom is bonded covalently to two oxygen atoms through double bonds. A carbon atom has four electrons in its outer most energy level. It requires four extra electrons to complete the octet. The carbon atom in carbon dioxide shares two pairs of electrons with each of the oxygen atoms. This way, both the carbon atom and the oxygen atoms acquire an electron octet in their outer most energy levels.



(c) Formation of triple bond e.g. in nitrogen molecule

Nitrogen molecule is a diatomic molecule. The two atoms of nitrogen share three pairs of electrons each atom contributing three electrons to complete the octet, since each nitrogen atom has five electrons in its outer most energy level. The bond formed is called **a triple bond**. It is represented between the nitrogen atoms, that as three lines is \equiv ,



11.4 Bonding and lone pairs of electrons

In many covalent compounds, the participating atoms have two types of electron pairs in their outer most shell. These are;

(i) *Bonding pairs*

This is the shared electron pair or pairs of electrons that are involved in the covalent bond.

(ii) *Lone pairs*

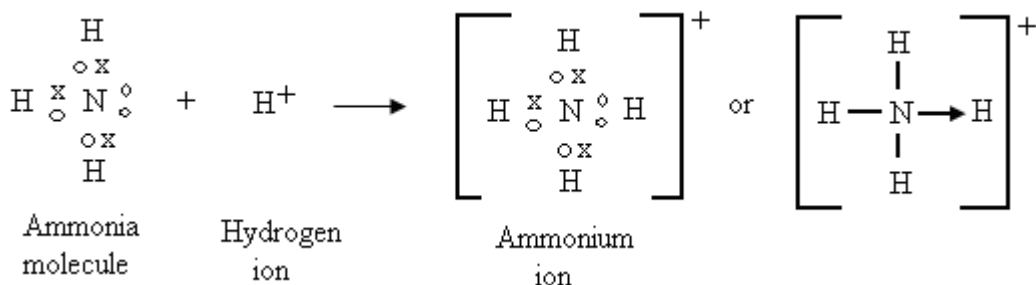
The pair of electrons that are not involved in the covalent bond.

Co-ordinate (dative) bonding

A co-ordinate or Dative bond is another variety of covalent bond where the two shared electrons are both supplied by one of the participating atoms. A co-ordinate bond is formed (in single cases at least) when one of the participating atoms possesses a lone pair of electrons (i.e. a pair not concerned in its existing valency bonds. An example is seen in ammonium ion.

Formation of ammonium ion (NH_4^+)

Ammonium ion is formed when an ammonia molecule reacts with a hydrogen ion (proton) from an acid. All the bonds in ammonia molecule are covalent. The nitrogen atom has a lone pair of electrons, which is not involved in the existing bonds. The ammonia supplies the lone pair to the hydrogen ion (proton). The hydrogen ion which has no electrons, uses the two electrons on the nitrogen atom to complete its duplet. At the same time the nitrogen atom retains its octet since it only shares but does not lose the electron pair to the hydrogen ion. The result is a co-ordinate bond. It is often indicated by an arrow pointing from the donor to the acceptor atom.



Characteristics Properties of Covalent compounds/substances

Covalent compounds/substances have the following characteristics properties.

1. They consist of molecules.
2. They are poor conductors of heat and electricity.
3. They are mostly gases and volatile liquids.
4. They have low melting and boiling points.
5. They are soluble in organic solvents such as ethanol, ether, benzene, carbon tetrachloride.

The Differences between electrovalent (ionic) compounds and covalent compounds

Electrovalent or Ionic Compounds	Covalent Compounds
Crystalline solids except mercury	Mostly liquids and gases
High melting and boiling points	Low melting and boiling points
Consist of oppositely charged ions	Consist of molecules
Conduct electricity in aqueous or molten state	Non-electrolytes. (Do not conduct electricity)
Insoluble in organic solvents e.g. ethanol, benzene, ether etc.	Soluble in organic solvents e.g. ethanol, ether, benzene, but insoluble in polar solvents e.g. water

11.5 Structure of covalent substances

Covalent substances can be classified into two types, namely:

- Simple molecular structures.
- Giant atomic structure

(a) Simple molecular structures

Simple molecular structures consist of small molecules made up of a few atoms. The atoms are joined by strong covalent bonds. However, the separate molecules are only held together by weak intermolecular forces, mainly *van der Waal's* forces. Because of this, only little energy is required to separate molecules into discrete particles. Examples of substances with simple molecular structures are *iodine* and *naphthalene*.

Characteristics of simple molecular structures

- (i) They exist as crystals of low melting points e.g. Naphthalene has a melting point of 81°C while that of iodine is 114°C .
- (ii) They are soft and flaky (easily broken).
- (iii) They are soluble in organic compounds like ethanol, benzene, propanol and ether but are insoluble in polar solvents e.g. water.
- (iv) They are poor conductors of heat and electricity.

(b) Giant atomic structures (Giant molecular or covalent) structures

The giant covalent structures consist of millions of similar or different atoms held together by strong covalent bonds. Examples of giant covalent structures are;

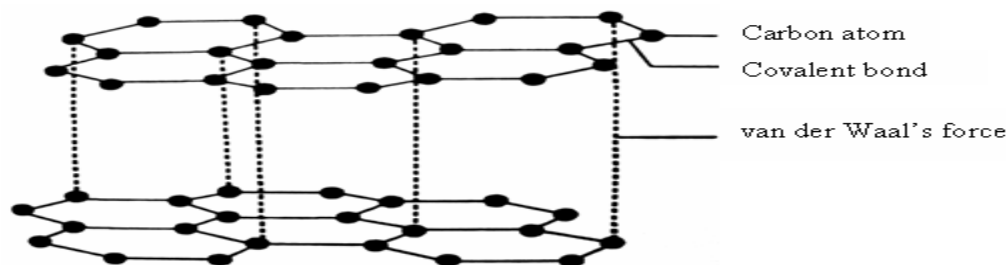
- (i) Graphite,
- (ii) Diamond and
- (iii) Silicon (IV) oxide.

The first two consist of millions of similar atoms (i.e. carbon atoms) only. While the last, consists of different atoms (i.e. Si & O atoms).

(i) Structure of graphite

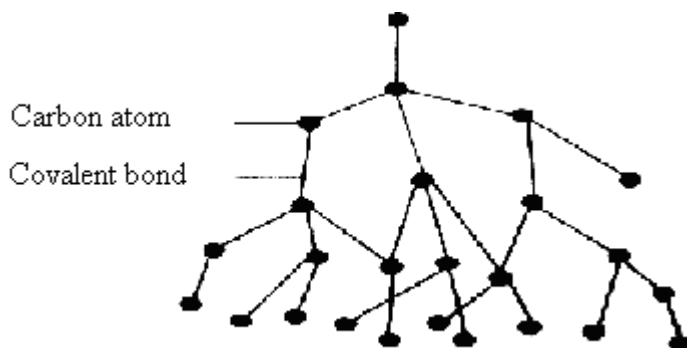
In a graphite structure, each carbon atom is bonded hexagonally to three other carbon atoms by covalent bonds in two dimensions resulting into parallel layers. Each layer is a giant molecule. The layers are held together by weak *van der Waal's* forces. Because of this, the layers are able to slide over each other on application of slight pressure. This helps to explain the softness and greasy feel of graphite. When heated, graphite sublimates at 4200°C .

Of the four valence electrons in each carbon atom, only three are involved in bond formation. The remaining electron is free (delocalized) & is able to carry charge along the layers of the graphite structure. Thus, unlike diamond, graphite is a good conductor of electricity.



(ii) Structure of diamond

In diamond, each carbon atom is covalently and tetrahedrally bonded to three other carbon atoms using strong covalent bonds through out the structure in three dimensions. Thus a diamond crystal is a giant molecule containing millions of carbon atoms. This explains why diamond is very hard.



All the valency electrons of the carbon atoms are involved in bond formation, leaving no free or delocalized electrons. This makes diamond a poor conductor of electricity and heat.

(iii) Structure of silicon (IV) oxide, Si_2O

Silicon (IV) oxide has a structure similar to that of diamond. However, it consists of silicon atoms alternating with oxygen atoms. All the atoms are held together by strong covalent bonds.

Like diamond, silicon (IV) oxide is hard and has a high melting point. It has no delocalized electrons and is therefore a non-conductor of electricity and a poor conductor of heat.

(c) Metallic bonding and the giant metallic structure

In metals, atoms are kept together by a type of bonding called metallic bonding. The metal atoms lose their valence electrons (outer most electrons). The atoms become positively charged ions because they have more protons than electrons. The lost electrons form a "sea" or "cloud" of electrons which surround the positively charged ions and keep them together since they tend to repel each other. The electrons are free to move any where within the structure.

When millions of metal atoms are held together through metallic bonding they form a *giant metallic structure* which is a crystalline solid except mercury which exists in liquid form.

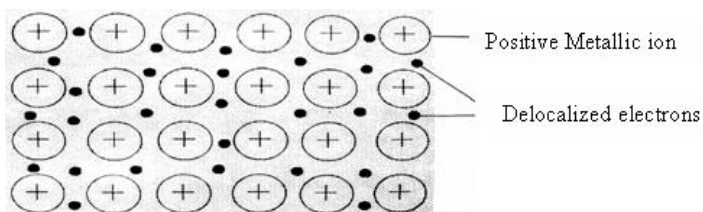


Figure showing metallic bonding

Properties of Metallic structure

1. Conductivity

They are good conductor of heat and electricity.

Explanation

The delocalized electrons are able to carry charge within the metallic structure both in solid and molten states.

Heat and electrical conductivity increases with increase in the number of valence electrons. E.g. Aluminium with valence of three (3) has a higher conductivity than magnesium with valence 2.

2. Hardness

The hardness of a metal depends on the strength of the metallic bond, which in turn is determined by the number of valence electrons in its atoms and the atomic radius.

3. Density

Metals have high densities.

4. Melting and boiling points

The melting and boiling points of metals are determined by the strength of the metallic bond. And the strength of metallic bond also depends on the valence electrons. E.g. Aluminium with 3 electrons in the outermost energy level has strong metallic bond, resulting to high melting and boiling points. In contrast, alkali metals with 1 electron in their outermost energy levels have weaker metallic bonds. Thus they have lower melting and boiling points.

5. Malleability

Metals are malleable. (i.e. can be hammered into sheets).

6. Ductility

They are ductile. (i.e. can be stretched into wires).

The malleability and ductility properties are due to the fact that layers of ions can be moved over one another without breaking the metallic structure.

11.6 Trends of Bonding across a Period

The bonding characteristics change gradually as one moves from left to the right of the period. This can be seen clearly by studying period 3 elements.

The table below shows period three elements and their electronic configurations.

Group	I	II	III	IV	V	VI	VII	VIII
Element	Na	Mg	Al	Si	P	S	Cl	Ar
Electronic Configuration	2:8:1	2:8:2	2:8:3	2:8:4	2:8:5	2:8:6	2:8:7	2:8:8

Period 3 Oxides

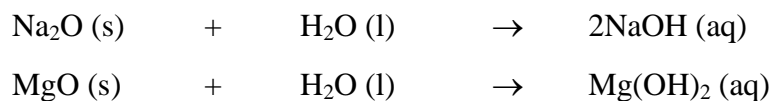
The properties of the oxides of period 3 metals decrease across the period from left to right.

(a) Sodium oxide and Magnesium oxide

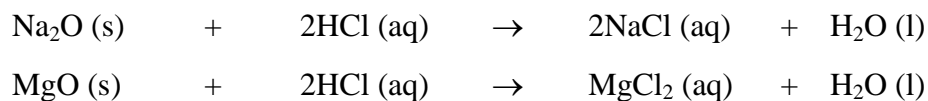
The oxides of sodium and magnesium are strongly ionic compounds with high melting and boiling points. Sodium oxide (Na_2O) exists as Na^+ and O^{2-} ions while magnesium oxide (MgO) exists as Mg^{2+} and O^{2-} ions.

Sodium oxide dissolves readily in water with evolution of heat forming a strong alkaline solution of sodium hydroxide.

Magnesium oxide dissolves slightly in water forming a weakly alkaline solution of magnesium hydroxide.

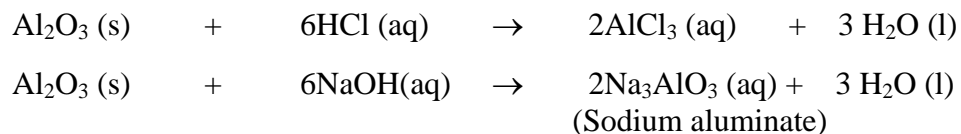


Both oxides are basic and react with acids to form salt and water only.



(b) Aluminium Oxide (Al₂O₃)

Aluminium Oxide is also a strongly ionic compound. It is however, insoluble in water. It is an amphoteric oxide. (i.e. it has both basic and acidic properties).



NB: *In having acidic properties, aluminium oxide behaves like a non-metallic oxide.*

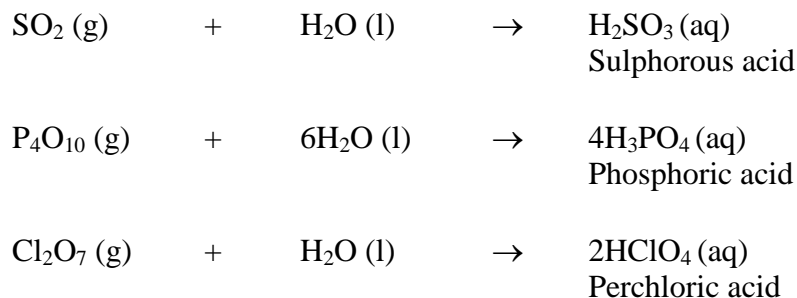
(c) Silicon, phosphorous, sulphur and chloride oxides

(i) Silicon oxide (SiO₂)

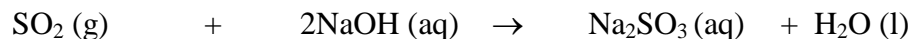
Silicon oxide exists as a giant atomic structure comprising of silicon atoms alternating with oxygen atoms. All the atoms are held together by strong covalent bonds. It is insoluble in water and does not react with the common acids and alkalis under ordinary conditions.

(ii) Phosphorous, Sulphur and Chloride oxides

The oxides of phosphorous, sulphur and chlorine are covalent compounds. They include: Sulphur dioxide (SO₂), Sulphur trioxide (SO₃), Phosphorous (V) oxide (P₄O₁₀) and Chlorine pent oxide (Cl₂O₇). They dissolve in water to form acids.



The oxides are acidic and react with dilute alkalis to form salt and water only.



Some properties of Period 3 oxides

Oxide	Melting point (°C)	Boiling point (°C)	Structure	Nature of oxide
Na ₂ O	1132	1275	Giant ionic	Basic
MgO	2852	3600	Giant ionic	Basic
Al ₂ O ₃	2072	2980	Giant atomic	Amphoteric
SiO ₂	1610	2230	Giant atomic	Amphoteric
P ₂ O ₅	580	300	Simpler molecular	Acidic
SO ₂	-75	-10	Simpler molecular	Acidic
Cl ₂ O ₇	-60	10	Simpler molecular	Acidic

NB: The figures in the table are *not* to be memorised.

11.7 Period 3 chlorides

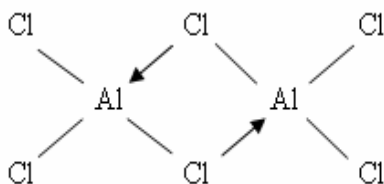
(a) Sodium and Magnesium Chlorides

Sodium and Magnesium Chlorides are strongly ionic compounds. Sodium chloride exists as Na⁺ and Cl⁻ ions. While magnesium chloride exists as Mg²⁺ and Cl⁻ ions. Both sodium and magnesium chlorides have high melting and boiling points. They dissolve in water to form neutral solutions.

Both are electrolytes when in molten and aqueous state.

(b) Aluminium chloride

Aluminium chloride exists as a covalent molecule in the anhydrous state. In this state it exists as a dimer (two molecules combined chemically) with the formula Al₂Cl₆. Its structure is as shown below.



The bonds shown with arrows are coordinate bonds with the electron pairs being donated by chlorine atoms.

Al₂Cl₆ has the following characteristics.

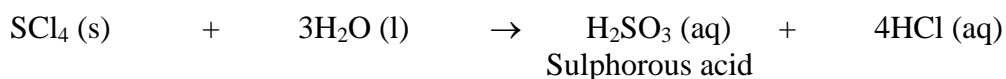
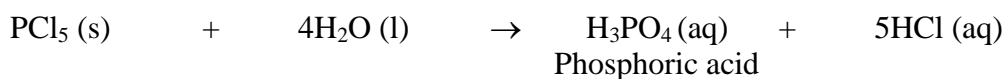
- (i) It is soluble in organic solvents e.g. ethanol.
- (ii) When heated, it sublimes at a temperature of 183 °C.
- (iii) It reacts vigorously with water and it is hydrolysed to give an acidic solution.

Hydrolysis

Hydrolysis, is a type of chemical reaction in which a molecule of water, formula $H-O-H$, reacts with a molecule of a substance AB , in which A and B represent either atoms or groups of atoms. In the reaction the water molecule breaks into the fragments H^+ and OH^- ; and the molecule AB breaks into A^+ and B^- ; the fragments then join to give the final products AOH and HB . This kind of reaction is called a double decomposition.

(c) Silicon, Phosphorous and Sulphur chloride

The chlorides of Phosphorous (PCl_3 and PCl_5) and Sulphur (SCl_4) are covalent compounds. They react with water forming two acids one of which is hydrochloric acid.



Some properties of Period 3 chlorides

Chloride	Melting point (°C)	Boiling point (°C)	Structure	Reaction with water	Nature of solution
NaCl	801	1465	Giant ionic	Dissolves	Neutral
MgCl ₂	714	1437	Giant ionic	Dissolves	Neutral
Al ₂ Cl ₆	183		Covalent	Hydrolysed	Acidic
SiCl ₄	-70	57	Simpler molecular	Hydrolysed	Acidic
PCl ₃	-94	76	Simpler molecular	Hydrolysed	Acidic
SCl ₂	-78	59	Simpler molecular	Hydrolysed	Acidic

Self-Check 11

The electronic configurations of some elements are shown below. Study them carefully and use it to answer questions 1 and 2.

M = 2:6, **N** = 2:5, **P** = 2:8:3, **Z** = 2:8:2

- The pair of elements that can form an ionic compound is
A. Z and M. **B.** Z and P. **C.** Z and N. **D.** M and N.
- Which one of the following pairs of elements are metals?
A. Z and M. **B.** Z and P. **C.** Z and N. **D.** M and N.
- The atomic numbers of elements, P, Q and R are 12, 8 and 7 respectively. Which of the following compounds are ionic?
A. PQ and Q_3R_2 . **B.** PQ and P_3R_2 .
C. Q_3R_2 and RQ. **D.** P_3R_2 and RQ_2 .
- Which one of the substances has a giant ionic structure?
A. Sodium chloride. **B.** Hydrogen chloride. **C.** Graphite. **D.** Diamond.
- How many Cl^- ions surround each Na^+ ion in a sodium chloride crystal?
A. 4 **B.** 6 **C.** 8 **D.** 10
- Two elements have atomic numbers 14 and 8 respectively. The compound formed when they react is most likely to be
A. crystalline with a high melting point.
B. a good conductor of electricity.
C. basic in nature.
D. a molecular gas.
- Which one of the following structures has a giant atomic structure?
A. Sodium chloride. **B.** Hydrogen chloride.
C. Diamond. **D.** Monoclinic sulphur.
- Which one of the following oxides has the most ionic structure?
A. P_2O_5 . **B.** SiO. **C.** Al_2O_3 . **D.** Na_2O .
- Which one of the following pairs of elements can form an ionic bond?
A. Hydrogen and oxygen. **B.** Potassium and bromine.
C. Copper and zinc. **D.** Hydrogen and chlorine.
- Which one of the following statements is **not** correct about electrovalent compounds?
A. They are soluble in water.
B. They are strong electrolytes in solid state
C. They have high melting and boiling points.
D. They are insoluble in organic solvents.

In each of the questions **11** to **15** one or more of the answers given may be correct. Read each question carefully and then indicate the correct answer according to the following:

A. If 1, 2 and 3 only are correct.

B. If 1 and 3 only are correct.

C. If 2 and 4 only are correct.

D. If 4 only is correct.

Instruction Summarised			
A	B	C	D
1, 2, 3	1, 3	2, 4	4
only correct	only correct	only correct	only correct

- 11.** Ionic compounds are generally
1. conductors of electricity when in molten state only.
 2. soluble in water.
 3. soluble in all solvents.
 4. have high melting points.
- 12.** Covalent compounds
1. are usually solids at room temperature.
 2. are formed by sharing of electrons.
 3. conduct electricity when molten.
 4. generally have low melting points.
- 13.** Magnesium is in group II of the Periodic Table. Magnesium oxide has
1. an ionic structure.
 2. a molecular structure.
 3. a high melting point.
 4. a low melting point.
- 14.** The atomic configurations of the atoms of elements P, Q, R, S, T and U are as follows:
- P=2:8:2 Q=2:8:3 R=2:8:4 S=2:8:5 T=2:8:6 U=2:8:7

Which of the pair(s) of elements will form a covalent compound?

1. T and R.
 2. P and U.
 3. T and S.
 4. Q and U.
- 15.** The electronic structure of elements P, Q, R and S are as shown in the table below:

Element	electronic configuration
P	2:8:2
Q	2:8:7
R	2:8:6
S	2:8:1

Which one of the following pairs of elements will combine to form ionic compound(s)?

1. P and Q.
2. P and R.
3. Q and S.
4. Q and P.

SECTION B

16. Part of the Periodic Table is shown below. The letters are not the usual symbols for the elements.

								VIII
I	II		III	IV	V	VI	VII	
							T	
P	Q				S		U	
							W	V

- (a) Which is the least reactive element?
- (b) Which one of the elements T, U and W reacts most vigorously with Q?
- (c) Write the formula of compound formed between Q and S.
- (d) The compound formed between P and W was dissolved in water. State whether the resultant solution was acidic, basic or neutral.
- (e) Which two elements represented in the table can react as reducing agents?
17. (a) An element X is in Group II of the Periodic Table.
- (i) State the type of bond that exists in the chloride of X.
- (ii) Write the formula of the ion formed by X.
- (b) The nitrate of X was strongly heated.
- (i) State what was observed.
- (ii) Write an equation for the reaction.
18. Some elements in Period 3 of the Periodic Table are shown in the table below.

Group	I	II	III	IV	V	VI	VII	O
Element	E	Y			T	X	Q	Z

- (a) Write the formula of the compound formed when
- (i) T reacts with Q.
- (ii) E reacts with X.
- (b) State the type of bonding;
- (i) Between the atoms of Y.
- (ii) When X is reacted with oxygen.
- (iii) Between Y and Q.
19. An atom X of an element, atomic mass 31 contains 15 protons.
- (a) (i) State the number of neutrons in X.
- (ii) Write the electronic configuration of X.
- (b) State the group in the Periodic Table the element belongs.

- (c) (i) Write the formula of a compound that can be formed between X and chlorine.
- (ii) State the bond type in the compound in (c)(i).
- (d) An atom Y contains 17 neutrons and 15 protons. What word is used to describe the relationship between X and Y?

20. Figure 1 shows part of the Periodic Table. The letters used are not the correct symbols of the elements.

I							VIII
	II	III	IV	V	VI	VII	
			P		T	R	
S							Q

- (a) Which of the elements is/are metal(s)?
- (b) Suppose element P reacts with element T,
- (i) Write the formula of the compound formed between P and T.
- (ii) What would be the type of bond formed between P and T?
- (c) (i) Which element in the table is least reactive?
- (ii) Explain your answer in (c)(i).
- (d) (i) Suggest a compound formed between any two elements shown, which would conduct electricity.
- (ii) Give a reason for your answer in (d)(i).

CHAPTER TWELVE

ACIDS, BASES AND SALTS

Learning objectives

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

1. (a) Define the terms acid and base.
(b) Differentiate between:
 - (i) Strong and weak acid.
 - (ii) Strong and weak base.(c) Give examples of strong and weak acids and bases.
 2. (a) Define the term basicity.
(b) State the properties of acids and bases.
 3. (a) Define normal salt and acid salt.
(b) Describe the preparation of soluble and insoluble salts.
 4. (a) Define the term solubility as applied to salts.
(b) Describe an experiment to determine the solubility of a salt at a given temperature.
(c) State the effect of temperature on solubility of solids and gases.
(d) Solve problems involving solubility of salts.
 5. State the effect of heat on nitrates.
 6. (a) Define hard water and soft water.
(b) State:
 - (i) the causes of temporary hardness water and permanent of hardness water.
 - (ii) Methods of softening hard water.
 - (iii) Advantages and Disadvantages of Hard water.
-

12.1 Acids

Definition: *An acid is a compound which when dissolved in water, releases hydrogen ions (H^+) as the only positive ions.*

Types of acids

There are two types of acids, namely:-

- (i) Strong acids and
- (ii) Weak acids

Strong & Weak Acids

(a) Strong acids

A strong acid is one in which there is complete ionization.

No molecules exist in the solution.

The acids that belong to this class are stable and are called *mineral acids*.

The table 13.1 below shows the examples of the strong acids.

Acid	Ionization equation
Hydrochloric acid	$\text{HCl(aq)} \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
Nitric acid	$\text{HNO}_3(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$
Sulphuric acid	$\text{H}_2\text{SO}_4(\text{aq}) \rightarrow 2\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$

Table 13.1

(b) Weak acids

A weak acid is one in which there is incomplete ionization.

In weak acids most of the acid solution exists in form of unionized molecules.

The acids that belong to this class are *unstable*. The examples are shown in the table 13.2 below.

Acid	Ionization equation
Carbonic acid	$\text{H}_2\text{CO}_3(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{HCO}_3^-(\text{aq})$
Sulphurous acid	$\text{H}_2\text{SO}_3(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{HSO}_3^-(\text{aq})$
Ethanoic acid	$\text{CH}_3\text{COOH(aq)} \rightarrow \text{H}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$

Table 13.2

Some everyday acids are:

- | | |
|---------------|--------------------|
| Citric acid | - in lemons |
| Tartaric acid | - in baking powder |
| Carbonic acid | - soda water and |
| Acetic acid | - vinegar |

12.11 Basicity of an acid

Basicity of an acid is the number of hydrogen ions that can be formed from one molecule of the acid.

Acids are further classified into:-

- monobasic acid,
- dibasic acid and
- tri-basic acid.

The table 13.3 below shows the examples of such acids.

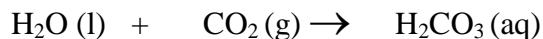
Acid	Ionization equation	Basicity
Hydrochloric acid	$\text{HCl (aq)} \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$	Monobasic
Nitric acid	$\text{HNO}_3(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$	
Sulphuric acid	$\text{H}_2\text{SO}_4(\text{aq}) \rightarrow 2\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$	Dibasic
Phosphoric acid	$\text{H}_3\text{PO}_4(\text{aq}) \rightarrow 3\text{H}^+(\text{aq}) + \text{PO}_4^{3-}(\text{aq})$	Tribasic

Table 13.3

(a) Preparation of acids

Acids can be prepared in the laboratory by the following methods.

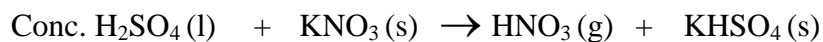
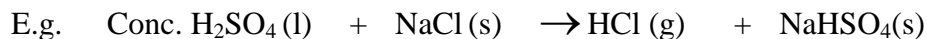
1. By the action of water on an acid anhydride (the acidic oxide of a non-metal).



2. By synthesis (direct combination of elements)



3. By displacing a volatile acid by a less volatile one.



(b) Properties of acids

1. **Taste.** Acids have sour, sharp taste.
 2. **Effect on indicators:** Acids turn blue litmus/solution red.
 3. **Effect on carbonates and hydrogen carbonates:**
Most acids react with carbonates and hydrogen carbonates evolving bubbles of carbon dioxides.
E.g. $2\text{HCl (aq)} + \text{Na}_2\text{CO}_3(\text{s}) \rightarrow 2\text{NaCl (aq)} + \text{H}_2\text{O (l)} + \text{CO}_2(\text{g})$
 $\text{H}_2\text{SO}_4(\text{aq}) + 2\text{KHCO}_3(\text{s}) \rightarrow \text{K}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O (l)} + 2\text{CO}_2(\text{g})$
 4. **Reactions with Oxides and hydroxides.**
Acids react with oxides and hydroxides of metals to form salt and water only.
E.g. $2\text{HCl (aq)} + \text{CuO (aq)} \rightarrow \text{CuCl}_2(\text{aq}) + \text{H}_2\text{O (l)}$
 $\text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaOH (aq)} \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O (l)}$
 5. **Reactions with metals.**
Acids react with many metals to form salt and hydrogen gas.
E.g. $2\text{HCl (aq)} + \text{Zn (s)} \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g})$
- NB:** (i) Nitric acid is an oxidizing agent and does not form hydrogen.
(ii) Magnesium displaces hydrogen from very dilute nitric acid.
(iii) Copper and mercury do not displace hydrogen from any dilute acid.

12.2 Bases and Alkalis

(a) Bases

A base is a compound which contains oxide (O^{2-}) or hydroxide (OH^-) ions and react with an acid to form salt and water only.

(b) Alkalis

An alkali is a compound which, when dissolved in water, forms hydroxide ions (OH^-) as the only negatively charged ions.

Types of alkalis

There are two types of alkalis, namely:

- Strong alkalis and
- Weak alkalis

(i) **Strong alkalis** A strong alkali is one in which there is complete ionization.

No molecules exist in the solution.

Examples are shown in the table 13.4 below.

Chemical name	Common name	Ionization equation
Sodium hydroxide	Caustic soda	$\text{NaOH (aq)} \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$
Potassium hydroxide	Caustic potash	$\text{KOH (aq)} \rightarrow \text{K}^+(\text{aq}) + \text{OH}^-(\text{aq})$
Calcium hydroxide	Slaked lime	$\text{Ca(OH)}_2 (\text{aq}) \rightarrow \text{Ca}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$

Table 13.4

Calcium hydroxide is only slightly soluble but since it completely ionized it is a strong alkali.

(ii) **Weak alkalis:**

A weak alkali is one which does not ionize completely.

It exists as molecules with only a few ions and its ionization is reversible.

Example of weak alkali is ammonium hydroxide. Its ionization is shown in table 13.5 below.

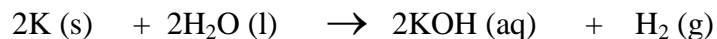
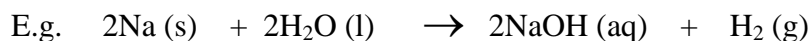
Chemical name	Common name	Ionization equation
Aqueous ammonia	Ammonia solution	$\text{NH}_3 (\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$
Ammonium hydroxide		$\text{NH}_4\text{OH (aq)} \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$

Table 13.5

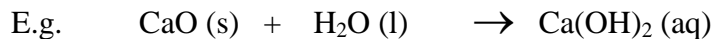
NB: Sodium hydroxide and potassium hydroxide are caustic alkalis (caustic means corrosive or burning) and they react with the skin. Therefore, they should be handled with care.

12.21 (a) Preparations of Alkaline solutions

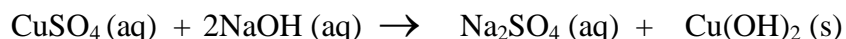
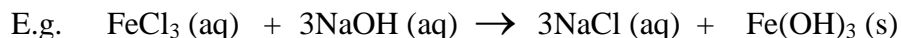
1. Action of water on a metal.



2. Action of water on a basic oxide.



3. Action of alkaline solution on solution of salts.



(b) Properties of Alkalis

1. *Feel and taste.*

Alkalis feel soapy and taste bitter.

2. *Effect on indicators:*

Alkalis turn red litmus solution or litmus paper blue.

3. *Reaction with ammonium salts*

Alkalis (and bases) react with ammonium salts to form ammonia, which has a choking smell and turns litmus blue.

4. *Reaction with acids.*

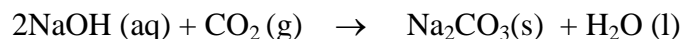
Alkalis neutralize acids to form a salt and water.

5. *Reaction with salts.*

Alkalis precipitate many insoluble hydroxides from solutions of salts.

6. *Reaction with air.*

Sodium hydroxide is deliquescent. It quickly becomes wet and soon forms a solution, which then absorbs carbon dioxide from air and forms crystalline sodium carbonate.



Self-Check 12.1

- Which one of the following substances will dissolve in water to give a solution that would turn red litmus paper blue?
A. Sodium chloride. B. Sodium hydroxide.
C. Sodium sulphate. D. Sodium nitrate.
- The following acids are completely ionised in water **except**
A. hydrochloric acid. B. nitric acid.
C. sulphuric acid. D. carbonic acid.
- The oxide that dissolves **most readily** in water is
A. sodium oxide. B. calcium oxide.
C. lead (II) oxide. D. copper (II) oxide.
- Which one of the following gases dissolves in water to form an acid solution?
A. Nitrogen monoxide. B. Nitrogen.
C. Carbon dioxide. D. Ethene.
- An acid accidentally spilled on a pupil's clothes is best neutralised by a solution of pH
A. 1 B. 4 C. 5 D. 8
- Which one of the following aqueous solutions will produce hydrogen gas with a magnesium ribbon? That of
A. pH=14 B. pH=7 C. pH=11 D. pH=2
- Most metals react with dilute mineral acids to form
A. hydrogen gas only. B. the salt of the metal and water.
C. the salt of the metal only. D. the salt of the metal and hydrogen gas.
- An acid X_2SO_4 ionises as $X_2SO_4(aq) \rightarrow 2H^+_{(aq)} + SO_4^{2-}(aq)$. The basicity of the acid is
A. 1 B. 2 C. 3 D. 4
- Which one of the following acids will react with calcium carbonate to produce the least volume of carbon dioxide gas
A. Dilute sulphuric acid. B. Dilute nitric acid.
C. Dilute hydrochloric acid. D. Dilute ethanoic acid.
- Which one of the following dissolves in water to give a solution with a pH less than 7?
A. $(NH_4)_2SO_4$. B. KCl.
C. Na_2CO_3 . D. CH_3CO_2Na .

12.3 Salts

Definition: *A salt is a compound containing positive metallic ion(s) and negative ion(s) derived from an acid (acid radical).*

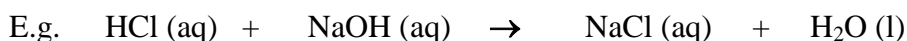
Or *a salt is a compound formed when an acid reacts with a base.*

Or *a salt is a substance formed when either all or part of the ionizable hydrogen atom(s) of an acid is replaced by a metallic ion or ammonium ion.*

Formation of salts

Salts are formed by a process called *neutralization*.

Neutralization is the reaction between an acid and a base to form salt and water only.



Types of Salts

Salts are classified into:-

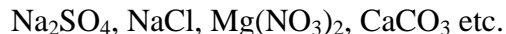
- (i) Normal salts
- (ii) Acid salts
- (iii) Basic salts

(a) Normal Salts

A normal salt is one where all the hydrogen ions in the acid have been replaced by the a metallic ion.

NB: *In normal salts the acid radical (negative ions) are not capable of further ionization to yield H^+ ions.*

Examples of normal salts include:



(b) Acid Salts

An acid salt is one in which not all the hydrogen ions in the acid have been replaced by metallic ions.

NB: *The acid radicals are capable of further ionization to yield H^+ ions.*

Examples of acid salts include:

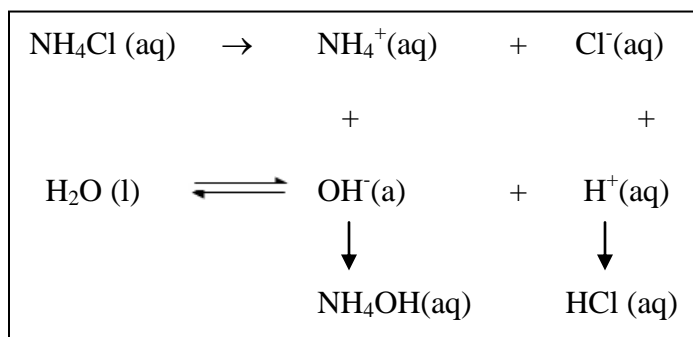
- NaHSO_4 ,
- NaHCO_3 ,
- $\text{Mg}(\text{HCO}_3)_2$ etc.

Acidic salts

Acidic salts are salts formed from weak bases. When they dissolve in water, the solution formed is acidic.

E.g. - Ammonium salts: - NH_4Cl ,
- $(\text{NH}_4)_2\text{SO}_4$ and
- NH_4NO_3

In water, the salts ionize to form ammonium ions and their respective acid radicals. The ammonium ion combines with the hydroxide ions from water to form ammonium hydroxide, a weak base. While the acid radicals combine with the hydrogen ions from water to form strong mineral acids. The basic effect of the weak base is over shadowed by the strong acid formed. Thus, the resulting solution is acidic solution. See the illustration below.



(c) Basic Salts

A basic salt is one in which O^{2-} and OH^- ions are retained together with metallic ions and the negative ions of acids.

Examples:

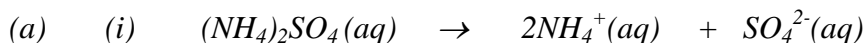
- | | |
|----------------------------|--|
| - Basic zinc chloride | - $\text{Zn}^{2+}\text{OH}^-\text{Cl}^-$ |
| - Basic magnesium chloride | - $\text{Mg}^{2+}\text{OH}^-\text{Cl}^-$ |
| - Basic lead carbonate | - $3\text{Pb}^{2+}.2\text{OH}^-.2\text{CO}_3^{2-}$ |

Note that the charges balance.

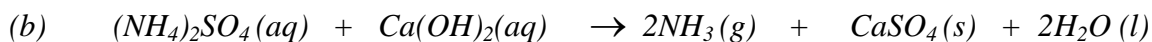
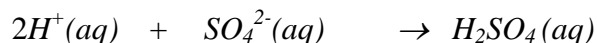
Example

1. (a) Ammonium sulphate dissolves in water to form an acidic solution.
 - (i) Write the equation for the reaction.
 - (ii) Explain why the solution is acidic.
- (b) Write an equation for the reaction that takes place when ammonium sulphate is heated with calcium hydroxide.
- (c) State two tests for the gas in (b) above.

Solution



(ii) The hydrogen ions from water combine with the sulphate ions from the salt to form sulphuric acid.



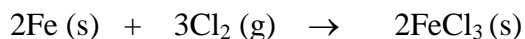
(c) It turns damp or moist red litmus paper blue.
It forms white dense fumes with hydrogen chloride gas.

12.4 Preparation of Salts

There are three major methods of preparing salts. They include:

- (i) Crystallization,
- (ii) Precipitation (double decomposition) and
- (iii) Direct combination.

E.g. FeCl_3 - from *Iron* and *Chlorine*.



The first two methods depend on the solubility of the salt in water. It is therefore important for a student to master the solubility of the common salts in water. The solubility of the common salts is shown below.

Table 13.6 showing the solubility of the common salts

Soluble salts	Insoluble salts
1. All group I/ammonium salts	
2. All nitrates	
3. All chlorides except	Lead (II) chloride and silver chloride. NB: <i>Lead (II) chloride is soluble in hot water.</i>
4. All sulphates except	Lead (II) sulphate, Barium sulphate and Calcium sulphate. NB: <i>Calcium sulphate is only slightly soluble.</i>
5. All group I and Ammonium Carbonates	All other carbonates

Table 13.6

12.41 Methods of preparation of salts

(a) Soluble salts

Soluble salts are prepared by using *crystallization method*.

The process involves heating the solution of the soluble salt to a crystallization point and cooling the concentrated solution to obtain the crystals. Cooling allows the crystals of the salt to grow.

The solution of the soluble salt is obtained from the following reactions:

- (i) Action of acid on a soluble base.
E.g. Preparation of NaCl from NaOH and HCl
- (ii) Action of acid on an insoluble base.
E.g. Preparation of CuSO₄ from CuO and H₂SO₄
- (iii) Action of acid on a metallic carbonate.
E.g. Preparation of ZnSO₄ from ZnCO₃ and H₂SO₄
- (iv) Action of acid on a metal.
E.g. Preparation of ZnSO₄ from Zn and H₂SO₄

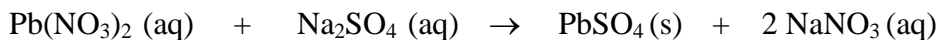
(b) Insoluble salts

The method of preparing insoluble salts is called *double decomposition*. Double decomposition method involves *precipitation* of the salt.

In this method, solutions of two soluble salts are mixed and the metallic ions exchange their acid radicals. The result is that the required insoluble salt is precipitated and the other salt remains in aqueous state.

E.g. Preparation of PbSO₄ (insoluble) from Pb(NO₃)₂ and Na₂SO₄ (both soluble).

Pb(NO₃)₂ - provides the Pb²⁺ ions and
Na₂SO₄ - provides the SO₄²⁻ ions.



12.42 (a) General steps of preparing soluble salts

Preparation of soluble salt from insoluble base (oxide), metallic carbonate, or metal

Procedure:

- (i) Add the metallic oxide, metallic carbonate, or metal to an appropriate acid (acid that provides the required acid radical) till the solid is in excess, warm if necessary.
- (ii) Filter off the excess solid.
- (iii) Saturate the filtrate by evaporation and cool the solution to crystallize.
- (iv) Filter and wash the crystals with distilled water.
- (v) Dry the crystals using filter paper or sunlight.

(b) General steps of preparing insoluble salts

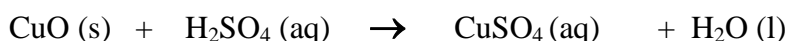
- (i) Mix a solution of soluble salt containing the metallic ion of the required salt with any soluble salt containing the appropriate acid radical. Precipitate forms.
- (ii) Filter the precipitates.
- (iii) Wash the precipitates with distilled excess water to dissolve any soluble impurities.
- (iv) Dry the precipitates using filter paper or sunlight.

12.43 Examples of preparation of soluble salts

1. Describe how you can prepare pure crystals of Copper (II) sulphate from Copper (II) oxide and dilute sulphuric acid

Procedure

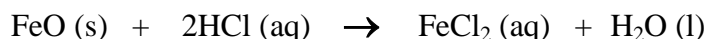
- (i) Add copper (II) oxide into a known volume of dilute sulphuric acid in a beaker a little at a time while stirring with a glass rod until no more of the solid dissolves. (This shows that all the acid is neutralized).



- (ii) Filter off the excess solid.
- (iii) Saturate the filtrate by evaporation and cool the solution to crystallize.
- (iv) Filter and wash the crystals with distilled water.
- (v) Dry the crystals using filter paper or sunlight.

2. Describe how you can prepare pure crystals of iron (II) chloride from Iron (II) oxide in the laboratory

- (i) Add iron (II) oxide into a known volume of dilute hydrochloric acid in a beaker a little at a time while stirring with a glass rod until no more of the solid dissolves. (This shows that all the acid is neutralized).



- (ii) Filter off the excess solid.
(iii) Saturate the filtrate by evaporation and cool the solution to crystallize.
(iv) Filter and wash the crystals with distilled water.
(v) Dry the crystals using filter paper or sunlight.

3. Describe how you can prepare pure crystals of Sodium chloride from Sodium hydroxide and dilute hydrochloric acid in the laboratory

Procedure

- Step I:**
- Pipette 25 cm³ of sodium hydroxide solution into a conical flask and add two drops of phenolphthalein indicator.
 - Titrate it with hydrochloric acid from burette until the mixture just changes colour.
 - Find the volume, v , of the acid used (final volume – initial volume).

- Step II:**
- Pipette another 25 cm³ of the hydroxide into a clean beaker.
 - Measure exactly the same volume, v , of the acid used in step I and add it on to the sodium hydroxide in the beaker.
 - Mix them thoroughly without the indicator.

- Step III:**
- Saturate the solution by evaporation and cool it to crystallize.
 - Filter and wash the crystals with distilled water.
 - Dry the crystals using filter paper or sunlight.

12.44 Examples of preparation of insoluble salts

1. Describe how you can prepare lead (II) chloride in the laboratory

Procedure

- (i) Mix a solution of lead (II) nitrate and a solution of any soluble chloride e.g. sodium chloride solution or dilute hydrochloric acid.
White precipitate forms according to the equation.
- $$\text{Pb(NO}_3)_2 \text{ (aq)} + 2\text{NaCl (aq)} \rightarrow \text{PbCl}_2 \text{ (s)} + 2\text{NaNO}_3 \text{ (aq)}$$
- (ii) Filter the precipitates.
(iii) Wash the precipitates with distilled water to dissolve any soluble impurities.
(iv) Filter and dry the precipitates using filter paper or sunlight.

NB: *Pure crystals can be obtained by dissolving in minimum hot water and cooling to crystallization.*

(c) **Preparation of salts by method of direct synthesis**

In this method, the constituents in the required salt are made to react directly. This is done by passing dry gas over the heated metal in a combustion tube. The salts commonly prepared by the method of direct synthesis are:

- Iron (III) chloride (FeCl_3),
- Iron (II) chloride (FeCl_2) and
- Iron (II) sulphide (FeS)

Examples of preparation of salts by direct synthesis method

1. Preparation of iron (III) chloride by action of chlorine on iron

- (i) Place a coil of iron wire or iron fillings in a hard glass tube and arrange the apparatus as shown in figure 13.1 below.

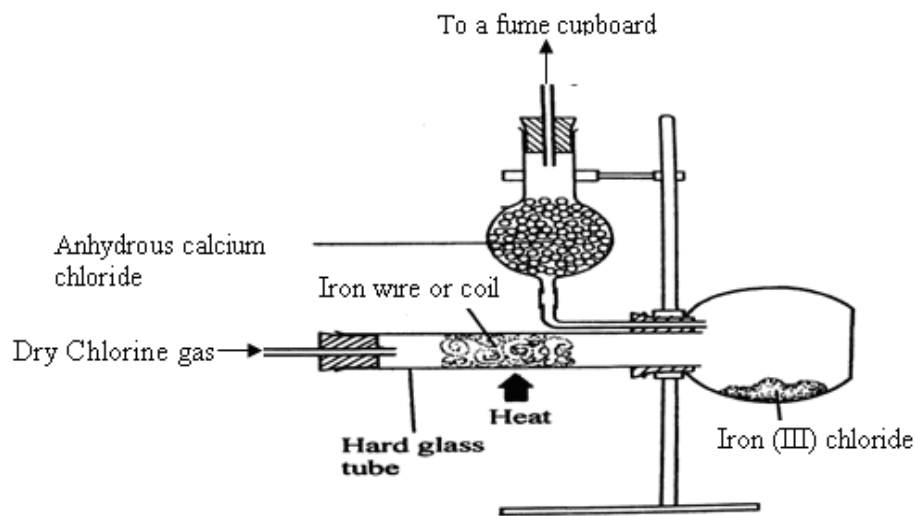
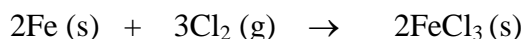


Figure 13.1

- (ii) Pass a stream of dry chlorine gas through the apparatus to sweep off air in the apparatus.
- (iii) Heat the wire until when it is red hot.
- (iv) Pass the dry chlorine over the hot iron.

The elements react according to the equation:



Observation

- (i) The wire glows.
- (ii) The reaction continues without application of flame.
- (iii) Black crystals of anhydrous iron (III) chloride collect in the small bottle, which acts as a condenser.

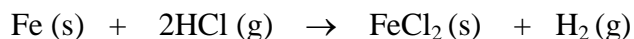
Conditions:

- (i) The iron must be heated strongly.
- (ii) The chlorine gas must be dry.
- (iii) The experiment should be carried in a fume cupboard.

NB: - *Heating can be stopped as soon as reaction starts.*
- *The black crystals of anhydrous iron (III) chloride should be removed and kept in a desiccator, as they are very deliquescent.*

2. Preparation of iron (II) chloride by action of Hydrogen chloride on iron

- (i) Place a coil of iron wire or iron fillings in a hard glass tube.
- (ii) Pass a stream of dry hydrogen chloride gas through the apparatus to sweep out air from it.
- (iii) Heat the wire strongly until it becomes red hot.
- (iv) Pass the dry hydrogen chloride over the hot iron as shown in the diagram below.



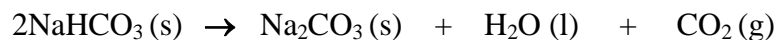
Observation: - White crystals of anhydrous iron (II) chloride are formed.

Solvay process (Manufacture of soda ash, sodium carbonate)

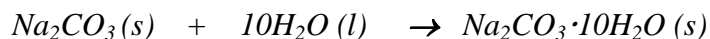
Very concentrated brine is saturated with ammonia gas in a tower and the ammoniacal brine is run down Solvay towers. Dry carbon dioxide gas is forced up through the mixture of brine and ammonia solution (ammoniacal brine). Sodium hydrogen carbonate is formed and precipitates out. The precipitation is quickened by cooling the lowest part of the tower.



The precipitate is filtered and washed free from ammonium compounds. It is then heated to convert it to sodium carbonate and the CO_2 is used again.



NB: *The salt formed is anhydrous sodium carbonate.*
If the crystalline form (washing soda) is required, the anhydrous sodium carbonate is dissolved in hot water and cooled to obtain the crystals.



The crystals are removed and allowed to dry.

Uses of sodium carbonate

Three of the important uses of sodium carbonate are the following:

- (i) Manufacture of glass.
- (ii) Manufacture of water-glass.
- (iii) Domestic water-softening.

12.45 Water of Crystallization

Some substances form crystals that contain water chemically combined with them. The water is necessary for their regular shape. The products of such compounds are called *hydrates* and the water is *water of hydration* or *water of crystallization*.

Definition

Water of crystallization is the definite amount of water with which some substances chemically combine when they form crystals from their solution in water.

Examples of hydrates are shown in table 13.7 below.

Name of compound	Chemical Formula of Compound
Calcium sulphate-2-water	$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$
Copper (II) sulphate-5-water	$\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$
Calcium Chloride-6-water	$\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$
Zinc sulphate-7-water	$\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$
Iron (II) sulphate-7-water	$\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$
Magnesium sulphate-7-water	$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$
Sodium carbonate-10-water	$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
Sodium sulphate-10-water	$\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$

Table 13.7

(a) Effect of heat on hydrates

When hydrates are heated strongly, the water of crystallization comes off and the crystals lose their shape and colour (if coloured) thus leaving the compounds without water of crystallization. Crystals without water of crystallization are called *anhydrous*.

(b) Effect of water in the atmosphere on salts

Salts behave in different ways when exposed to ordinary air. Some lose their water of crystallization to the atmosphere. While others absorb water vapour from the atmosphere. The changes observed when salts are exposed to air are classified as follows.

(i) Hygroscopy

Hygroscopy is the process in which a substance absorbs water vapour from the atmosphere.

The substances that absorb water vapour from the atmosphere are called *hygroscopic substances*. Some hygroscopic substances absorb a little water. For example, when anhydrous copper (II) sulphate is left in open, it absorbs a little water and changes colour from white to blue. While others take up a lot of water and dissolve in it to form a solution. Such hygroscopic substances are called *deliquescent substances*.

Examples of hydrates are shown in table 13.8 below.

Hygroscopic substance	
Name of Compound	Chemical Formula
Calcium oxide	CaO
Anhydrous Copper (II) sulphate	CuSO₄
Copper (II) oxide	CuO
Con. Sulphuric acid	H₂SO₄
Ethanol	C₂H₅OH
All deliquescent substances	

Table 13.8

(ii) Deliquescence

Deliquescence is the absorbing of water from the atmosphere by a solid to form a solution.

For example, when sodium hydroxide pellets are left in open air, they are quickly covered with drops of water. Thereafter, more water vapour is taken up and a colourless solution of sodium hydroxide is formed. The solid substance is called deliquescent substance.

All deliquescent substances are hygroscopic, but hygroscopic substances are not necessarily deliquescent.

Examples of deliquescent substances are shown in table 13.9 below.

Deliquescent substance	
Name of Compound	Chemical Formula
Potassium hydroxide	KOH
Sodium hydroxide	NaOH
Sodium nitrate	NaNO ₃
Calcium chloride	CaCl ₂
Iron (III) chloride	FeCl ₃
Phosphorous (V) oxide	P ₂ O ₅

Table 13.9

(iii) Efflorescence

Efflorescence is the giving up of water of crystallization by a crystal to the atmosphere. For example, when crystals of sodium carbonate-10-water are left in open air for a few hours, the crystals lose their water of crystallization to the atmosphere and the transparent crystalline structure collapses and changes to white powder. Substances such as sodium carbonate-10-water that give up their water of crystallization to the atmosphere are called efflorescent substances.

Examples of deliquescent substances are shown in table 13.10 below.

Deliquescent substance	
Name of Compound	Chemical Formula
Sodium carbonate-10-water	Na ₂ CO ₃ •10H ₂ O
Sodium sulphate-10-water	Na ₂ SO ₄ •10H ₂ O
Copper (II) sulphate-5-water	CuSO ₄ •5H ₂ O

Table 13.10

12.5 Solubility of Salts

Definition *The solubility of a solute in a solvent at a particular temperature is the number of grams of the solute required to saturate 100 g of the solvent at that temperature.*

Note: - *Solubility denotes a limit, that is, the maximum amount which can normally be held in solution.*
- *It can also sometimes be expressed in grams of solute per litre (dm^3) of solution at a given temperature.*

Experiment 13.51 To determine the solubility of a given salt at a given temperature

Apparatus/Requirements

Distilled water, evaporating dish, a spatula, a beaker, a boiling tube, a thermometer, weighing machine and source of heat

Procedure

- (i) Heat water in a beaker to a temperature slightly above the required temperature.
- (ii) Fill a boiling tube with distilled water and using a spatula, add crystals of the salt into the water in the boiling tube, stir well to dissolve the salt.
- (iii) Repeat procedure (ii) until no more salt dissolves (i.e. when the solution is saturated).
- (iv) Remove the flame, leave the apparatus undisturbed for the crystals to settle and cool.
- (v) Weigh a clean evaporating dish and record its mass, say x g.
- (vi) When the temperature is exactly as required, rapidly decant a little of the saturated solution of the salt into the dish and weigh the dish again and record its mass, say y g.
- (vii) Evaporate the solution carefully to dryness and weigh the dish plus the dry salt, say z g.
- (viii) Calculate the solubility of the salt (i.e. the mass of salt dissolved in 100 grams of water) at the given temperature.

Spacemen results

Mass of empty dish	= x g
Mass of empty dish + solution	= y g
Mass of solution only	= $(y - x)$ g

$$\begin{aligned}
\text{Mass of empty dish + dry salt} &= z \text{ g} \\
\text{Mass of salt only} &= (z - x) \text{ g} \\
\text{Mass of water} &= \text{Mass of solution} - \text{Mass of salt} \\
&= (y - x) - (z - x) \\
&= y - x - z + x \\
&= (y - z) \text{ g}
\end{aligned}$$

Calculations

$(z - x)$ g of salt saturates $(y - z)$ g of water.

1 g of water would be saturated by $\left(\frac{z - x}{y - z}\right)$ g of salt.

100 g of water would be saturated by $\left(\frac{z - x}{y - z}\right) \times 100$ g of salt.

\therefore the solubility of the salt = $\left(\frac{z - x}{y - z}\right) \times 100$ g per 100 g of water.

By repeating the experiment at various temperatures, several values can be obtained.

Solubilities in grams of solute per 100 g at various temperatures of some common substances are shown in the table below.

Temperature/ $^{\circ}\text{C}$	0	20	40	60	100
Sodium chloride (NaCl)	35.70	36.00	36.60	37.30	39.20
Sodium carbonate (Na_2CO_3)	0.70	21.50	48.00	47.00	45.50
Potassium nitrate (KNO_3)	14.00	31.60	61.40	106.00	245.00
Calcium sulphate (CaSO_4)	0.18	0.20	0.21	0.20	0.20
Calcium hydroxide [$\text{Ca}(\text{OH})_2$]	0.13	0.12	0.10	0.08	0.05

Table 13.10

NB: These figures are *NOT* to be memorized.

12.51 (a) Solubility Curves

A solubility curve is a graph showing how solubility varies with temperature. Some solubility curves are shown in figure 13.11 below.

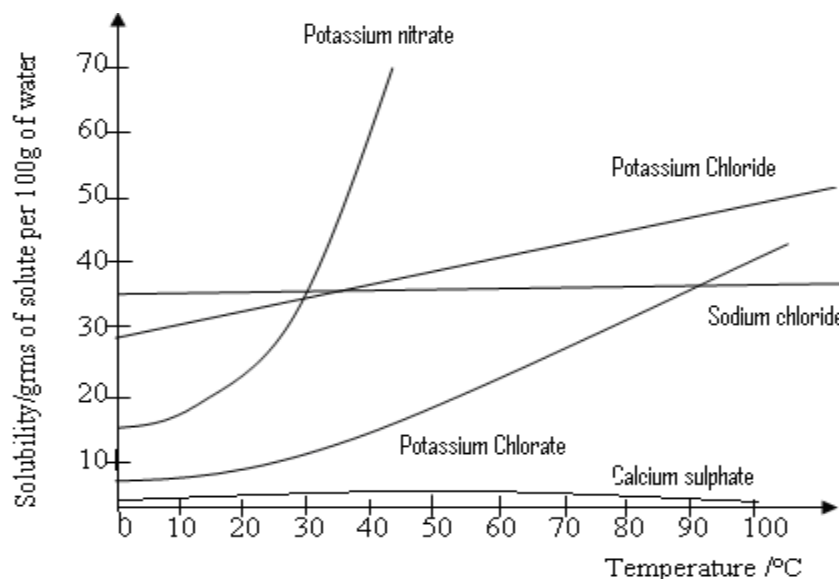


Figure 13.11

Using the graph in figure 13.2 above, answer the following questions:

- (a) For which salt does the solubility increase most rapidly with rise in temperature?
- (b) For which is there a decrease in solubility with increase in temperature?
- (c) For which salt does the solubility change little with rise in temperature?

(b) Effect of temperature on solubility

Solubility of most substances increase and decreases with temperature.

If the temperature of a saturated solution is increased, the solution becomes unsaturated and more solute dissolve in it. And when the temperature of the saturated solution is reduced to a lower temperature, the solubility also decreases and some of the solute crystallises out leaving the amount the solution can hold at that particular temperature.

12.52 Calculations on solubility

Examples

1. 50 g of water was saturated by 30.7 g of potassium nitrate at 40 °C.
Calculate; (i) the solubility of the salt.
(ii) the mass of the crystals that precipitates if the solution is cooled to a temperature of 20 °C (Given that the solubility of the potassium nitrate is 31.6 g at 20 °C).

Solution

- (i) 50 g of water dissolved 30.7 g of potassium nitrate.
1 g of water dissolves $\frac{30.7}{50}$ g of potassium nitrate.
 \therefore 100 g of water dissolves $\frac{30.7}{50} \times 100 \text{ g} = 61.4 \text{ g}$
 \therefore the solubility of potassium nitrate at 40 °C is 61.4 g/100 g of water.

- (ii) 100 g of water dissolves 31.6 g of potassium nitrate at 20 °C.

1 g of water will dissolve $\frac{31.6}{100}$ g of potassium nitrate.

\therefore 50 g of water dissolves $\frac{31.6}{100} \times 50 \text{ g} = 15.8 \text{ g}$.

That is, at 20 °C the mass of potassium nitrate in solution = 15.8 g

The mass of potassium nitrate crystals deposited = $30.7 - 15.8 = 14.9 \text{ g}$

12.53 Applications of solubility

Solubility is applied in the separation of mixture of salts in solution by a process called *fractional crystallisation*.

In this method, the mixture of the solution is cooled. The salt with low solubility crystallises out first and the crystals are filtered off. The filtrate (solution of the other salt) is evaporated to dryness.

12.6 Effect of heat on salts

The effect of heat on salts varies widely. Some salts are very stable and do not undergo any chemical change on heating. Other salts are unstable and undergo decomposition when heated.

(a) Nitrates

Nitrates are salts of nitric acid.

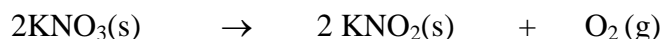
The properties of nitrates vary according to the position of the metal in the reactivity series. The lower the metal is in the series, the more readily and completely the nitrate of that metal decomposes.

(i) Metallic nitrates that decompose to form two products

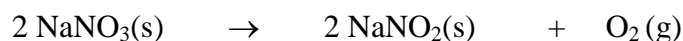
Potassium nitrate (KNO₃) and Sodium nitrate (NaNO₃)

Potassium nitrate and Sodium nitrate melt to form a colourless liquid. On further heating, they slowly decompose to form a pale-yellow *metallic nitrite* and *oxygen* as shown in the following equations:

Potassium nitrate \rightarrow Potassium nitrite + Oxygen



Sodium nitrate \rightarrow Sodium nitrite + Oxygen



Ammonium nitrate (NH₄NO₃)

Ammonium nitrate, NH₄NO₃, decomposes on heating to form di-nitrogen oxide (N₂O) and water.



NB: *Samples of ammonium nitrate can be explosive and it is best done in Situ by using a mixture of ammonium chloride and potassium nitrate.*

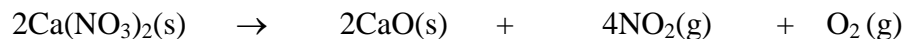
(ii) Metallic nitrates that decompose to form three products

The metallic nitrates from *calcium* to *copper* decompose to form the *oxide of the metal*, brown fumes of *nitrogen dioxide* and *oxygen*.

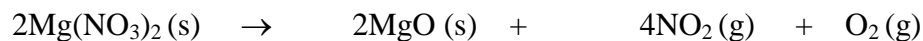
Metallic nitrate \rightarrow *Metal oxide* + *Nitrogen dioxide* + *Oxygen*

Example:

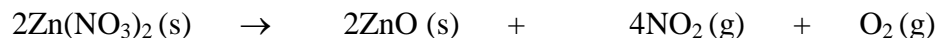
⇒ Calcium nitrate → Calcium oxide + Nitrogen dioxide + Oxygen



⇒ Magnesium nitrate → Magnesium oxide + Nitrogen dioxide + Oxygen

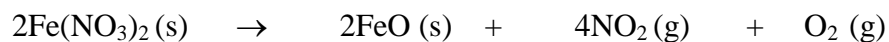


⇒ Zinc nitrate → Zinc oxide + Nitrogen dioxide + Oxygen

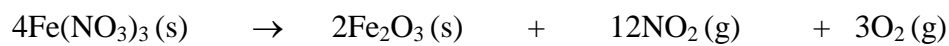


NB: *The zinc oxide is yellow when hot and white when cold.*

⇒ Iron (II) nitrate → Iron (II) oxide + Nitrogen dioxide + Oxygen



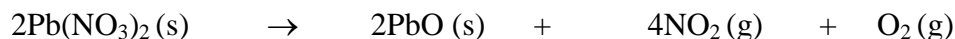
⇒ Iron (III) nitrate → Iron (III) oxide + Nitrogen dioxide + Oxygen



⇒ **Lead (II) nitrate**

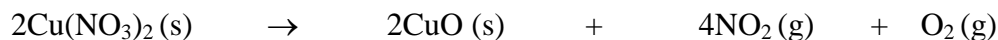
When lead (II) nitrate is heated, it decrepitates. That is it decomposes with a **crackling** sound giving off brown fumes of nitrogen dioxide and leaving a reddish brown solid when hot and yellow when cold.

Lead (II) nitrate → Lead (II) oxide + Nitrogen dioxide + Oxygen



NB: *The crackling sound is due to formation of gas molecules that form inside the crystals and splits them.*

⇒ Copper (II) nitrate → Copper (II) oxide + Nitrogen dioxide + Oxygen



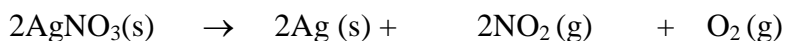
(iii) **Metallic nitrates that decompose to form three (3) products, metal, nitrogen dioxide and oxygen are:**

- Silver nitrate and
- Mercury

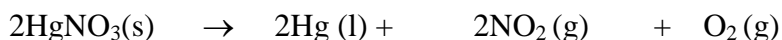
Why do these nitrates form metals when heated?

The nitrates of these metals decompose to form metal, nitrogen dioxide and oxygen. But the oxides are unstable to heat and therefore the oxides decompose to the metal and oxygen.

Silver nitrate \rightarrow Silver + Nitrogen dioxide + Oxygen



Mercury nitrate \rightarrow Mercury + Nitrogen dioxide + Oxygen



The summary of the effect of heat on nitrates

Metal	Effect of heat	Solubility
K Na	Nitrates of these metals are decomposed by heat to give two products : - The <i>nitrite</i> and <i>oxygen</i> .	All Nitrates are soluble in water
Ca Mg Al Zn Fe Pb Cu	Nitrates of these metals are decomposed on heating to give three products : - The <i>oxide of the metal</i> , - <i>Nitrogen dioxide</i> and - <i>Oxygen</i>	
Hg Ag	Nitrates of these metals are decomposed on heating to give three products : - <i>Metal, nitrogen dioxide</i> and <i>oxygen</i> .	

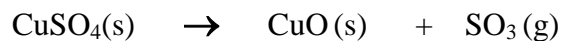
Table 13.11

(b) **Carbonates** See Chapter 13, page 228 section 13.45.

(c) **Sulphates**: Most sulphates are stable on heating.

(i) **Anhydrous copper (II) sulphate**

Anhydrous copper (II) sulphates decompose at very high temperatures to form sulphur trioxide and copper (II) oxide.



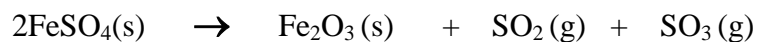
Observation: The blue crystals turn white and then black.

(ii) *Iron (II) sulphate-7-water*

Green Iron (II) sulphate-7-water crystals at first lose their water of crystallization.

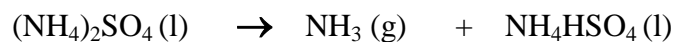


The white anhydrous salt decomposes on heating more strongly, forming reddish-brown iron (III) oxide, sulphur dioxide and sulphur trioxide.



(iii) *Ammonium sulphate*

Ammonium sulphate decomposes at its melting point (235°C) forming ammonia and either the hydrogensulphate, a colourless liquid which solidifies on cooling, or sulphuric acid.



Self-Check 12.2

- The substance that will undergo a permanent change when heated strongly is
A. I₂.
B. Cu(NO₃)₂.
C. NH₄Cl.
D. Na₂CO₃.
- Which one of the following nitrates will produce nitrogen dioxide when strongly heated?
A. Potassium nitrate.
B. Sodium nitrate.
C. Zinc nitrate.
D. Ammonium nitrate.
- Which of the following salts **cannot** be prepared by precipitation method?
A. Lead (II) nitrate.
B. Lead (II) chloride.
C. Lead (II) sulphate.
D. Lead (II) carbonate.
- Which one of the following salts can be prepared by precipitation method?
A. Zinc nitrate.
B. Copper (II) sulphate.
C. Barium sulphate.
D. Iron (III) chloride.
- Which one of the following is an acid salt?
A. CaSO₄.
B. NH₄Cl.
C. KNO₃.
D. NaHCO₃.
- The substance that decomposes when heated strongly is
A. potassium carbonate.
B. potassium hydroxide.
C. potassium chloride.
D. potassium nitrate.

7. Which one of the following pairs of substances will produce a precipitation when their aqueous solutions are mixed together?
 A. Sodium carbonate and ammonium sulphate.
 B. Nitric acid and ammonia.
 C. Potassium chloride and calcium nitrate.
 D. Lead nitrate and Sulphuric acid.
8. Which one of the following salts is insoluble in water?
 A. Potassium chloride.
 B. Calcium chloride.
 C. Sodium chloride.
 D. Silver chloride.
9. Which one of the following salts does **not** decompose on heating?
 A. Sodium carbonate.
 B. lead nitrate.
 C. Calcium carbonate.
 D. Zinc nitrate.
10. Which one of the following salts can be prepared by precipitation?
 A. Ammonium nitrate.
 B. Silver chloride.
 C. Sodium carbonate.
 D. Zinc sulphate.
11. Which one of these salts is the **least** soluble in water?
 A. $\text{CaCl}_{2(s)}$.
 B. $\text{CaCO}_{3(s)}$.
 C. $\text{Ca}(\text{NO}_3)_{2(s)}$.
 D. $\text{CaSO}_{4(s)}$.
12. Which one of the following salts is the most soluble in water with increasing temperature?
 A. Potassium nitrate.
 B. Sodium chloride.
 C. Sodium nitrate.
 D. Potassium chloride.
13. When ammonium nitrate is heated, it produces
 A. nitrogen dioxide.
 B. ammonia.
 C. dinitrogen oxide.
 D. nitrogen monoxide
14. Which one of the following salts is soluble in water?
 A. Lead carbonate.
 B. Sodium carbonate.
 C. Barium carbonate.
 D. Calcium carbonate.
15. Which one of the following salts is soluble in water?
 A. Lead chloride.
 B. Calcium sulphate.
 C. Sodium nitrate.
 D. Barium carbonate.
16. The product formed when silver nitrate is heated until there is no further change is
 A. silver oxide, nitrogen dioxide and oxygen.
 B. silver metal, nitrogen dioxide and oxygen.
 C. silver metal and nitrogen dioxide.
 D. silver oxide and nitrogen dioxide.
17. Which one of the following carbonates decomposes when heated to give an alkaline colourless gas?
 A. Calcium carbonate.
 B. Zinc carbonate.
 C. Potassium carbonate.
 D. Ammonium carbonate.
18. When copper (II) nitrate is strongly heated the gases evolved are
 A. oxygen and nitrogen.
 B. oxygen and nitrogen dioxide.
 C. nitrogen and ammonia.
 D. ammonia and nitrogen dioxide.

19. The solubility of a solute can be increased by
 A. agitating the mixture. B. adding more solute.
 C. adding more solvent. D. increasing the temperature.
20. Which one of the following carbonates decomposes to leave a metal?
 A. Potassium carbonate. B. Magnesium chloride.
 C. Zinc carbonate. D. Silver carbonate.
21. Which one of the following substances is a deliquescent substance?
 A. Calcium chloride. B. Sodium carbonate.
 C. Cobalt (II) chloride. D. Iron (II) sulphate.
22. Which one of the following salts can be prepared by direct synthesis?
 A. Sodium chloride. B. Sodium sulphate.
 C. Sodium carbonate. D. Sodium nitrate.
23. The solubility of copper (II) sulphate at 30 °C is 25 g per 100 g of water. The mass of copper (II) sulphate that would crystallise if a solution containing 50 g of copper (II) sulphate in 100 g of water at 60 °C is cooled to 30 °C is
 A. 12.5 g. B. 25.0 g. C. 50.0 g. D. 75.0 g.
24. Lead (II) chloride can **best** be prepared by the reaction between
 A. lead metal and concentrated hydrochloric acid.
 B. lead oxide and dilute hydrochloric acid.
 C. lead nitrate and dilute hydrochloric acid.
 D. lead carbonate and dilute hydrochloric acid.
25. Which one of the following nitrates does **not** decompose completely on heating?
 A. Copper nitrate. B. Lead nitrate.
 C. Sodium nitrate. D. Silver nitrate.
26. Which one of the following nitrates does **not** produce nitrogen dioxide when heated strongly?
 A. KNO_3 . B. $\text{Ca}(\text{NO}_3)_2$.
 C. NH_4NO_3 . D. $\text{Cu}(\text{NO}_3)_2$.
27. Which one of the following salts can be best prepared by precipitation?
 A. Sodium chloride. B. Lead (II) chloride.
 C. Potassium chloride. D. Copper (II) chloride.
28. Which one of the following solutions forms a precipitate when heated?
 A. Calcium hydrogen carbonate. B. Potassium hydrogen carbonate.
 C. Sodium hydrogen carbonate. D. Ammonium carbonate.
29. The solubility of salt W is 35 g per 100 cm^3 of water at 20°C. The mass of W in 40 cm^3 of water at the same temperature is
 A. 7.0 g. B. 14.0 g. C. 87.5 g. D. 114.3 g.
30. Which one of the following mixtures would not form a precipitate?
 A. Barium nitrate and sodium sulphate. B. Lead nitrate and potassium iodide.
 C. Copper nitrate and sodium sulphate. D. Silver nitrate and potassium bromide

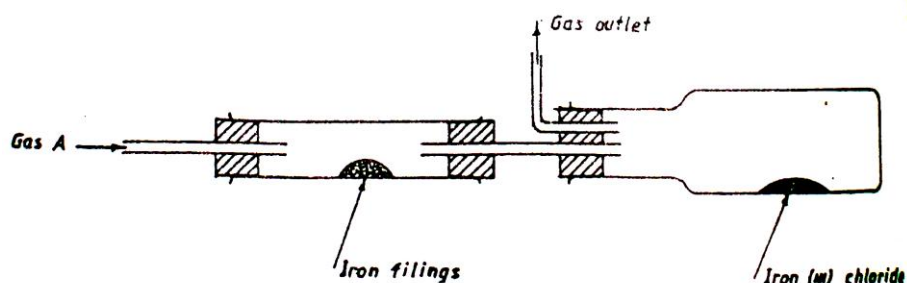
SECTION B

31. (a) Define the following terms, in each case give an example.
(i) a normal salt
(ii) an acid salt
- (b) A mixture containing copper (II) sulphate and copper (II) carbonate was shaken with excess water and filtered.
(i) Identify the residue.
(ii) The dry residue was heated strongly. State what was observed.
(iii) Write an equation for the reaction.
- (c) State what would be observed if dilute nitric acid followed by a few drops of barium nitrate solution was added to a solution of sodium sulphate.
32. (a) What is meant by the term “solubility”?
- (b) Describe how the solubility of potassium chloride can be determined in the laboratory.
- (c) The table below shows the solubilities of potassium chloride and potassium nitrate at various temperatures.

Temperature (°C)	0	20	40	60
Solubility of potassium chloride (g)	28.2	33.5	38.8	44.7
Solubility of potassium nitrate (g)	12.9	31.8	61.2	108.2

- (i) On the same axes, plot graphs of solubilities of potassium chloride and potassium nitrate against temperature.
- (ii) Determine the temperature at which the concentrations of the two salts are equal.
- (iii) Which of the two salts dissolves more rapidly with increase in temperature?
- (iv) State what would happen if a saturated solution of potassium chloride at 40°C was cooled to 30°C.
- (d) Of what industrial application is the study of solubility of salts?
33. (a) Describe how you would prepare pure crystals of lead (II) nitrate in the laboratory starting from lead (II) oxide. Write an equation for the reaction that takes place.
- (b) State what happens when lead (II) nitrate is strongly heated.
34. (a) Describe how a dry sample of copper (II) sulphate may be prepared from copper (II) oxide.
- (b) Crystals of lead (II) nitrate were heated gently and then strongly until there was no further change. Write equation(s) for the reaction(s) that took place.

35. The diagram below shows the apparatus that can be used to prepare anhydrous iron (III) chloride.



- (a)
 - (i) Name the gas A.
 - (ii) State the conditions for the reaction between iron filings and gas A.
 - (iii) Describe what would be observed during the reaction.
 - (iv) Write equation for the reaction.
- (b) Describe how you would prepare pure crystals of iron (II) chloride in the laboratory.

12.7 Hard and Soft Water

Hard water: *Hard water is the water that does not readily form lather with soap.*

Soft water: *This is one that readily forms lather with soap.*

Causes of hardness of water

Hardness of water can be:

- (i) Temporary hard water.
- (ii) Permanent hard water.

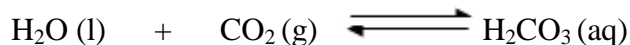
(a) Temporary hardness

This is caused by the presence of:

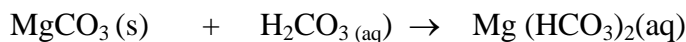
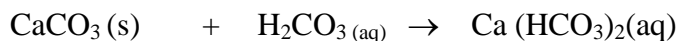
- Calcium hydrogen carbonate ($\text{Ca}(\text{HCO}_3)_2$) and
- Magnesium hydrogen carbonate ($\text{Mg}(\text{HCO}_3)_2$)

Formation of the hydrogen carbonate

As rain droplets fall through the atmosphere, they dissolve carbon dioxide in the atmosphere to form dilute carbonic acid.



If the carbonic acid formed passes through rocks containing calcium carbonate or magnesium carbonate, some dissolve in it to form the hydrogen carbonates.



(b) Permanent hardness

Permanent hardness is caused due to the presence of:

- Calcium sulphate (occurs naturally as *gypsum* $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$)
- Magnesium sulphate.

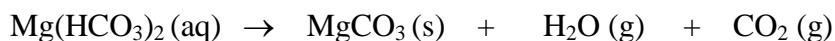
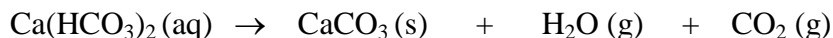
Calcium and magnesium sulphates are slightly soluble in water.

If rain water passes through rocks containing these sulphates, some of them dissolve in water causing it to be hard water. The sulphates react with soap.

12.71 Methods of removal of hardness

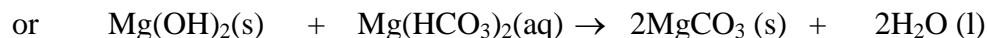
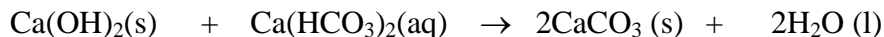
(i) Boiling

This method removes temporary hardness only. Heat decomposes the hydrogen carbonate into calcium carbonate or magnesium carbonate and carbon dioxide according to the equation:



(ii) Addition of calcium hydroxide (slaked lime)

Slaked lime reacts with the calcium hydrogen carbonate and magnesium hydrogen carbonate to form insoluble carbonates of calcium and magnesium. The salts formed are precipitated.

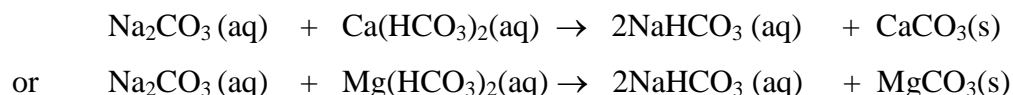


This method can only be used to treat temporary hardness.

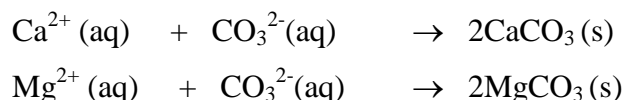
(iii) Addition of sodium carbonate

A solution of sodium carbonate precipitates both calcium and magnesium carbonates. Hence, it is used for the treatment of both temporary and permanent hardness.

Removal of temporary hardness



Removal of permanent hardness



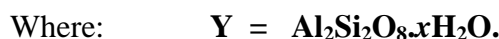
NB: *The solid carbonates formed are filtered off.*

(iv) Distillation

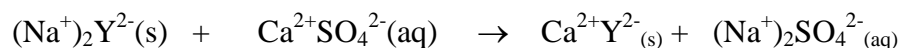
This method treats both temporary and permanent hardness in water. The water is boiled to form steam which condenses back to water. The dissolved solids remain in the distillation as residue.

(v) Permutit method of softening permanent hardness of water

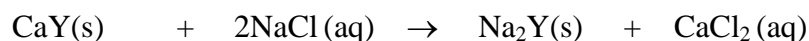
Permutit is a complex substance (hydrated sodium aluminium silicate) regarded as Na_2Y .



This method operates on the principle of *ion-exchange*. When a dissolved calcium salt runs over the permutit, ion-exchange occurs.



The calcium permutit formed is made fresh to sodium permutit by running concentrated sodium chloride (brine) solution over it and washing away the soluble calcium chloride formed.



NB: *- In the treatment of larger supplies of water (but not so large as to be treated by lime method) the permutit is used.*

12.72 Advantages and disadvantages of hard water

(a) Advantages of hard water

- (i) The calcium compounds in hard water may help to form healthy teeth and bones.
- (ii) It helps to form healthy animal shells and eggs.

(b) Disadvantages of hard water

- (i) It wastes soap.
- (ii) The scum formed when soap reacts with hard water makes dirty marks on clothes.
- (iii) It forms fur in kettles and pans. The fur is a bad conductor of heat and therefore wastes fuel.
- (iv) It forms boiler scale inside boilers. The boiler scale wastes fuel since it is a bad conductor of heat and it may cause damage to the pipes and boiler as it may block them.

12.73 Water Treatment

Water treatment refers to the process of making impure water safe to drink. This is done

- by:
- Sedimentation,
 - Filtration and
 - Chlorination.

The suspended particles are removed by sedimentation in the storage lakes or reservoirs. However, not all particles settle and those remaining are removed by filtration through filter beds consisting of stones, gravel and fine sand. The beds remove even the smallest solid particles and also some bacteria on the particles. The last stage in water treatment is chlorination. This is the addition of chlorine or one of its compounds to kill disease-causing germs.

NB: *In some areas, the water supply is very hard and the water is made soft by adding calcium hydroxide into the reservoirs.*

Self-Check 12.3

SECTION A

1. Spring water decomposes on boiling to produce white solid particles. The solid particles are
 - A. calcium sulphate
 - B. calcium hydrogen carbonate
 - C. magnesium carbonate
 - D. magnesium hydrogen carbonate
2. Which one of the following pairs of compounds can cause temporary hardness of water?
 - A. Sodium hydrogen carbonate and calcium hydrogen carbonate.
 - B. Sodium hydrogen carbonate and magnesium hydrogen carbonate.
 - C. Potassium hydrogen carbonate and magnesium hydrogen carbonate.
 - D. Magnesium hydrogen carbonate and calcium hydrogen carbonate.
3. Which of the following reagents is used for softening hard water?
 - A. Na_2CO_3
 - B. Na_2SO_4
 - C. CaCO_3
 - D. CaSO_4

4. Which one of the following compounds does not cause hardness of water?
 A. Calcium sulphate. B. Sodium hydrogen carbonate.
 C. Magnesium sulphate. D. Calcium hydrogen carbonate.
5. Which one of the following salts is used in water purification?
 A. Ammonium nitrate. B. Magnesium carbonate.
 C. Iron (III) sulphate. D. Potassium chloride.
6. Which one of the following processes is not used to remove hardness in water?
 A. Treatment with sodium carbonate. B. Addition of calcium hydroxide
 C. Distillation of water. D. Addition of sodium structure.
7. Which one of the following is most common in hard water?
 A. SO_4^{2-} B. CO_3^{2-}
 C. Na^+ D. Mg^{2+}
8. Which one of the following compounds does not cause hardness of water?
 A. Calcium sulphate. B. Sodium carbonate.
 C. Magnesium sulphate. D. Calcium hydrogen carbonate.
9. How is permanent hardness of water removed?
 A. By boiling. B. By adding slaked lime.
 C. By adding washing soda. D. By adding ammonia.
10. One advantage of hard water is that
 A. it does not contain bacteria.
 B. it forms lather more easily with soap.
 C. it contains calcium compounds which help to form healthy bones.
 D. it forms scales in boilers which prevent the boilers from leaking.
11. Which one of the following types of water would take the greatest amount of soap solution to form lather using the same volume of each type?
 A. Rain water. B. Distilled water.
 C. Sea water. D. River water.
12. Potassium aluminium sulphate (potash alum) is used during the purification of water for
 A. removing colouring matter. B. killing harmful bacteria.
 C. removing suspended matter. D. making water soft.
13. Permanent hard water can be softened by
 A. boiling the water. B. adding calcium hydroxide.
 C. adding aqueous ammonia. D. adding sodium carbonate.

Each of the questions 14 to 18 consists of an assertion (statement) on the left-hand side and a reason on the right-hand side.

- Select:**
- A. if both the assertion and the reason are **true** statements and the reason is a correct explanation of the assertion.
 - B. if both the assertion and the reason are **true** statements but the reason is **not** a correct explanation of the assertion.
 - C. if the assertion is **true** but the reason is **not** a correct statement.
 - D. if the assertion is **not** correct but the reason is a **true** statement.

Instructions Summarised:

Assertion	Reason
A. True	True (reason is a correct explanation.)
B. True	True (Reason is not a correct explanation.)
C. True	Incorrect.
D. Incorrect	True.

14. Hard water requires a lot of soap to form a lather **because** Some of the soap is initially used in removing calcium ions
15. Permanent hardness of water is caused by the presence of magnesium and calcium ions in water **because** These elements form sulphate compounds.
16. Water purified by filtration is made suitable for drinking by adding alum (potassium aluminium sulphate) **because** Alum kills all the bacteria in the water
17. Calcium hydrogen carbonate causes permanent hardness of water **because** calcium hydrogen carbonate is an acid salt
18. The presence in water of dissolved sodium hydrogen carbonate does not make water hard **because** all sodium salts are soluble

In each questions 19 to 20 one or more of the answers given may be correct. Read each question carefully and then indicate the correct answer A, B, C or D according to the following:

A. if 1, 2, 3 only are correct.

B. if 1, 3 only are correct

C. if 2, 4 only are correct.

D. if 4 only is correct.

Instructions summarised:

Instruction Summarised			
A	B	C	D
1, 2, 3	1, 3	2, 4	4
only correct	only correct	only correct	only correct

19. Which of the following ions can cause hardness in water?
 1. Mg^{2+} 2. Fe^{2+} 3. Ca^{2+} 4. Pb^{2+}
20. Which one of the following compounds are responsible for causing fur in kettles used for boiling water?
 1. Calcium sulphate 2. Calcium carbonate
 3. Magnesium sulphate 4. Magnesium carbonate

SECTION B

21. (a) Name one substance that causes:
- (i) Permanent hardness of water
 - (ii) Temporary hardness of water.
- (b) State one disadvantage of using hard water.
- (c) (i) Write equation to show how permutit (Na_2Y) removes permanent hardness in water.
- (ii) State the principle on which it works.
22. Table II shows results obtained when soap solution was added to 10 cm^3 of water samples P, Q and R in separate containers.

Table II

	Before boiling			After boiling		
Sample of water	P	Q	R	P	Q	R
Volume of soap solution required to form permanent lather (cm^3)	2	8	5	2	8	3

- (a) Identify which sample was, rain water, temporary hard water and permanent hard water. Give reasons for your answer.
- (i) Rain water.
 - (ii) Temporary hard water.
 - (iii) Permanent hard water.
- (b) Name one substance which can cause permanent hardness in water.

CHAPTER THIRTEEN

CARBON AND ITS COMPOUNDS

Learning objectives

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

1. (a) Define the following terms:
 - (i) Allotropy
 - (ii) Allotropes
 - (b) State the allotropes of carbon.
 - (c) Describe the structures of Graphite and Diamond.
 - (d) State: (i) the physical properties of graphite and diamond.
 - (ii) Uses of graphite and diamond.
 2. (a) Describe the laboratory preparation of carbon dioxide.
 - (b) State the physical properties and chemical properties of carbon dioxide.
 - (c) Describe the chemical test for carbon dioxide.
 3. (a) Describe the laboratory preparation of carbon monoxide.
 - (b) State the Physical properties and chemical properties of carbon monoxide.
 4. (a) State the effect of heat on carbonates and hydrogen carbonates.
 - (b) Describe chemical test for carbonates.
-

13.0 Introduction

Carbon is a nonmetallic chemical element known by the symbol C. It is the fundamental building block of material in living organisms and is important to many industries. It belongs to group IV and period 2 in the periodic table.

Occurrence

Carbon occurs naturally as: - - Coal,
- Mineral oil,
- Carbonates and bicarbonates,
- Organic matter and
- Carbon dioxide in the air to the extent (0.03 – 0.04) %
by volume.

(a) Allotropy

Allotropy is the existence of an element in more than one form (without change of state).

The various forms are called allotropes.

Allotropes

Allotropes are two or more structural forms of the same element existing in the same physical state.

Allotropes of carbon

Carbon exists in three allotropic forms, namely:

- (i) Diamond
- (ii) Graphite
- (iii) Amorphous carbon

Diamond and graphite are crystalline while amorphous carbon is non-crystalline.

(b) Polymorphism

Polymorphism is the existence of a substance which can crystallize in more than one form.

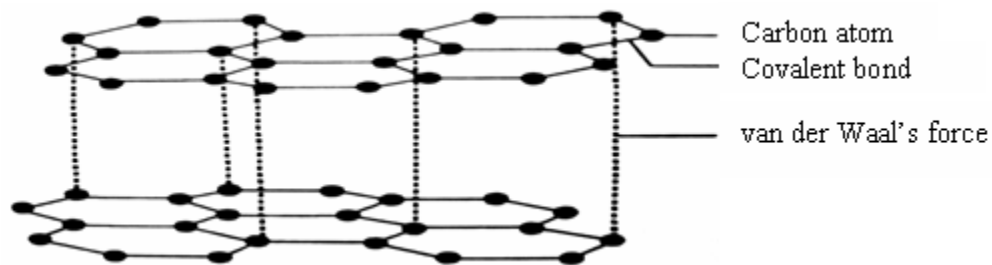
The various forms are called polymorphs. Carbon is therefore both *allotropic* and *polymorphic* element.

13.1 (a) Structures of graphite and diamond

Graphite

In a graphite structure, each carbon atom is bonded hexagonally to three other carbon atoms by covalent bonds in two dimensions resulting into parallel layers. Each layer is a giant atomic molecule. The layers are held together by weak intermolecular forces called van der Waal's forces. Because of this, the layers are able to slide over each other on application of slight pressure. This helps to explain the softness and greasy feel of graphite. Of the four valence electrons in each carbon atom, only three are involved in bond formation. The remaining electron is free (delocalized) and is able to carry charge along the layers of the graphite structure. This makes graphite a good conductor of electricity along the layers. But it is a poor conductor across the layers.

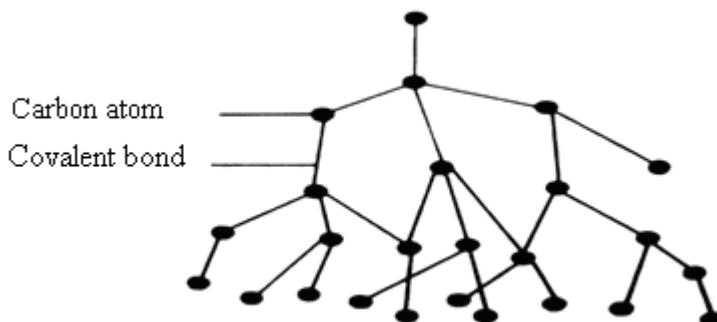
Structure of graphite



Diamond

In diamond, each carbon atom is bonded tetrahedrally to four other carbon atoms using strong covalent bonds through out the structure to form a three-dimensional lattice. All the valence electrons of the carbon atoms are involved in bond formation, leaving no free or delocalized electrons. This makes diamond a non-conductor of electricity and poor conductor of heat.

Tetrahedral structure of diamond



(b) Physical properties of graphite and diamond

	Diamond	Graphite
Appearance	Colourless, transparent and sparkling	Grey-black, opaque, shiny
Density(g/cm ³)	Highest 3.5	2.3
Hardness	Hardest natural substance known	Soft, greasy or soapy
Conductivity of heat and electricity	Poor conductor	Good conductor

13.2 Uses of carbon

(a) Diamond

- (i) Due to its hardness, diamond is used for making:
 - Abrasive tools,
 - Cutting tools e.g. glass cutters and
 - Drilling tools, e.g. drill bits used in quarrying and water drilling.
 - Surfaces that can withstand shock and pressure.
- (ii) Cut diamond sparkles brilliantly, this makes it treasured gemstone in jewelry.

(b) Graphite

Graphite is used in the following.

- (i) For making pencils.
- (ii) It is used as a lubricant in:
 - Dynamos,
 - Electric motors and
 - Fast moving parts of machinery
- (ii) It is used as electrodes in electrolysis.

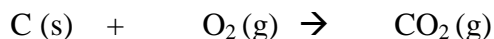
(c) Amorphous

- (i) Wood charcoal is used:
 - as source of fuel.
 - in gas masks to absorb poisonous gas.
- (ii) Animal charcoal is used to remove the brown colour from crude sugar and make it white.

13.3 Carbon dioxide

Occurrence:

Carbon dioxide occurs naturally in air to the extent of about 0.03% by volume. It is formed during the combustion of all common fuels:- coal, coke, coal-gas, water-gas, petrol, paraffin, diesel oil- all of which contain carbon.



In the laboratory, it is prepared by the action of dilute hydrochloric acid on marble chips (calcium carbonate).

Laboratory preparation of carbon dioxide

In the lab, carbon dioxide is prepared by dropping dilute hydrochloric acid from a dropping funnel on to marble chips placed in the flat bottom flask. Effervescence occurs liberating a colourless gas which is collected over water or by downward delivery method.



Diagram showing laboratory preparation of carbon dioxide

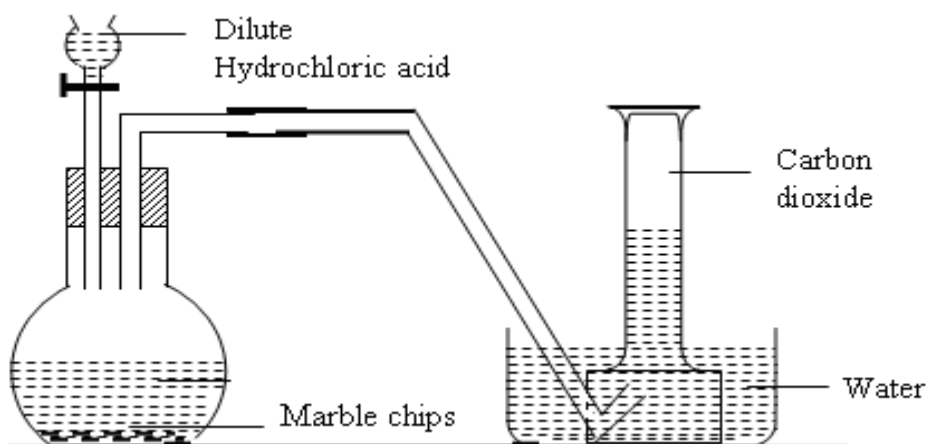


Figure 14.1

- NB:**
- (i) *The gas can also be prepared by pouring conc. hydrochloric acid down a thistle funnel on to marble chips covered with water.*
 - (ii) *If the gas is required dry, it is passed through solution of potassium hydrogen carbonate to absorb hydrogen chloride gas and then passed through anhydrous calcium chloride as shown in the diagram in figure 14.2 below.*

Diagram showing laboratory preparation of dry carbon dioxide

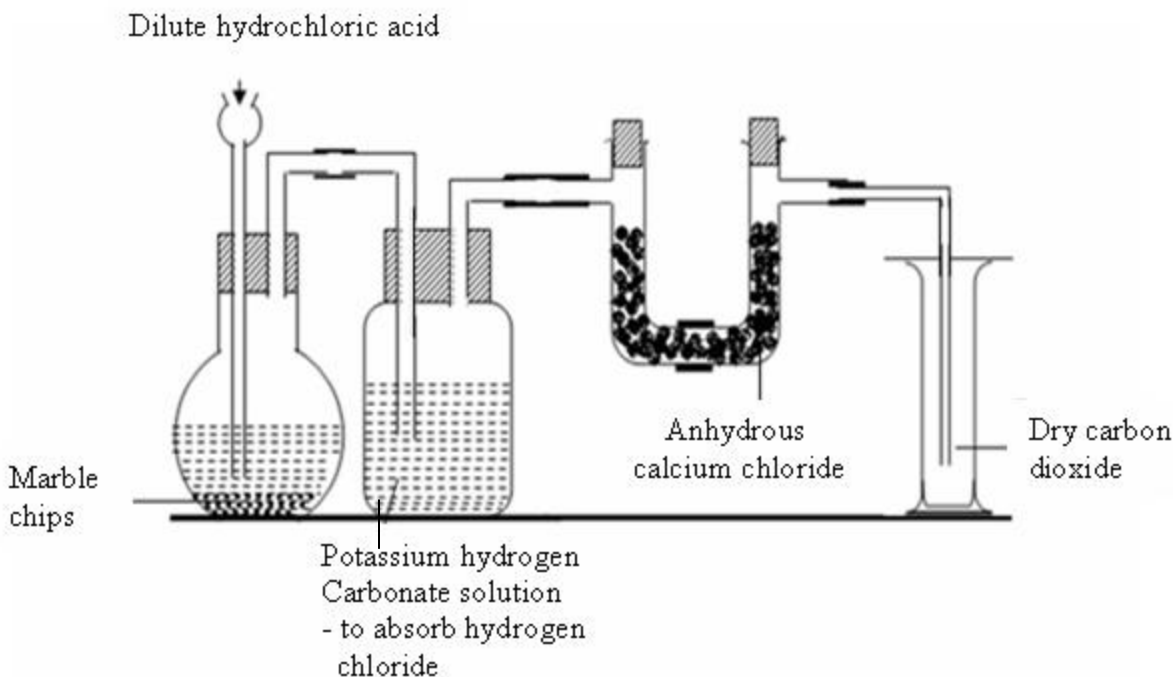


Figure 14.2

Properties of carbon dioxide (CO₂)

(a) Physical properties

- It is a colorless gas.
- It has a sharp pleasant taste.
- It is slightly soluble in water forming carbonic acid (weak acid).
- It is 1½ times denser than air.
- It does not burn; as a result, it is used as fire extinguisher.

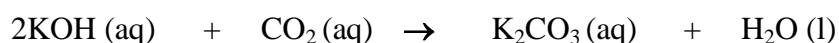
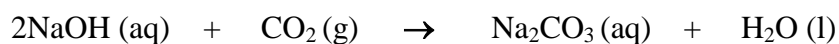
(b) Chemical properties

(i) Reaction with alkalis

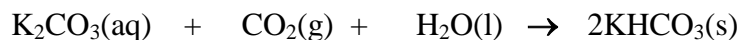
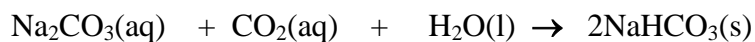
Carbon dioxide reacts with alkalis to form carbonates of metals and water.

⇒ *With Sodium hydroxides and Potassium hydroxides*

Carbon dioxide reacts with Sodium hydroxide and Potassium hydroxide to form soluble sodium carbonate potassium carbonate according to:



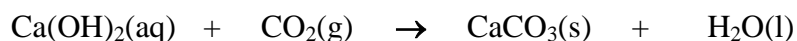
If excess carbon dioxide is bubbled through the alkalis, the carbonates formed react with it to form sodium hydrogen carbonates and potassium hydrogen carbonates respectively.



- NB:**
- *These reactions are the reactions for the preparation of sodium hydrogen carbonate and potassium hydrogen carbonates.*
 - *The precipitates formed are filtered and washed with a little cold water and then heated to a constant mass in an evaporating dish to obtain fine white powder of the salts.*

⇒ Calcium hydroxide

When carbon dioxide is bubbled through a solution of calcium hydroxide, the solution first turns milky. This is due to the formation of a white precipitate of calcium carbonate according to the equation.

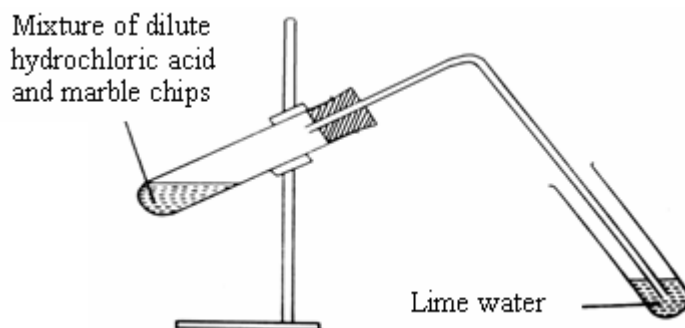


If excess carbon dioxide is passed through, the solution turns clear again. This is because the insoluble calcium carbonate dissolves to form soluble calcium hydrogen carbonate.



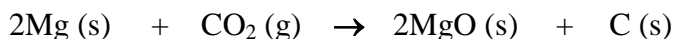
Test for carbon dioxide (CO₂)

When carbon dioxide is passed through lime water (a solution of calcium hydroxide), the solution turns milky.



(ii) Reaction with burning magnesium

When a piece of burning magnesium is lowered into carbon dioxide in a gas jar, it continues to burn for a short time with a splattering flame (making a series of short explosive sounds), and black specks of carbon form on the sides of the jar. The magnesium burns to form white magnesium oxide.



Uses of carbon dioxide

1. Fire extinguisher

Carbon dioxide is a fire extinguisher (i.e. it is used for putting out burning fires).

2. Aerated drinks

Carbon dioxide is dissolved under pressure to give a pleasant taste to some drinks.
E.g. Coke, lemonade, e.t.c.

3. Refrigerator

Carbon dioxide is used as a refrigerant in refrigerators to “deep freeze” foods and drinks.

4. Making of rain

Pieces of solid carbon dioxide (dry ice) are dropped into clouds to cool them to form rain.

5. Cooking

Bubbles of carbon dioxide form inside the dough of cakes and ‘rises’ when heated in an oven during making of bread.

6. Manufacture of soda ash

Carbon dioxide is used in the manufacture of sodium carbonate in solvay process.

13.4 Carbon monoxide (CO)

Carbon monoxide gas can be prepared from three main sources mainly:-

- Oxalic acid,
- Sodium formate and
- Carbon dioxide

The most convenient source is oxalic acid.

Laboratory preparation of carbon monoxide from oxalic acid

Method I

In the laboratory, carbon monoxide is prepared by pouring concentrated sulphuric acid down a thistle funnel on to crystals of oxalic acid ($\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$) in a round bottom flask and the mixture is warmed gently. The white crystals dissolve, with effervescence liberating a mixture of carbon monoxide and carbon dioxide. The mixture of the two gases is passed through a concentrated solution of potassium hydroxide to absorb carbon dioxide and the gas (carbon monoxide) is collected over water.

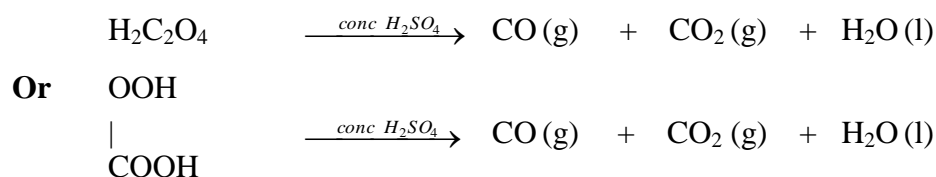


Diagram showing the laboratory preparation of carbon monoxide

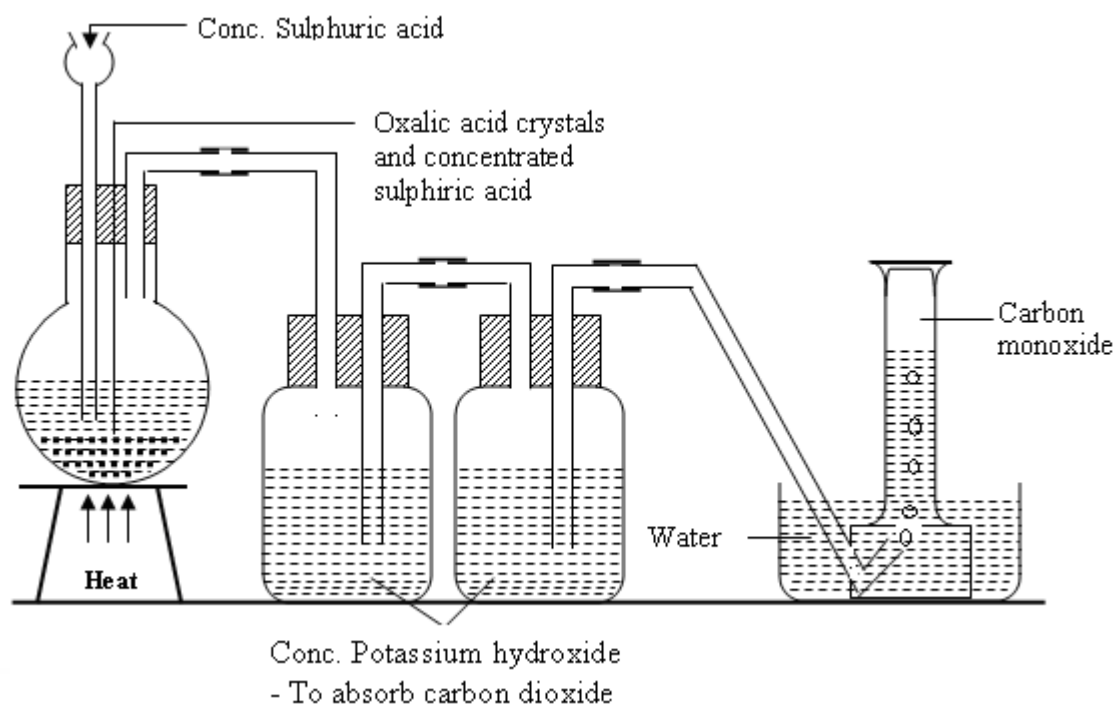


Figure 14.1

NB: The hot concentrated sulphuric acid removes the elements of water from the molecule of oxalic acid, leaving a mixture of equal volumes of carbon monoxide and carbon dioxide

Method II: Preparation of carbon monoxide from sodium methanoate (sodium formate)

In the laboratory, carbon monoxide can be prepared by dropping concentrated sulphuric acid on to crystals of sodium formate in a flat-bottomed flask. The reaction takes place in the cold, effervescence occurs and carbon monoxide is collected over water.



Diagram showing laboratory preparation of carbon monoxide

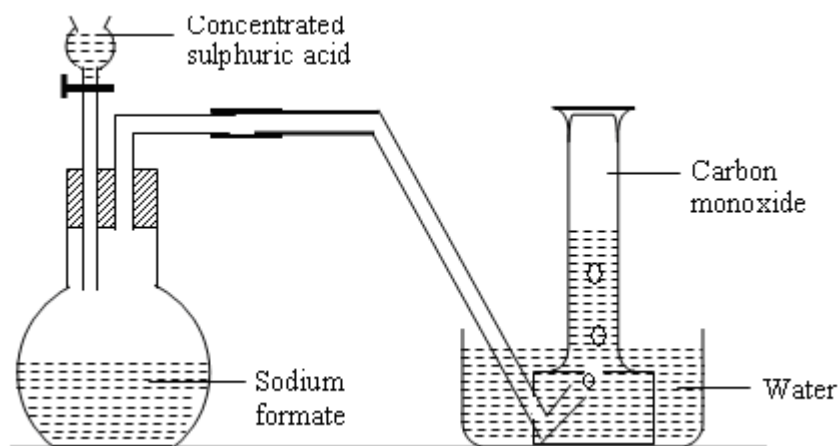


Figure 14.1

Properties of carbon monoxide

(a) Physical properties of CO

- (i) It is colourless.
- (ii) It is odourless (does not smell).
- (iii) Lighter than air.
- (iv) It does not support combustion.
- (v) It is a neutral gas (i.e. has no effect on litmus paper).
- (vi) It is a poisonous gas.
- (vii) It is slightly soluble in water.

If inhaled, it combines with haemoglobin of the blood to give brick red substance called carboxy haemoglobin. Too much of it kills.

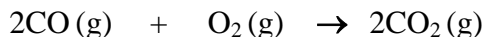
Test for carbon monoxide

Carbon monoxide burns in air with *blue flame* forming carbon dioxide (the latter will turn lime water milky or turbid).

(b) Chemical properties of Carbon monoxide (CO)

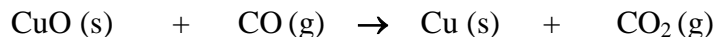
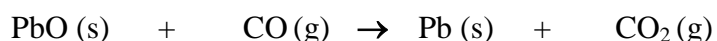
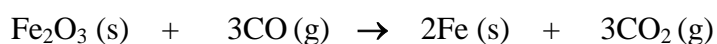
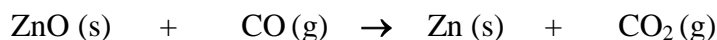
(i) Combustion

Carbon monoxide burns in air with a blue flame, forming carbon dioxide according to the equation.



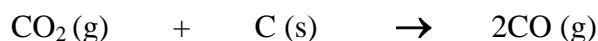
(ii) Reduction of oxides of metals below carbon in reactivity series

Carbon monoxide is a strong reducing agent. It reduces (i.e. readily removes oxygen from) heated oxides of zinc, iron, lead and copper to form the metals and is itself oxidized to carbon dioxide.



(iii) Reduction of carbon dioxide to carbon monoxide

Carbon dioxide is reduced by passing it over red-hot coke.



13.45 Carbonates and Hydrogen Carbonates

(a) Carbonates

Carbonates may be regarded as salts derived from carbonic acid, H_2CO_3 , which is formed when carbon dioxide is dissolved in water.

Properties of carbonates

Summary of the important properties of the common carbonates

Metal	Solubility	Action of heat	Action of acid
K Na	Carbonates of these metals are soluble in water	Carbonates of these metals do not decompose on heating	Carbonates react with acids to form or liberate carbon dioxide gas
Ca Mg Zn Fe Pb Cu	Carbonates of these metals are insoluble in water	Carbonates of these metals decompose on heating to form oxides of the metals and carbon dioxide gas	

NB. (i) **Aluminium* -forms no carbonate.

- (ii) Ammonium carbonate, $(\text{NH}_4)_2\text{CO}_3$, is soluble in water and when heated, decomposes to carbon dioxide (CO_2), ammonia (NH_3), and water.



Test for carbonates

Add dilute nitric acid to suspected carbonate in a test tube.

Observation

If carbonate is present, effervescence occurs liberating a colourless gas which turns lime water milky.

(b) Hydrogen carbonates

These are acid salts of carbonic acid formed by the replacement of only one of the hydrogen ions by a metal ion.

Properties of Hydrogen carbonates

(i) Physical state:

Solid: There are three solid hydrogen carbonates namely:-

- Potassium hydrogen carbonate (KHCO_3)
- Sodium hydrogen carbonate (NaHCO_3)
- Ammonium hydrogen carbonate (NH_4HCO_3)

There are also three hydrogen carbonates that exist only in solution. They include:-

- Lithium hydrogen carbonate (LiHCO_3).
- Calcium hydrogen carbonate ($\text{Ca}(\text{HCO}_3)_2$).
- Magnesium hydrogen carbonate ($\text{Mg}(\text{HCO}_3)_2$).

(ii) Solubility

The three hydrogen carbonates that exist in solid state are all soluble in water.

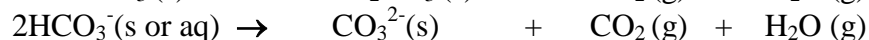
(iii) Action of acids

Like carbonates, all hydrogen carbonates react with acids to form carbon dioxide, water & salts.



(iv) Action of heat

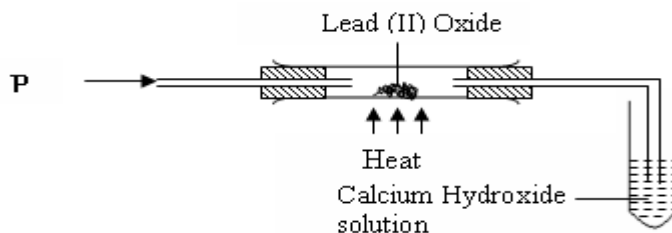
All hydrogen carbonates decompose to form the metal carbonate, carbon dioxide and water when heated.



Self-Check 13

- Which one of the following substances is formed when magnesium burns in carbon dioxide?
A. Magnesium carbonate.
B. Magnesium nitride.
C. Carbon monoxide.
D. Carbon.
- The substance that will undergo a permanent change when heated strongly is
A. I_2 .
B. $CuCO_3$.
C. NH_4Cl .
D. Na_2CO_3 .
- Which one of the following processes is **not** involved in the carbon cycle?
A. Respiration.
B. Combustion.
C. Photosynthesis.
D. Lightning.
- Which one of the following is **not** a property of carbon dioxide?
A. It is slightly soluble in water.
B. It forms a precipitate with lime water.
C. It extinguishes burning magnesium ribbon.
D. It sublimates when solid.
- Which one of the following statements about graphite is false?
A. It has a layer structure.
B. It refracts light rays strongly.
C. It is an excellent lubricant in oil.
D. It has a high density.
- Which one of the following statements about sodium hydrogen carbonate is correct?
A. It decomposes to carbon dioxide and hydrogen.
B. It does not exist in solid form.
C. Its solution turns milky in carbon dioxide.
D. It reacts with acids to yield carbon dioxide.
- Carbon monoxide can be obtained from carbon dioxide by
A. passing heating carbon dioxide over heated carbon.
B. passing carbon dioxide over heated calcium.
C. heating a mixture of carbon dioxide and steam.
D. passing carbon dioxide over heated copper.
- Carbon dioxide is normally, **not** prepared by the action of
A. nitric acid on calcium carbonate.
B. nitric acid on sodium hydrogen carbonate.
C. sulphuric acid on calcium carbonate
D. sulphuric acid on sodium carbonate.
- Which one of the following gases reduces hot Copper (II) oxide to copper:
A. Carbon dioxide.
B. Carbon monoxide.
C. Nitrogen dioxide.
D. Nitrogen monoxide.
- Which one of the following best explains why graphite and diamond differ?
A. Graphite is impure carbon.
B. Their densities are not the same.
C. All the four valency electrons of graphite are used in covalent bonding.
D. Their atomic structures are different.

11. Which one of the following carbonates decomposes when heated to give an alkaline colourless gas?
A. Calcium carbonate.
B. Zinc carbonate.
C. Potassium carbonate.
D. Ammonium carbonate.
12. Which one of the following metal oxides is reduced by carbon?
A. Magnesium oxide.
B. Sodium oxide.
C. Calcium oxide.
D. Iron oxide.
13. Which of the following substances does **not** exhibit allotropy?
A. Carbon.
B. Sulphur.
C. Phosphorus.
D. Chlorine.
14. Which one of the following carbonates decomposes to leave a metal?
A. Potassium carbonate.
B. Zinc carbonate.
C. Magnesium chloride.
D. Silver carbonate.
15. A white solid formed when lime water is left exposed to air for sometime is
A. calcium carbonate.
B. calcium oxide.
C. calcium hydroxide.
D. calcium chloride.
16. A mixture of carbon monoxide, ammonia and carbon dioxide was bubbled through lime water. Which one of the following gases was not absorbed?
A. Carbon dioxide.
B. Carbon monoxide.
C. Carbon monoxide and ammonia.
D. Carbon dioxide and ammonia.
17. Which one of the substances below conducts electricity in the solid state?
A. Graphite.
B. Sulphur.
C. Iodine.
D. Phosphorus.
18. Which one of the following substances given below can burn in carbon dioxide?
A. Magnesium.
B. Aluminium.
C. Zinc.
D. Lead.
19. Graphite is used as an electrode in electrolysis because it
A. has hexagonal carbon rings.
B. is soft.
C. has mobile electrons.
D. is opaque.
20. A colourless gas P, was passed over red-hot lead (II) oxide as shown in the diagram below



When the gas that came off was bubbled through calcium hydroxide solution, a white precipitate was formed P is

- | | |
|--------------|---------------------|
| A. hydrogen. | B. carbon dioxide. |
| C. butane. | D. carbon monoxide. |

SECTION B

21. (a) Name two crystalline allotropes of carbon.
(b) State two differences between the allotropes you have named in (a).
(c) Give one use of each of the allotropes.
22. (a) (i) Define the term *allotropes*.
(ii) Name *one* example of an element that shows allotropy other than carbon.
(b) (i) Give the allotropes of carbon.
(ii) State *two* properties of one of the allotropes of carbon you named in (i).
(iii) Explain how the allotrope is used due to its properties named in (b)(ii) above.
23. (a) (i) Draw a labelled diagram to show how carbon dioxide can be prepared in the laboratory.
(ii) Write an ionic equation for the reaction leading to the formation of carbon dioxide.
(b) Carbon dioxide was passed through calcium hydroxide solution. Describe and explain the reaction that took place.
(c) (i) State what would be observed if burning magnesium ribbon was lowered into a jar of carbon dioxide.
(ii) Write equation for the reaction that takes place in (c)(i).
(d) Water was added to the product in (c) and the resultant mixture tested with litmus solution.
24. Carbon dioxide was bubbled into a fairly concentrated solution of sodium hydroxide and no visible change was observed at first. On further bubbling, a white precipitate was formed.
(a) Explain the observation.
(b) Write equations for the reaction.

ANSWERS TO SELF-CHECK QUESTIONS

CHAPTER 2

Self-Check 2

SECTION A

1. A.	2. C.	3. D.	4. D.	5. A.	6. C
7. C.	8. B.	9. D.	10. B.	11. B.	12. D.

SECTION B

13. (a) *Physical change is a change where no new substance is formed.*
(b) *Chemical change is a change where a new substance is formed.*
(c) (i) *Chemical change* *A new substance is formed.*
(ii) *Physical change* *Because the salt can be obtained by a physical means.*
(iii) *Physical change* *Because no new substance is formed.*
 The ammonium chloride change back solid.
(vi) *Chemical change* *A new substance is formed thus the process cannot be reversed..*

CHAPTER 3

Self-Check 3

SECTION A

1. A.	2. C.	3. A.	4. B.	5. A.	6. A.	7. A.	8. D.	9. B.	10. D.
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SECTION B

11. (a) *Chromatography:*
The dyes are absorbed to different extent on a paper and also differ in their solubilities in a solvent (Water or acetone) being used.
(b) *Fractional distillation:*
Liquids with different boiling points, usually close together, are separated by this method.
(c) *Sublimation:*
Iodine sublimates when heated whereas potassium chloride does not.
(d) *Evaporation:*
Copper (II) sulphate is soluble in water. Therefore the mixture can be dissolved in water, filtered and evaporated to obtain the crystals of copper (II) sulphate.
(e) *Magnetic method:*
Iron fillings are magnetic while sulphur is non-magnetic.

12. (a) Chromatography
 (b) (i) Orange
 (ii) Phenolphthalein changes to purple in alkaline solution.
 (c) Ink or extract from pigment of plant leaf.
13. (i) Red and green. In figure 3, the chromatography for food-colour is a combination of marks made by these two dyes.
 (Key: The dyes are in line with the food-colour spots).
- (ii) Till B
- (iii) Filter paper must dip in the solvent.
 Leave the arrangement for some time: then remove the paper and dry it.
 Different solubilities of dyes in the solvent: dyes are absorbed to different extents on the paper.
- (iv) Purity of drugs/food or separation of pigments from plants.
14. (a) The water connections to the condenser are not ideal: the condenser may not fill completely with the water and the cooling will be less efficient. Reverse the water connections to the condenser.
- (b) To produce even boiling throughout the liquid
- (c) To measure the temperature of vapors
- (d) To reduce the evaporation of the distillate, as it is exposed to the smaller surface area.
- (e) Dissolved salts remain as residue in the flask: they are not volatile. Take a little distillate in a clean evaporating dish. And heat it. All the liquid will evaporate, leaving no residue.
- (f) Ethanol is inflammable: it cannot be heated directly on burner flame. Use sand – bath or water-bath. Ethanol and water both would vaporize when heated, though more of ethanol would do so. The distillate will not be pure ethanol. A fractionating column should be used when substances evaporate together.
 Ethanol will distill off as it boils at 78°C whereas B.P of water is 100°C .
- (g) (i) Air can be separated like this : First the air is liquidified by cooling to 200°C and then the liquid air is fractionally distilled in an illustrated column. Nitrogen, having low boiling point (-195°C) begins to evaporate first. Oxygen, having higher boiling point (-183°C) comes off later.
- (ii) Crude oil is fractionally distilled to produce a number of substances with different boiling points which are closer in value. When crude oil is heated to 400°C , a mixture of gases evaporates. In a tower with fractionating columns, the substances with lower boiling points rises towards the top, while those with high boiling points condense near the bottom.

15. (a) *The black shiny solid iodine changes to violet vapour of iodine and condenses on a cooler surface to black shining plates.*
- (b) *Ability to sublime*
- (c) *Any one of the following mixtures:*
- *Sodium chloride and Ammonium chloride,*
 - *Sodium chloride and Iron (III) chloride*

CHAPTER 4

Self-Check 4

SECTION A

1. B.	2. A.	3. D.	4. C.	5. C.
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SECTION B

6. (a) (i) *An element is a substance that cannot be split into two or more simpler substances by any chemical means.*
- (ii) *A metal is an element which ionizes by loss of electron(s).*
- (iii) *A non-metal is an element which ionizes by gain of electron(s).*
- (b) *See chapter 4 table 4.1 on page 51.*
7. (a) (i) *Atomicity of an element is the number of atoms in one of its molecules.*
- (ii) *A radical is a group of two or more atoms that exist in several compounds but do not exist on its own.*
- (iii) *Valence is the number of hydrogen atom(s) which combine with or displace one atom of an element or a group of radical.*

(b)

Name of Element or radical	Chemical Symbol/ formula	Valency
<u>Iron (II) (Ferum)</u>	Fe	2
<u>Sodium</u>	<u>Na</u>	<u>1</u>
<u>Sulphate</u>	<u>SO₄</u>	<u>2</u>
<u>Carbonate</u>	CO ₃	<u>2</u>
<u>Hydroxide</u>	OH	<u>1</u>

8. (a) (i) 3 (ii) 2 (iii) 1 (iv) 2
- (b) Write the full meaning of the following compounds.
- (i) Means one molecule of sodium hydroxide containing one atom of sodium, one atom of oxygen and one atom of hydrogen that are chemically combined together.
- (ii) Means one molecule of sulphuric acid containing two atoms of hydrogen, one atom of sulphur and four atoms of oxygen that are chemically combined together.
- (iii) Means one molecule of magnesium sulphate containing one atom of magnesium, one atom of sulphur and four atoms of oxygen that are chemically combined together.
- (iv) Means one molecule of magnesium carbonate containing one atom of magnesium, one atom of carbon and three atoms of oxygen that are chemically combined together.
9. (i) Na_2SO_4 (ii) MgCl_2 (iii) $\text{Al}(\text{NO}_3)_3$ (iv) $\text{Al}_2(\text{SO}_4)_3$
10. (a) Metal + Oxygen \rightarrow Metallic oxide
- (b) Non-metal + Oxygen \rightarrow Non-metallic oxide
- (c) Base + Acid \rightarrow Salt + Water
- (d) Acid + Metallic carbonate \rightarrow Salt + Carbon dioxide + water

CHAPTER 5

Self-Check 5

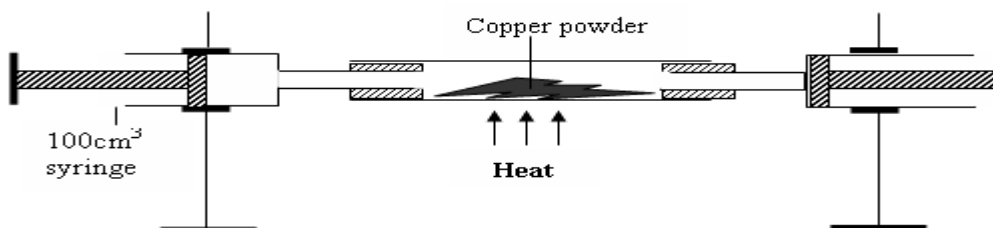
SECTION A

1. C.	2. C.	3. A.	4. B.	5. C.	6. A.	7. D.	8. B.	9. B.	10. A.
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SECTION B

11. (a) Atmosphere is the air surrounding the earth.
- (b) 21% & 78%.
- (c) (i) Combustion, respiration, decaying, germination etc.
- (ii) Photosynthesis.

12. (a) *Iron, water and oxygen.*
 (b) *Water*
 (c) *It weakens and destroys materials made of iron.*
 (d) (i) *Iron coated with zinc.*
 (ii) *For roofing houses.*
13. (a) (i) *Prevents re-entry of oxygen in the water.*
 (ii) *To drive oxygen from it.*
 (iii) *No observable change.*
 (iv) *Oiling, greasing, painting, alloying, plastic covering, galvanising, electroplating (any one).*
- (b) *See chapter 5 page 79.*
- (c) *Oiling, greasing, painting, alloying, plastic covering, galvanising, Electroplating*
14. (a) (i) *Pass air through lime water. The colourless lime water turns milky showing the presence of carbon dioxide.*
 (ii) *Pass air over a dry blue cobalt chloride paper. The blue paper turns pink.*
 or *Pass air over anhydrous copper (II) sulphate. The white anhydrous copper (II) sulphate turns blue showing the presence of water vapour.*
 (iii) *Pass air over heated copper metal. The brown copper metal turns black.*
- (b) *Test the apparatus for leaks by pulling air through from one syringe to the other. Then let the volumes read 0 cm³ and 100 cm³ and then set the apparatus as shown in the diagram below.*



- *Heat the glass tube vigorously and at the same time continuously pass air slowly from one syringe to the other through the hot copper. After 4-5 minutes, stop heating and cool the glass tube with a damp cloth.*
- *Record the volume of the gas in the two syringes.*
- *Repeat the procedure until no further change of the volume takes place.*
- *Record the final volume of the remaining gas.*

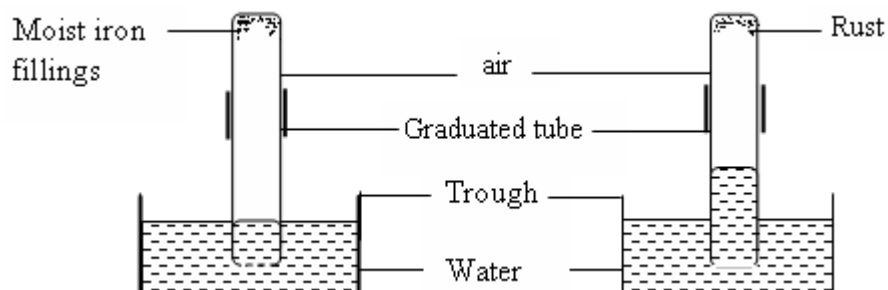
Calculation:

$$\text{Vol. of oxygen used to oxidize copper} = 100 - \text{Residual air}$$

$$\text{Percentage of oxygen in air} = \left(\frac{100 - \text{vol. of remaining air}}{100} \right) \times 100$$

If the experiment is accurately performed, the value obtained is approximately 21%.

or - Set up the experiment as shown below.



- Note the initial level of water.
- Leave the apparatus to stand for some days.
- Equalize the levels of water both inside and out side the tube to have atmospheric pressure inside.
- Note the final level.

$$\text{Percentage of oxygen in air} = \left(\frac{\text{Vol. of oxygen used}}{\text{Initial vol. of air}} \right) \times 100$$

- (c) Aid in breathing where natural oxygen supply is insufficient;
In the oxyacetylene flame which is used in welding and cutting metals.

15. (a) (i) - Its constituents can be separated by physical means.
- Proportions of the constituents are not fixed in every where.
- (ii) Nitrogen 78% and Carbon dioxide 0.03%.
- (ii) Sulphur dioxide - SO_2

CHAPTER 6

Self-Check 6

SECTION A

1. B.	2. C.	3. C.	4. A.	5. D.
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SECTION B

6. (a) (i) *Hygroscopic substance is a substance which absorbs water from the atmosphere.*
- (ii) *Deliquescent substance is a solid substance which absorbs water vapour from the atmosphere to form a solution.*
- (iii) *Deliquescence is the process of absorbing water from the atmosphere by a solid to form a solution.*
- (b) *Calcium oxide, Conc. Sulphuric acid, Phosphorous (V) oxide, Silicon (IV) oxide (Silica gel) and Anhydrous or Fused Calcium Chloride.*
7. (a) (i) *Hydrogen gas and ammonia gas.*
- (ii) *Carbon dioxide, Carbon monoxide, sulphur dioxide, chlorine.*
- (iii) *Carbon monoxide, Carbon dioxide, sulphur dioxide etc.*
- (b) (i) *Hydrogen gas and ammonia gas are lighter than air.*
- (ii) *Because the gases (Carbon dioxide, Carbon monoxide, sulphur dioxide, chlorine are denser than air.*
- (iii) *Carbon monoxide is insoluble in water; oxygen and carbon dioxide are slightly soluble in water.*

CHAPTER 7

Self-Check 7

SECTION A

1. B.	2. A.	3. A.	4. C.	5. A.	6. B.	7. C.	8. B.	9. A.	10. B.
11. B.	12. D.	13. B.	14. A.	15. C.	16. C.	17. B.	18. D.	19. B.	20. D.

SECTION B

21. (a) (i) - Name = *Oxygen*
- Chemical formula = *O₂*
- (ii) *It relights a glowing splint.*
- (iii) $2\text{H}_2\text{O}_2\text{(aq)} \rightarrow 2\text{H}_2\text{O(l)} + \text{O}_2\text{(g)}$
- (b) *Manganese (IV) oxide.*

22. (a) P: Manganese (IV) oxide (MnO_2)
Q: Oxygen
- (b) $2\text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$
- (c) (i) Sulphur continues to burn with a blue flame.
(ii) Magnesium burns with a brilliant flame leaving a white powder behind.
- (d) (i) $\text{S}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g})$
 $2\text{Mg}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2\text{MgO}(\text{s})$
(ii) Solution formed by products in c(i).
This is because the products (SO_2), dissolves in water forming sulphurous acid according to the equation:
- $$\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{SO}_3(\text{aq})$$
- (e) (i) Fractional distillation of liquid air or electrolysis of water.
(ii) Oxygen is used in - oxy-hydrogen flames or oxy-acetylene flames for cutting metal, - making steel and
- burning fuel in rockets.
- hospitals for respiration.
- Mountain climbing.
23. (a) (i) Basic (ii) Amphoteric (iii) Amphoteric
(iv) Basic (v) Neutral (vi) Acidic
24. (a) (i) See chapter 7 page 95.
(ii) $2\text{KClO}_3(\text{s}) \rightarrow 2\text{KCl}(\text{s}) + 3\text{O}_2(\text{g})$
- (b) (i) - Sodium burns with a yellow flame forming a pale yellow solid called sodium peroxide.
 $2\text{Na}(\text{s}) + \text{O}_2(\text{l}) \rightarrow \text{Na}_2\text{O}_2(\text{s})$
- The product is sodium peroxide which is highly soluble in water and dissolves to form an alkaline solution.
(ii) Magnesium ribbon burns brightly in oxygen to form a white solid called magnesium oxide.
 $2\text{Mg}(\text{s}) + \text{O}_2(\text{l}) \rightarrow 2\text{MgO}(\text{s})$
(iii) When iron wool is strongly heated in air or oxygen, it burns with brick red flame forming a black residue called tri-iron tetra oxide.
 $3\text{Fe}(\text{s}) + 2\text{O}_2(\text{l}) \rightarrow \text{Fe}_3\text{O}_4(\text{s})$
- (c) Photosynthesis

25. (a) X. - Manganese (IV) oxide.
- (b) $2\text{H}_2\text{O}_2 (\text{aq}) \rightarrow 2\text{H}_2\text{O} (\text{l}) + \text{O}_2 (\text{g})$
- (c) (i) To increase the rate of reaction.
- (ii) A catalyst and slightly warm temperature.
- (d) Photosynthesis

CHAPTER 8

Self-Check 8

SECTION A

1. D.	2. B.	3. C.	4. B.	5. C.	6. A.	7. B.	8. C.	9. A.	10.A.
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SECTION B

11. (a) (i) Granulated zinc.
- (ii) $\text{Zn} (\text{s}) + 2\text{HCl} (\text{aq}) \rightarrow \text{ZnCl}_2 (\text{aq}) + \text{H}_2 (\text{g})$
- (iii) Hydrogen is slightly soluble in water.
- (iv) Plunge a burning splint into a mixture of air with hydrogen.
It burns with a pop sound.
- (b) (i) Moisture (i.e. the gas is not dry), it is not hydrogen chloride fumes because this dissolve in water.
- (ii) By drying the gas (passing it through concentrated sulphuric acid).
- (iii) Copper (II) sulphate.
- (c) (i) Brown residue is left behind and water drops are formed on the sides of the apparatus.
- (ii) Oxidation
- (iii) $\text{CuO} (\text{s}) + \text{H}_2 (\text{g}) \rightarrow \text{Cu} (\text{s}) + \text{H}_2\text{O} (\text{g})$
12. (a) (i) Heated magnesium and steam.
- (ii) $\text{Mg} (\text{s}) + \text{H}_2\text{O} (\text{g}) \rightarrow \text{MgO} (\text{s}) + \text{H}_2 (\text{g})$
- (b) The white ash dissolved to form a colourless solution.
- (c) $\text{Mg}(\text{OH})_2 (\text{aq}) + 2\text{HCl} (\text{aq}) \rightarrow \text{MgCl}_2 (\text{aq}) + \text{H}_2\text{O} (\text{l})$

13. (a) (i) *P* - Magnesium oxide, MgO .
(ii) *Q* - Hydrogen (H_2).
(b) *Q* can be tested by plunging a burning splint into a mixture of *Q* and air. It burns with a pop sound.
(c) $\text{MgO} (s) + 2\text{HCl} (aq) \rightarrow \text{MgCl}_2 (aq) + \text{H}_2\text{O} (l)$
14. (a) *Magnesium*.
(b) (i) *A solid white ash*.
(ii) $\text{Mg} (s) + \text{H}_2\text{O} (g) \rightarrow \text{MgO} (s) + \text{H}_2 (g)$
(c) (i) *Hydrogen*.
(ii) $2\text{H}_2 (g) + \text{O}_2 (g) \rightarrow 2\text{H}_2\text{O} (g)$
15. (a) (i) *A solution is a homogenous mixture of two or more substances*.
(ii) *A saturated solution of a solute at a particular temperature is one which can dissolve no more solute at that temperature*.
(iii) *A super saturated solution is one that contains more of the solute than it could hold at that temperature if crystals of the solute were present*.
(iv) *A suspension is a liquid containing small particles of solid spread throughout it and the particles settle on standing*.
- (b) (i) *A saturated solution of a solute at a particular temperature is one which can dissolve no more solute at that temperature. While unsaturated solution is one which can dissolve more solute at the same temperature*.
(ii) *When no more sodium chloride can dissolve in the water*.
- (c) (i) - *A suspension contains solid particles which can be seen.*
- *Its solid particles settle on standing.*
- *Filtration separates it into a filtrate and residue.*
(ii) *A saturated solution of a solute at a particular temperature is one which can dissolve no more solute at that temperature. While super saturated solution is one that contains more of the solute than it could hold at that temperature if crystals of the solute were present.*

CHAPTER 9**Self-Check 9****SECTION A**

1. B.	2. A.	3. A.	4. C.	5. D.	6. D.	7. D.	8. A
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CHAPTER 10**Self-Check 10****SECTION A**

1. D	2. C	3. D	4. C	5. B
6. D	7. A	8. D	9. C	10. D
11. D	12. D	13. B	14. C	15. A
16. D	17. B	18. A	19. B	20. D
21. B.	22. B.	23. A.	24. A.	25. D.

SECTION B

26. (a) (i) $31 \quad (A = Z + n)$
 $= 15 + 16$
 $= 31$
(ii) 2:8:5
- (b) (i) XCl_3 (ii) Covalent bond. (Both X and Chlorine are non-metals)
- (c) It would turn blue litmus paper red or would have no effect on the red litmus paper.
27. (a) (i) Y (ii) X
- (b) 2:8:8:1
- (c) (i) 4
(ii) The electronic configuration of W = 2:4, since the number of electrons in the outermost shell 4)
28. (a) 2:8:3 (Atomic number of an element = Number of electrons for a neutral atom. Therefore, the number of electrons in Q = 13)
- (b) Group III.

- (c) *Q is a metal therefore, it would conduct electricity.*
- (d) (i) Q_2O_3
- (ii) *Electrovalent or ionic bond.*
29. (a) (i) $x = 3$
(Atomic no. of an element = No. of electrons for a neutral atom)
- (ii) $A = Z + n$
 $= 13 + 14$
 $= 27$
- (b) (i) A 2:6 (ii) A^{2-} 2:8
- (iii) C 2:8:3 (iv) C^{3+} 2:8
- (c) *A and B*
30. (a) (i) X. 2:8:2
- (ii) Y. 2:6
- (b) (i) *W is a solid at room temperature.* (ii) *Ionic or electrovalent.*

CHAPTER 10

Self-Check 10.2

SECTION A

1. A.	2. B.	3. A.
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SECTION B

4. (a) *Group II.*
- (b) (i) *M sunk to the bottom of the container dissolving with effervescence liberating a colourless gas that burns with a pop sound.*
- (ii) $M(s) + 2H_2O(l) \rightarrow M(OH)_2(aq) + H_2(g)$
- (c) *It strengthens bones.*
5. (a) (i) W_2O
- (ii) *The resultant solution is alkaline.*
W belongs to group I in the periodic Table. It is an alkali metal and the oxide is highly soluble in water to form an alkaline solution.
- (b) XZ_3 (c) W

6. (a) (i) *A and C (in metallic elements, ions are smaller than their parent atoms)*
 (ii) *B (ionic radius is larger than that of the parent atom.)*
 (iii) *Going down a group, the size of the atom as well as that of its ion increases. Hence C is more reactive and will have greater atomic number.*
- (b) (i) *Going from left to right of the period, each element has one more electron than the previous element. Notice that this extra electron goes into the same energy level (3rd in this case) and doesn't cause any increase in size of the atom. But, the extra positive charge on the nucleus, caused by the extra proton (going from left to right, atomic number increases by 1), increases the attraction of the electrons and makes atoms slightly smaller across the period.*
- (ii) *Oxides of metals are basic or amphoteric (aluminum oxide; metal having higher valence); non-metal oxides are acidic with low melting points –most of them are gasses. SiO₂ has a giant structure and is a solid with very high melting point. It is an amphoteric oxide.*
- (iii) - Na₂O and MgO are basic;
 - Al₂O₃ and SiO₂ are amphoteric and
 - P₂O₅ is acidic.
- (c) - SO₂, and SO₃;
 - Cl₂O, Cl₂O₆ and Cl₂O₇
- (d) *These two ions have the same number of electrons, in the same shells but the greater nuclear charge in K⁺ ions pulls the electrons closer to the nucleus and it is therefore, smaller than that of Cl⁻ ions.*

CHAPTER 11

Self-Check 11

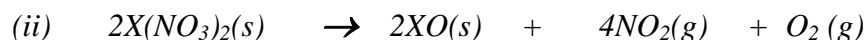
SECTION A

1. A	2. B	3. B	4. A	5. B
6. A	7. C	8. C	9. B	10. B
11. C	12. C	13. B	14. B	15. A

SECTION B

16. (a) V (b) T (c) Q₃S₂ (d) Neutral. (e) P and Q
17. (a) (i) Ionic or Electrovalent bond
 (ii) X²⁺.

(b) (i) Radish brown fumes were seen and solid substance formed.



18. (a) (i) TQ_3 . (ii) E_2X .

(b) State the type of bonding;

(i) Metallic bond.

(ii) Covalent bond.

(iii) Ionic or electrovalent bond.

19. An atom X of an element, atomic mass 31 contains 15 protons.

(a) (i) 16 (ii) 2 : 8 : 5 (b) V

(c) (i) XCl_3 (ii) Covalent bond. (d) Isotope

20. (a) S (b) (i) PT_2 . (ii) Covalent bond.

(c) (i) Q

(ii) It belongs to group VIII (inert gases), therefore has fully filled outer shell.

(d) (i) SR or ST

(ii) The type of bond is ionic. Therefore, in water it would be an excellent electrolyte.

CHAPTER 12

Self-Check 12.1

1. B	2. D	3. A	4. C	5. D	6. D	7. D	8. B	9. D	10. A
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Self-Check 12.2

1. B.	2. C.	3. A.	4. C.	5. D.	6. D.	7. D.	8. D.	9. A.	10. B.
11. B.	12. A.	13. C.	14. B.	15. C.	16. B.	17. D.	18. B.	19. D.	20. D.
21. A.	22. A.	23. B.	24. C.	25. C.	26. C.	27. B.	28. A.	29. B.	30. C.

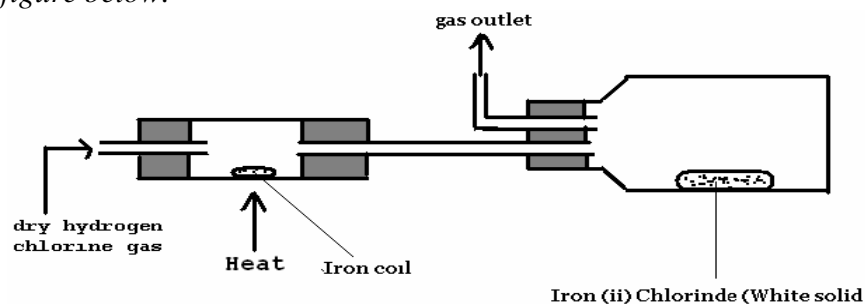
SECTION B

31. (a) (i) A normal salt is one where all the hydrogen ions in the acid have been replaced by a metallic ion.

Examples: Na_2SO_4 , $NaCl$, $Mg(NO_3)_2$, $CaCO_3$ etc.

- (ii) An acid salt is one in which not all the hydrogen ions in the acid have been replaced by metallic ions.
Examples: NaHSO_4 , NaHCO_3 , $\text{Mg}(\text{HCO}_3)_2$ etc.
- (b) (i) Copper (II) carbonate.
(ii) White solid turned black.
(iii) $\text{CuCO}_3(\text{s}) \rightarrow \text{CuO}(\text{s}) + \text{CO}_2(\text{g})$
- (c) White precipitate insoluble in the acid formed.
32. (a) Solubility is the amount of solute dissolved in 100 cm^3 of water at a given temperature.
- (b) See the notes on page 200.
- (c) For (i) and (ii), draw the graph and obtain the answers from it!!
(iii) Potassium nitrate.
(iv) White precipitates of Potassium chloride would be seen in the solution.
- (d) Separation of mixture of salts.
33. (a)
- Add lead (II) oxide into a known volume of dilute sulphuric acid in a beaker a little at a time while stirring with a glass rod until no more of the solid dissolves. (This shows that all the acid is neutralized).
 $\text{PbO}(\text{s}) + 2\text{HNO}_3(\text{aq}) \rightarrow \text{Pb}(\text{NO}_3)_2(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 - Filter off the excess lead (II) oxide.
 - Saturate the filtrate by evaporation and cool the solution to crystallize.
 - Filter and wash the crystals with distilled water.
 - Dry the crystals using filter paper or sunlight.
- (b) - It decomposes with a crackling sound giving brown fumes of a gas.
- Leaving a reddish brown solid which turns to yellow on cooling.
 $\text{Pb}(\text{NO}_3)_2(\text{s}) \rightarrow \text{PbO}(\text{s}) + 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$
34. (a)
- Add copper (II) oxide into a known volume of dilute sulphuric acid in a beaker a little at a time while stirring with a glass rod until no more of the solid dissolves. (This shows that all the acid is neutralized).
 - $\text{CuO}(\text{s}) + 2\text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CuSO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 - Filter off the excess copper (II) oxide.
 - Saturate the filtrate by evaporation and cool the solution to crystallize.
 - Filter and wash the crystals with distilled water.
 - Dry the crystals using filter paper or sunlight.
- (b) $\text{CuSO}_4(\text{s}) \rightarrow \text{CuO}(\text{s}) + \text{SO}_3(\text{g})$

35. (a) (i) Chlorine
- (ii) - Chlorine must be dry
 - Iron must be heated strongly
 - Experiment can be carried out in a fume cupboard as chlorine is poisonous
 - Heating can be stopped as soon as iron starts to react.
- (iii) Iron glows red as it reacts with the chlorine, black crystals sublime into the bottle.
- (iv) $2\text{Fe} (s) + 3\text{Cl}_2 (g) \rightarrow 2\text{FeCl}_3 (s)$.
- (b) Coil rust free wire around a pencil. Place the coil in a combustion tube as in the figure below.



- Pass dry Hydrogen Chloride gas to sweep air out of the apparatus.
 - Heat the iron strongly and pass the hydrogen chloride over it.
 - Stop heating as soon as the iron starts to react.
 - Warm the end of the combustion tube near the receiver.
- White solid of anhydrous iron (II) Chloride is formed.



Self-Check 12.3 SECTION A

1. C.	2. D.	3. A.	4. B.	5. D.	6. D.	7. D.	8. B.	9. C.	10. C.
11. C.	12. B.	13. D.	14. A.	15. A.	16. A.	17. D.	18. B.	19. B.	20. B.

21. (a) (i) Calcium sulphate or Magnesium sulphate.
 (ii) Calcium hydrogen carbonate or Magnesium hydrogen carbonate.
- (b) - It wastes soap.
 - It leads to waste of fuel since it forms fur and boiler scales, which are bad conductors of heat.
 - The boiler scale may block the boilers hence leading to damage of them.
- (c) (i) $\text{Na}_2\text{Y} (a) + \text{MgSO}_4 (aq) \rightarrow \text{Na}_2\text{SO}_4 (aq) + \text{MgY} (s)$
 or $\text{Na}_2\text{Y} (a) + \text{CaSO}_4 (aq) \rightarrow \text{Na}_2\text{SO}_4 (aq) + \text{CaY} (s)$
 (ii) Ion exchange.

22. (a) (i) *P* Because the volume of the soap solution required is the smallest.
- (ii) *R* Because after boiling, the volume of soap solution required decreased from 5 cm³ to 3 cm³.
- (iii) *Q* because the volume of soap solution required is the greatest (8 cm³) and the volume did not change.
- (b) Calcium sulphate or Magnesium sulphate

CHAPTER 13

Self-Check 13

SECTION A

1. D.	2. B.	3. D.	4. C.	5. B.	6. D.	7. B.	8. C	9. B.	10. D.
11. D.	12. D.	13. D.	14. D.	15. A.	16. B.	17. A.	18. A.	19. C.	20. D.

SECTION B

21. (a) Graphite and diamond.
- (b) Any two of the following:
- Graphite conducts electricity while diamond does not conduct electricity.
 - Graphite is soft while diamond is hard.
 - Graphite is opaque while diamond is transparent.
- (c) **Graphite** Any **one** of the following uses:
- It is used as an electrode in electrolysis.
 - It is used as lubricant.
 - It is used for making pencils.
- Diamond** Any **one** of the following uses:
- It is used for making cutting tools.
 - It is used for making drilling tools
 - It is used for making abrasive tools.
 - It is used for making jewellery.
22. (a) (i) Allotropes are two or more structural forms of the same element existing in the same physical state.
- (ii) Sulphur or Phosphorous.
- (b) (i) Graphite, diamond and amorphous carbon.
- (ii) Diamond
- It is the hardest substance known.
 - It is transparent and sparkles.
- or Graphite
- It is soft.
 - It has mobile electron and therefore conducts electricity.

(iii) *Diamond*

- As a result of its hardness diamond is used for making cutting and drilling tools.

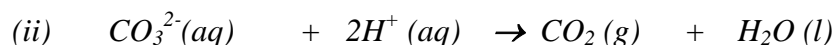
- Due to its transparency and sparkling property, it is used for making jewellery.

or *Graphite*

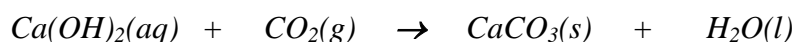
- Due to its softness, graphite is used as a lubricant.

- As a result of having mobile electrons, it is used as electrodes in electrolysis and dry cells.

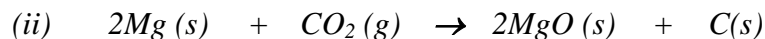
23. (a) (i) **See page 223.**



(b) *When carbon dioxide is passed through a solution of calcium hydroxide, the solution first turns milky due to the formation of a white precipitate of calcium carbonate according to the equation.*



(c) (i) *When a piece of burning magnesium ribbon is lowered into a jar of Carbon dioxide, it continues to burn for a short time with a splattering flame, and black specks of carbon form on the sides of the jar. The magnesium burns to form white magnesium oxide.*



(d) *The litmus solution turned blue.*

24. (a) *Carbon dioxide reacts with Sodium hydroxide to form soluble sodium Carbonate and water.*

On further bubbling, the carbonate formed reacts with the carbon dioxide to form white precipitate of sodium hydrogen carbonate. according to the equation:

