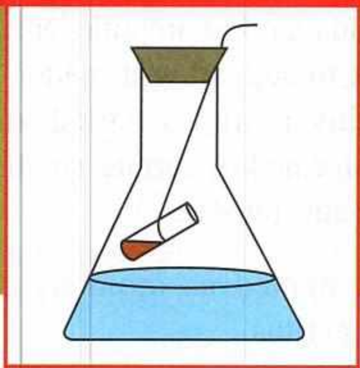
CONTENTS



Matter

1

Theme : Matter is composed of atoms/molecules and is found as solid, liquid and gas. These states are changed on the basis of interparticle space and interparticle collisions. Kinetic theory of matter has been postulated to explain the change of state. In a physical and chemical change, the total mass before and after the change remains the same which is known as the law of conservation of mass, which would help in understanding the behaviour of matter.

In this chapter you will learn :

- *Nature of matter.*
- *Main postulates of kinetic theory of matter.*
- *Interparticle space, interparticle attraction and collision.*
- *Law of conservation of mass (statement and explanation with examples).*
- *Interconversion of states of matter.*

LEARNING OUTCOMES

The children will be able to :

- describe the main postulates of kinetic theory of matter.
- explain the reason of change of one state of matter into another and vice-versa.
- define and explain the law of conservation of mass using an example.

INTRODUCTION

We know that chemistry is the branch of science that deals with the study of various kinds of **matter**, which includes their structure, composition and physical and chemical properties.

If we look around us, we see a number of things like, stones, sand, clay, water, minerals, plants, animals, pens, pencils, books, etc. They are all made up of different types of materials and are called **"matter"** in science.

In fact, matter includes all living and non-living things of which the universe is composed of.

WHAT IS MATTER ?

"Matter is anything which has mass, occupies space and can be perceived by our senses."

> There are different kinds of matter.
> - A book is made up of paper.
> - A bag is made up of plastic.
> - A dress is made up of cloth.
> - A hammer is made up of iron.

Air which we cannot see is also a kind of matter because we can feel it.

- If you blow air in a balloon, it gets inflated because air occupies space.
- If you switch on the fan, you feel air due to the movement of fan.

Thus, air, water, sugar, sand, oxygen, hydrogen, steel, coal, iron, wood, alcohol, milk, oil, etc. are all different kinds of matter because they have mass and occupy space.

MATTER OFFERS RESISTANCE TOO

If you try to walk during an air storm or lift a big stone, you experience some resistance.

WHAT IS MATTER MADE UP OF ? [NATURE OF MATTER]

Ancient philosophers had different beliefs about the nature of matter.

Greek philosophers believed that all kinds of matter are made up of fire, water, air and earth. While ancient Indian philosophers believed that all kinds of matter are made up of five elements — sky, air, fire, water and earth.

Maharshi Kannada, an Indian philosopher was perhaps the first to suggest that matter is made up of very tiny particles called anu (molecule) which is formed of further smaller particles called parmanu (atom).

> Parmanu was named as atom by a Greek thinker Democritus.

Later on, John Dalton also suggested that all kinds of matter are made up of extremely small particles called atoms.

An atom is the smallest particle of matter that exhibits all the properties of that matter. They usually do not have independent existence, therefore they combine with each other to form small particles called molecules. These molecules have independent existence.

Both these particles (atoms and molecules), are too small to be seen through a naked eye or through an ordinary microscope.

However, the number of particles in matter can be very large. *Example :* a drop of water contains 10^{21} particles of water.

CHARACTERISTICS OF MATTER

1. **Particles of matter are very small :** To show this, the following experiment is carried out.

Activity 1

Dissolve two or three crystals of blue vitriol (copper sulphate pentahydrate) in about 10 ml of water to get a clear transparent blue solution. Take four beakers and label them as

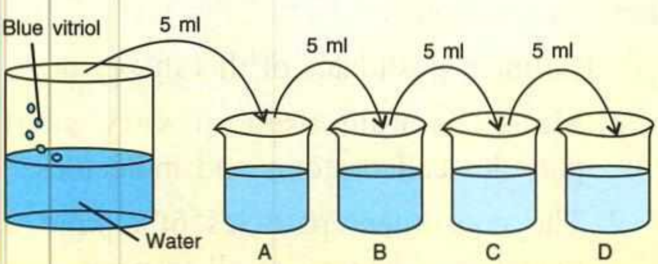
A, B, C and D. Fill each beaker with 50 ml of water. Now transfer 5 ml of solution to beaker A and stir it properly to get a uniform blue colour. Take 5 ml of solution from beaker A, transfer it to beaker B and stir well. Again transfer 5 ml of solution from B to C and then from C to D.

What do you observe ?

The solutions in all the beakers are coloured though they become fainter due to successive dilution.

Thus it is concluded that a small crystal of blue vitriol contains a very large number of tiny particles which show all the properties of the substance.

The whole process can be repeated for potassium permanganate crystals or ink to prove the nature of a particle.



The colour of the solution gets fainter with successive dilution.

2. Particles of matter have interparticle space between them.

Activity 2

Take 100 ml of water in a graduated cylinder and add 20 grams of sodium chloride to it. Dissolve the salt in water by proper stirring with the help of a glass rod. When all the salt has dissolved in water, you will notice that there is no increase in the level of water in the cylinder.

This shows that there must be some **space between the particles of water** in which the salt particles get accomodated when dissolved.

The space between the particles is called **interparticular space** or **intermolecular space**.

3. Particles of matter are in constant random motion : This can be proved by the following :

Brownian motion : Robert Brown gave the evidence for the existence and movement of particles in liquids. He suspended some pollen grains in water and looked into the water through a magnifying glass. He observed that the pollen grains were moving throughout the water in a zig-zag or irregular manner.

Why were the pollen grains moving in an irregular manner ?

This is because water is made up of tiny particles which are also in random motion. The pollen grains move in such a way because they collide with the moving particles of water.

This haphazard, random motion of suspended particles on the surface of a liquid or in air is called **Brownian motion**.

Since this phenomenon was first noticed by Robert Brown, it is called as Brownian motion.

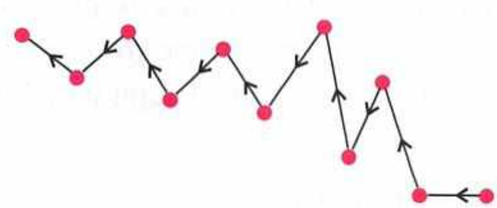
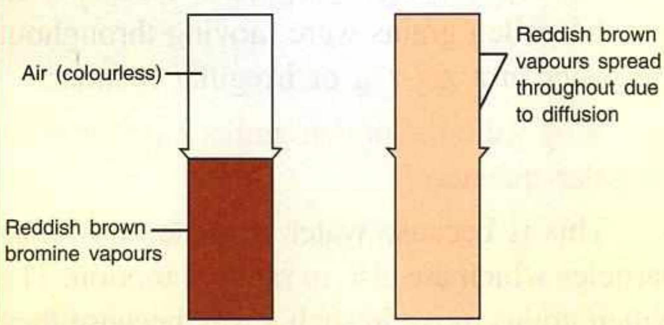


Fig. 1.1 Brownian motion

Diffusion : The intermixing of two or more substances due to the motion of their particles in order to get a uniform mixture is called **'diffusion'**. The rate of diffusion is the fastest in gases and the slowest in solids. It increases with an increase in temperature.

Activity 3

Take a gas jar full of bromine vapours (reddish brown) and invert another gas jar containing air over it. It is observed that after sometime, the reddish brown vapours of bromine also spread out into the upper jar. This mixing is called diffusion.



4. **Particles of matter attract each other :** There exists a force of attraction between the particles of matter which keeps them together. But the magnitude of force varies from one type of matter to another.

For example : A piece of chalk can be broken into pieces very easily while a piece of coal requires a greater force to break and a metal piece of copper or steel cannot be broken easily.

STATES OF MATTER

The magnitude of intermolecular force of attraction, intermolecular space and random motion of molecules of matter lead to the three main states of matter :

Solids, liquids and gases

Solids : A solid has a definite shape and a definite volume. *Example :* wood, stone, iron, ice, etc.

Liquids : A liquid has a definite volume but no definite shape. *Example :* water, alcohol, mustard oil, fruit juice, milk, etc.

Gases : A gas has neither a definite shape nor a definite volume. *Example :* air, hydrogen, oxygen, water vapour, etc.

KINETIC THEORY OF MATTER

The theory stating that any substance whether solid, liquid or gas is made up of tiny particles called atoms, molecules or ions which are in constant motion is called **"kinetic theory of matter"**.

The main postulates of this theory are :

1. Matter is composed of very small particles called atoms and molecules.
2. The constituent particles of a kind of matter are identical in all respects.
3. These particles have spaces or gaps between them which are known as interparticular or intermolecular spaces.
4. There exists a force of attraction between the particles of matter which holds them together. This force of attraction is known as interparticular or intermolecular force of attraction.
5. Particles of matter are always in a state of random motion and possess kinetic energy, which increases with an increase in temperature and vice-versa.

EXPLANATION OF STATES OF MATTER BASED ON KINETIC THEORY

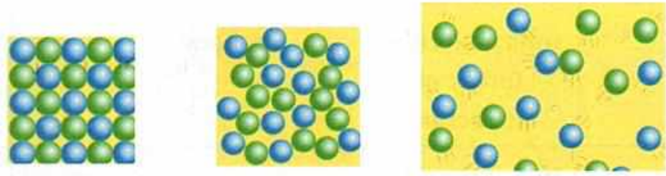
Solid state : In solids, the particles (molecules) are closely packed. There is a strong force of attraction between the particles (molecules) and the intermolecular space is almost negligible. The molecules are therefore not free to move. They only vibrate about their mean positions.

This makes solids hard and rigid and difficult to compress, giving them a fixed shape and size. Solids have low kinetic energy.

Liquid state : In case of liquids, the particles are not very closely packed. The intermolecular forces of attraction are not as strong as in the case of solids, thus the intermolecular spaces are larger. The molecules are loosely packed, hence liquids are more compressible. The particles are able to move freely and randomly. The kinetic energy is higher than that of solids.

This makes a liquid flow and take the shape of the container into which it is poured. Thus, liquids have a fixed volume but no definite shape.

Gaseous state : In case of gases, the intermolecular forces of attraction are negligible. The particles lie far apart from each other and the intermolecular spaces are therefore, very large. Hence they can be compressed. The forces of attraction are so weak that the particles of gases are free to move within the entire space available to them. They have high kinetic energy. During motion, these particles collide with each other and also with the walls of the container. As a result, gases have neither a fixed shape nor a fixed volume. They completely fill up the space available to them.



(a) Molecules lie closely packed together, with a great force of attraction between them.

(b) Molecules lie fairly less apart from each other, with a little force of attraction between them.

(c) Molecules lie very far from each other with a very little force of attraction between them.

Fig. 1.2 The position of molecules in the (a) solid, (b) liquid and (c) gaseous states.

Note : All substances that can flow are called **fluids**. Both gases and liquids are fluids. If a container having a liquid or a gas is opened, they both can flow out of the container. Gases can flow in all directions freely, but liquids cannot flow upwards or against the gravity on their own.

Table 1.1 Properties of solids, liquids and gases.

S.No.	Property	Solids	Liquids	Gases
1.	Mass	Definite	Definite	Definite
2.	Shape	Definite	No fixed shape.	No fixed shape.
3.	Volume	Definite	Definite	Indefinite
4.	Packing of particles	Very closely packed.	Less closely packed.	Least closely packed.

5.	**Intermolecular forces of attraction**	Strongest	Weaker than in solids	Weakest
6.	**Intermolecular space**	Least	More	Most
7.	**Fluidity**	Not possible	Can flow	Can flow
8.	**Compressibility**	Not possible	Negligible	Highly compressible
9.	**Rigidity**	Highly rigid	Less rigid	Not rigid
10.	**Kinetic energy**	Lowest	Higher	Highest

INTERCONVERSION OF STATES OF MATTER

"The phenomenon of change of one state of matter into another and then back to the original state, without any change in its chemical composition is called **interconversion of the states of matter**.

The change in the state of matter is caused by –

(i) change in temperature

(ii) change in pressure

Change of state of matter by changing the temperature

When a substance in solid state is heated, it changes into liquid state after sometime. On further heating, the liquid changes into gaseous state. On cooling, just the reverse happens.

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

This can be explained by the following examples :

- Water is a liquid under ordinary conditions but when it is kept in a deep freezer, it gets cooled and changes into ice at 0°C. Ice when kept at room temperature, again changes back into liquid water.

Similarly water on heating changes into steam at 100°C, which on cooling changes back into liquid water. But there is no change in the chemical composition of water when its state changes from liquid to solid or liquid to gaseous state.

$$\underset{\text{(solid)}}{\text{Ice}} \underset{\text{Cool}}{\overset{\text{Heat}}{\rightleftharpoons}} \underset{\text{(liquid)}}{\text{Water}} \underset{\text{Cool}}{\overset{\text{Heat}}{\rightleftharpoons}} \underset{\text{(gas)}}{\text{Steam}}$$

Another example is candle wax. When candle is burnt, the solid wax melts into liquid wax. On cooling, the molten wax again changes back into solid wax.

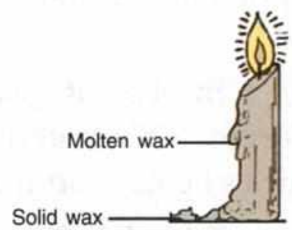


Fig. 1.3 Interconversion of the states of matter (wax) in a candle

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

How and why do the changes take place ?

This can be explained as follows : When a substance in solid state is heated, its constituent particles gain kinetic energy and start vibrating more vigorously. Eventually, a stage is reached at which particles gain enough energy to overcome the forces of attraction between them and they start moving, thus changing into liquid state from the solid state. On further heating, more kinetic energy is acquired, the forces of attraction become almost negligible, increasing the intermolecular spaces and ultimately the liquid state of matter changes into its gaseous state.

When cooling is done, the kinetic energy decreases. Thus, the spaces between the particles decrease and the forces of attraction increase. This also affects the movements of particles. Eventually, the gaseous state changes into liquid state which on further cooling changes into solid state.

There are some substances that directly change from the solid state to the gaseous state without passing through the liquid state. This process is called **sublimation** and such substances are called sublimable substances.

$$\text{Solid state} \underset{\text{Cool}}{\overset{\text{Heat}}{\rightleftharpoons}} \text{Gaseous state}$$

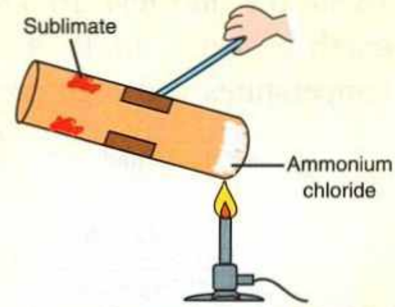
Example : Camphor, iodine, naphthalene, ammonium chloride, dry ice (solid carbon dioxide), etc.

Note : Naphthalene balls are used in bathrooms, wardrobes, etc. to keep the pests away. With the passage of time, they become smaller because they sublime and change into vapour state.

Activity 4

To show that ammonium chloride sublimes on heating.

Take a dry test tube. Hold it with a test tube holder and take a pinch of ammonium chloride in it. Heat the test tube.



Sublimation of ammonium chloride

What do you observe ?

After sometime, the solid ammonium chloride directly changes into white vapour, without changing into its liquid state. The vapour rises and solidifies on the upper cooler part of the test tube. This solid is called the sublimate.

This proves that there are substances which can change from solid to gaseous state on heating and vice versa. The process of changing directly from the gaseous to solid state without passing through the liquid state is called **deposition**.

Do all substances change their states ? Give four examples of substances which do not change their state on heating.

Change of state of matter by changing the pressure

Pressure is also one of the important factors for the change in the state of matter.

A gas can be changed into a liquid, and then into solid, by cooling it and by increasing its pressure. Thus, we can obtain liquid oxygen, liquid hydrogen, etc.

- Air contains mostly nitrogen and oxygen gases. When pressure is increased and temperature is decreased, air changes into its liquid state.
- LPG cylinders contain cooking gas in liquid state at high pressure (under normal conditions, it is a gas).
- Even the hardest rocks under the earth's crust melt at very high temperatures and high pressures.

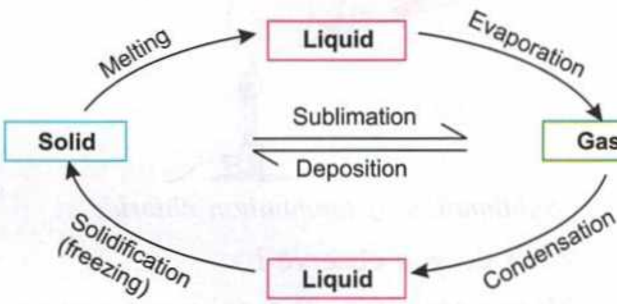


Fig. 1.4 Interconversion of states of matter

TERMS RELATED TO INTERCONVERSION OF STATES OF MATTER

Melting or fusion : The process by which a substance changes from solid state to liquid state is called melting or fusion.

Melting point : The fixed temperature at which a solid changes into a liquid upon heating at atmospheric pressure is called its melting point. The temperature remains constant till whole of the solid changes into liquid. Different substances have different melting points in their pure state. *For example,* melting point of ice is 0°C.

Evaporation : The process by which a substance changes from liquid state to vapour state is called evaporation or vaporisation. Evaporation takes place even at room temperature but it becomes faster on heating.

Boiling point : The fixed temperature at which a liquid starts changing into gaseous state upon heating is called its boiling point. At this temperature, the pressure of the liquid becomes equal to the atmospheric pressure. The temperature remains constant till whole of the liquid changes into vapour state. *Example :* The boiling point of pure water is 100°C.

Condensation : The process by which a substance in gaseous state changes into its liquid state is called condensation or liquefaction.

Condensation point : It is the temperature at which a gas starts changing into its liquid state upon cooling. *Example :* Condensation point of steam is 100°C.

Liquefaction : The process by which a gas is changed into its liquid state by applying pressure and lowering the temperature is called liquefaction.

Freezing : The process by which a substance in liquid state changes into a solid state is called freezing or solidification.

Freezing point : The temperature at which a liquid starts changing into its solid state upon cooling is called its freezing point. *Example :* Pure water freezes at 0°C under normal conditions.

Sublimation : The process by which certain substances change directly from solid to gaseous state on heating is called sublimation. *Example :* Camphor, ammonium chloride, iodine naphthalene, etc.

LAW OF CONSERVATION OF MASS

Lavoisier proposed the law of conservation of mass which states that —

"Matter can neither be created nor be destroyed in a chemical reaction".

However, it may change from one form to another during the reaction process.

It can also be stated as —

"In a chemical reaction, the total mass of the reactants is equal to the total mass of the products".

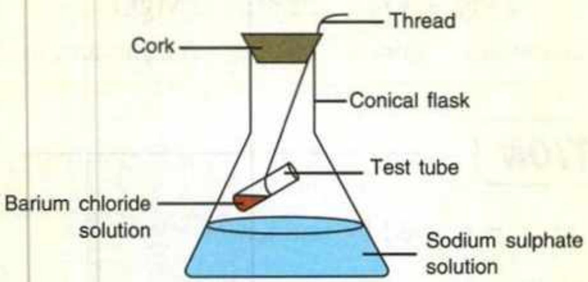
That means there is no change in the mass during a chemical reaction.

Experimental verification of the law of conservation of mass : This can be done by the following activities.

Activity 5

Things required : a conical flask, 10 ml test tube, two measuring cylinders, thread, cork, weighing balance, barium chloride solution and sodium sulphate solution.

$$\underset{(aq)}{BaCl_2} + \underset{(aq)}{Na_2SO_4} \rightarrow \underset{(ppt)}{BaSO_4} + \underset{(aq)}{2NaCl}$$



Experimental verification of law of conservation of mass

Take 5 ml of sodium sulphate solution in a measuring cylinder and pour into a concial flask. Take 5 ml of barium chloride solution in another measuring cylinder and pour into a 10 ml test tube. Tie the test tube with a thread and hang the test tube in the flask carefully so that the two solutions do not mix with each other. Put a cork in the mouth of the test tube so that the thread holding the test tube is held firmly in place.

Now carefully weigh the concial flask on a weighing balance and note the reading to get the mass of the concial flask and the substances.

Now tilt and twist the flask so that barium chloride solution mixes with sodium sulphate solution.

You will observe that a white insoluble solid (precipitate) of barium sulphate is formed along with a solution of sodium chloride. Wait for ten minutes to complete the reaction and the solid formed to settle down.

Weigh the contents again and note the reading.

You will observe that —

total mass of the apparatus + reactants

= total mass of the apparatus + products

Hence, the law of conservation of mass is verified.

It is clear from the above experiment that there is no change in the mass before and after a reaction. However, the substances undergo physical and chemical changes.

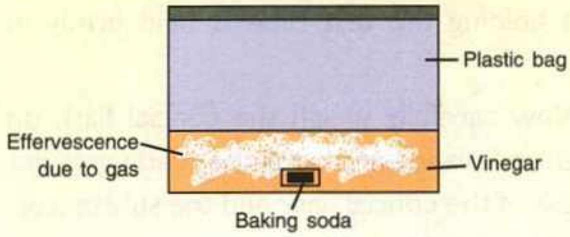
Activity 6

Things required : A plastic bag with zip lock, vinegar, baking soda and a weighing balance.

Take some baking soda on a paper. Take 1/4th cup of vinegar in the plastic bag. Weigh them separately.

Let the mass of plastic bag and vinegar + baking soda and paper = *W* g.

Now carefully put the baking soda with paper in the plastic bag and lock the zip of the



Experimental verification of law of conservation of mass

bag immediatley so that no gas escapes. Now shake the bag so that baking soda and vinegar come in contact. What do you observe now ?

A strong effervescence or fizz sound occurs indicating the formation of carbon dioxide gas. Wait till the fizz sound persists. Now weigh the plastic bag again. It is approximately W g.

Hence total mass of the reactants

= total mass of the products.

The reaction is, baking soda + vineger

= carbon dioxide + sodium acetate + water

$NaHCO_3 + CH_3COOH \rightarrow CO_2 + CH_3COONa + H_2O$

Note : To prove the law of conservation of mass, it is necessary to carry out the experiment in a closed container, otherwise if a product is gaseous, it may escape and the desired result will not be obtained.

- When wood is burnt in air, ash is formed. The mass of ash is less than that of wood. This is because the mass of air before the reaction and the mass of gaseous products formed after the reaction are not taken into consideration.
- Similarly when magnesium ribbon is burnt in air, a white solid — magnesium oxide is formed. The mass of magnesium oxide is more than the mass of magnesium. This is because the mass of oxygen used is not considered. If that is considered, the total mass of the reactants and the products is found to be almost equal.

$$\underset{\text{(magnesium)}}{2Mg} + \underset{\text{(air)}}{O_2} \xrightarrow{\text{heat}} \underset{\text{(magnesium oxide)}}{2MgO}$$

RECAPITULATION

- Matter is anything that has mass, occupies space and can be perceived by our senses.
- Matter consists of tiny particles called atoms and molecules.
- There is an intermolecular force of attraction between the particles of matter.
- The gaps between the particles are called intermolecular spaces.
- Particles of matter are always in random motion and they collide with each other.
- Matter exists in three states : solids, liquids and gases.
- Matter can change from one state to another when temperature and pressure are changed.
- The phenomenon of change of one state of matter into another and vice versa is called interconversion of states of matter.
- Matter can neither be created nor be destroyed. Only it can be changed from one form to another during a chemical reaction. This is known as law of conservation of mass.

EXERCISE

1. Define :
   (a) matter
   (b) intermolecular forces of attraction
2. What are the three states of matter ? Define each of them with two examples.
3. Explain interconversion of states of matter. What are the two factors responsible for the change of state of matter ?
4. State the main postulates of kinetic theory of matter.
5. What happens to water if
   (a) it is kept in a deep freezer ?
   (b) it is heated ?
   Explain the phenomenon of change of state of water.
6. (a) State the law of conservation of mass.
   (b) What do you observe when barium chloride solution is mixed with sodium sulphate solution ?
7. Give reasons :
   (a) A gas can fill the whole vessel in which it is enclosed.
   (b) Solids cannot be compressed.
   (c) Liquids can flow.
   (d) When magnesium is burnt in air, there is an increase in mass after the reaction.
8. Fill in the blanks :
   (a) The change of a solid into a liquid is called ....
   (b) The process in which a solid directly changes into gas is called ....
   (c) The change of water vapour into water is called ....
   (d) The temperature at which a liquid starts changing into its vapour state is ....
9. Give two examples for each of the following :
   (a) The substances which sublime.
   (b) The substances which do not change their state on heating.
10. Define :
    (a) Diffusion. (b) Brownian motion.
11. When sodium chloride is added to a definite volume of water and stirred well, a solution is formed, but there is no increase in the level of water. Why ?
12. What do you observe when a gas jar which appears empty is inverted over a gas jar containing Bromine vapours ? Name the phenomenon.
13. Why can a piece of chalk be broken easily into smaller pieces while a coal piece cannot be broken easily ?

Physical and Chemical Changes

2

Theme : There are different types of changes in our surroundings which are slow/fast, reversible/irreversible, periodic/non-periodic and physical/chemical. In a physical change, no new substance is formed while in a chemical change, a new substance with properties different from the elements forming that substance is formed. Learning of these changes helps in developing different scientific skills.

In this chapter you will :

- *Revise and review various types of changes which you have studied in previous classes.*
- *Learn physical and chemical changes, classification with examples.*

LEARNING OUTCOMES

The children will be able to :

- illustrate different changes occurring in nature with examples learned in previous classes.
- perform some activities to show some well known changes.
- differentiate between physical and chemical changes and classify the changes.

INTRODUCTION

Change is a universal phenomenon. Almost all substances undergo change. While some changes are easy to detect, some are so small that they are difficult to identify.

These changes have been taking place around us continuously, such as day and night, change of seasons, growth of plants and animals, ripening of fruits, burning of fuel, melting of ice, cooking of food, etc.

The change can be, natural and man-made, slow and fast, periodic and non-periodic, desirable and undesirable, small and large, reversible and irreversible, temporary and permanent, etc.

All these changes are exciting and each one of them has a reason behind it. Since most of the changes are different in nature, it is necessary to study them in detail.

SLOW AND FAST CHANGES

The changes that take longer time to complete are called **slow changes**. They take hours, days, months or even years to occur. *Examples :* Rusting of iron, change of seasons, formation of curd from milk, a child growing into an adult, formation of fossil fuels from dead plants and animals, etc.

Fig. 2.1 Rusting of iron

The changes that take place in a very short interval of time are called **fast changes**.

Examples : Bursting of a cracker, lighting of an electric bulb, switching on a fan, blinking of eyes, cutting of an apple, etc.

NATURAL AND MAN-MADE CHANGES

Some changes that take place in nature by themselves, are called **natural changes**.

Examples : Change of day and night, growing of a tree, earthquakes, eruption of volcanoes, etc.

Any change that occurs due to the efforts of human beings is called a **man-made change**.

Examples : Cooking of food, formation of steel from iron, brass from copper and zinc, jewellery from gold, etc.

COOKING OF FOOD

JEWELLERY FROM GOLD

Fig. 2.2

PERIODIC AND NON-PERIODIC CHANGES

Changes that are repeated at regular intervals of time are called periodic changes.

Examples : Change of day and night, change of seasons, etc.

Changes that are not repeated at regular intervals are called non-periodic changes. They occur at any time.

Examples : Earthquakes, landslides, an epidemic, a person becoming sick, etc.

REVERSIBLE AND IRREVERSIBLE CHANGES

When a change in a substance can be reversed by changing the conditions, it is said to be a reversible change.

Examples : Water freezing into ice on cooling and ice melting into water at room temperature.

$$\text{Water} \underset{\text{melting}}{\overset{\text{freezing}}{\rightleftharpoons}} \text{Ice}$$

If a substance cannot be brought back to its original state after a change, it is said to be an irreversible change.

Example : Burning a piece of paper into ash, cooking of food, formation of curd from milk, etc.

Activity 1

Write two examples for each of the following changes :

- Periodic
- Non-periodic
- Reversible
- Irreversible
- Slow
- Fast

In science, all possible changes can be classified into two broad categories.

1. Physical changes 2. Chemical changes

PHYSICAL CHANGE : CHANGE OF STATE

A physical change is one in which a substance alters temporarily in some or all of its physical properties, *viz.* state, shape, size, appearance, *etc.* but not in its chemical composition.

Characteristics of a physical change :

(i) *No new substance is formed.*

You might have observed that when a cube of ice is taken out of a refrigerator, it melts into water. If this water is kept back in the refrigerator, it re-freezes into ice. This indicates that the properties of water and ice are the same, *i.e.* their chemical composition is the same. They both have the formula H_2O. On melting of ice or on freezing of water, no new substance is formed. Only the physical state of the substance changes.

(ii) *The change is temporary and reversible, i.e. the substance returns to its original state on changing the conditions.*

Example 1 : You have noted that when a cube of ice is taken out of a refrigerator, it melts and changes into water. It turns again into ice when put back in the refrigerator. This shows that on removing the cause of change, the substance returns to its original state.

Example 2 : Gently heat some powdered sulphur in a hard glass test tube. It melts into a pale yellow liquid. Stop heating and allow the test tube to cool. Molten sulphur quickly changes back to the solid state.

The above examples prove that physical change is both temporary and reversible.

However, some physical changes like tearing of paper, plucking of flower, chopping vegetables, etc. are *irreversible* physical changes. Even though we cannot get the original substance back, such changes are physical changes as no new substance is formed.

(iii) *There is no change in mass during a physical change.*

Weigh a beaker containing some solid wax on a beam balance. Melt the wax and again weigh the beaker. You will find that the two weights are identical. This shows that *there is no change in the mass of the substance involved as a result of physical change.*

(iv) *There is usually no gain or loss of energy as a result of physical change.*

Water changes into steam by absorbing a certain amount of heat energy. The same amount of energy is given out when steam changes back into water by giving up its heat.

Therefore, we can say that there is no net gain or loss of energy as a result of physical change.

The characteristics of a physical change can be summarized as follows :

1. No new substance is formed.
2. There is no change in the chemical composition of the original substance.

3. The change is temporary and it can be reversed by reversing the conditions.
4. The change is only in the state, size, shape, colour, texture or the smell of some or all of the substances that undergo physical change.
5. There is no change in the mass of the substances involved in a physical change.
6. There is usually no loss or gain of energy as a result of physical change.

Examples of physical change

The formation of dew, melting of wax, melting of ice, sublimation of iodine and camphor, magnetisation of iron, breaking of glass, drying of wet clothes, crystallisation of salt or sugar, dissolution of sugar in water, glowing of an electric bulb, formation of vapour, etc. are just a few common examples of physical change.

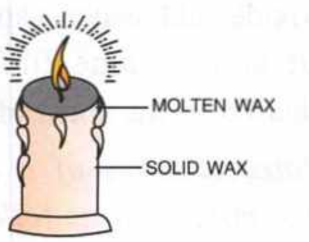


Fig. 2.3 Change of states of wax.

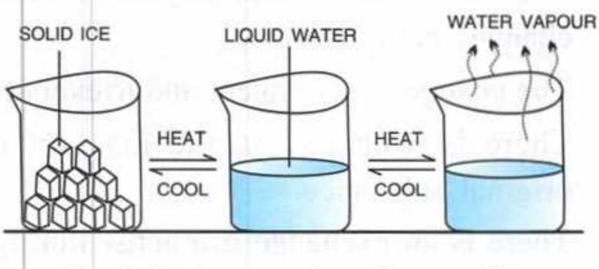


Fig. 2.4 Interconversion of states of ice.

Activity 2

Take a carrot and cut it into pieces.

A carrot cut into pieces

Now answer the following questions :

1. Is the change temporary ?

....

2. Is the change reversible ?

....

3. Why is cutting of carrot into pieces a physical change ?

....

Physical change : A physical change is a change in which no new substance is formed and the chemical composition of the original substance remains the same, even though some of its physical properties like colour, state, shape, size, *etc.* may change.

Chemical change : A chemical change is a permanent change in which new substances are formed whose chemical composition and physical and chemical properties are different from those of the original substance.

CHEMICAL CHANGE

In a chemical change, the original substance gives rise to one or more new substances with entirely different compositions and properties compared to those of the original substance.

Characteristics of a chemical change :

1. **New substances are formed :** Take some iron powder and sulphur powder in a test tube and heat them. A grey black solid is formed which is not attracted by a magnet and is insoluble in carbon disulphide. That means a new substance known as **iron sulphide** is formed which has properties completely different from the properties of iron and sulphur. [Iron is attracted by a magnet and sulphur is soluble in carbon disulphide].

$$\underset{\text{Iron}}{Fe} + \underset{\text{Sulphur}}{S} \xrightarrow{\text{heat}} \underset{\text{Iron sulphide}}{FeS}$$

2. **The change is permanent and irreversible :** When a piece of paper is burnt, a new substance *ash* is produced. Even when the burning is stopped, the ash cannot be changed back into paper. This shows that the formation of the ash from paper is a permanent and irreversible change.

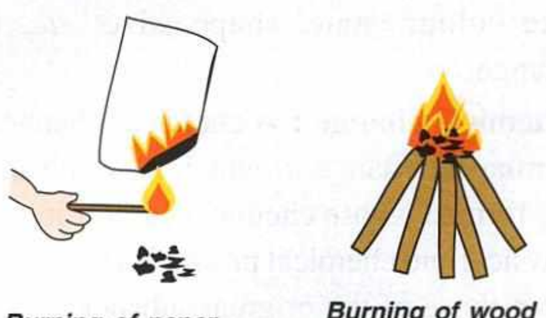


Fig. 2.5 Examples of chemical changes

3. **There is usually a change in the mass of the original substance :** Take a piece of magnesium in a crucible with a lid. Weigh it and then heat it by opening the lid after short intervals to let the air enter the crucible. When the whole of magnesium is burnt, cool the crucible and weigh it again. You will find that the final weight is more than the initial weight. This proves that when a chemical change takes place, there is a gain in the mass.

$$\underset{\text{Magnesium}}{2Mg} + O_2 \rightarrow \underset{\text{Magnesium oxide}}{2MgO}$$

In the above process, magnesium combines with oxygen to form magnesium oxide. Hence, mass is gained.

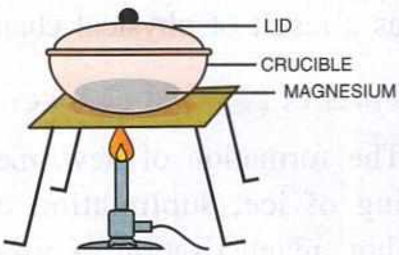


Fig. 2.6 Burning of magnesium in a crucible (weight before heating and after heating are different, showing that the final weight is more than the initial weight)

4. **Exchange of energy takes place :** When wood or paper is burnt in air, carbon dioxide and water vapour are produced, but at the same time energy is also released in the form of heat and light.

The characteristics of a chemical change can be summarized as follows :

1. New substance(s) is/are formed.
2. The composition of the original substance changes completely.
3. The change is permanent and irreversible.
4. There is a change in the mass of the original substance.
5. There is an exchange of energy during a chemical change which means that heat and light may be released or absorbed.

Examples of chemical change

The cooking of rice, the formation of curd from milk, the digestion of food, the formation of salt from acid and base, the burning of fuel, the liberation of gases, rotting of eggs, rusting of iron, *etc*. are some examples of chemical change.

The burning of candle is an example in which both physical and chemical changes take place simultaneously.

When a candle is lighted, some of the solid wax first melts and turns into liquid, then it burns to produce a flame. New substances CO_2 and water vapour are formed along with the evolution of light and heat energy. This shows a chemical change.

When some of the molten wax drops on the floor, it again solidifies which shows a physical change. Thus the melting of candle wax is a physical change and the burning of wax to produce CO_2 and H_2O is a chemical change.

Explain what will happen if an inflated balloon is brought near a lighted bulb ? Is it a phycial change or a chemical change ?

A lighted bulb

Inflated balloon brought near a lighted bulb

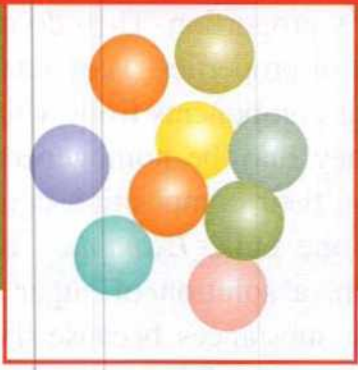
The inflated balloon bursts when brought near a lighted bulb. This is because, the air inside the balloon becomes hot, the kinetic energy of its particles increases the pressure on the inner walls of the balloon and it bursts. It is a physical change.

Table 2.1 Differences between physical and chemical changes

Physical change	Chemical change
1. In a physical change, no new substance is formed and the chemical composition of the substance remains the same. These are changes only in physical properties and state.	1. In a chemical change, a new substance(s) with entirely different chemical composition and properties is/are formed.
2. The change can be reversible or irreversible.	2. The change is permanent.
3. The change can be reversed by simple physical methods.	3. The change cannot be reversed by simple physical methods.
4. Heat or light energy may or may not be released or absorbed.	4. Heat and/or light energy are given out or absorbed.

EXERCISE

1. Define :
   (a) a physical change
   (b) a chemical change
2. Classify the following as a physical or a chemical change :
   (a) Drying of wet clothes
   (b) Manufacture of salt from sea water.
   (c) Butter getting rancid
   (d) Boiling of water
   (e) Burning of paper
   (f) Melting of wax
   (g) Burning of coal
   (h) Formation of clouds
   (i) Making of a sugar solution
   (j) Glowing of an electric bulb
   (k) Curdling of milk
   (l) Rusting of iron
   (m) Roasting of potatoes
   (n) Formation of alloys
3. Fill in the blanks :
   (a) The process of a liquid changing into a solid is called ....
   (b) A change which alters the composition of a substance is known as a .... change.
   (c) There is no change in the .... of the substance during a physical change.
   (d) The reaction in which energy is evolved is called ....
4. Give reason :
   (a) Freezing of water to ice and evaporation of water are physical changes.
   (b) Burning of a candle is both a physical and a chemical change.
   (c) Burning of paper is a chemical change.
   (d) Cutting of a cloth piece is a physical change, though it cannot be reversed.
5. Give four differences between physical and chemical changes.



Elements, Compounds and Mixtures

3

Theme : An important classification of matter is comprised of — elements, compounds and mixtures. Mixture is an important class of matter as most of the matter is found in the form of mixtures . There are various techniques by which components of a mixture can be separated.

In this chapter you will learn :

- *Elements, compounds and mixtures — a brief explanation.*
- *Separation of the components of a mixture.*
- *Emphasis on the principle of separations.*

LEARNING OUTCOMES

The children will be able to :

- recall previous knowledge related to elements, compounds and mixtures.
- classify substances into elements, compounds and mixtures on the basis of their properties.
- perform activities to separate components of a mixture.
- explain the principle involved in using a particular technique in separating a mixture.

INTRODUCTION

There are millions of substances in this world. To understand and study them systematically, they are classified in different ways. You have already learnt about **matter**, its classification into solids, liquids and gases and the interchange of the states of matter.

But matter can also be classified chemically on the basis of its composition, properties, similarity, dissimilarity and uses into elements, compounds and mixtures. You have already studied about these substances in classes VI and VII.

Let us revise and recall our previous knowledge.

Elements and compounds are pure and homogeneous substances while mixtures are impure and can be homogeneous or heterogeneous.

Pure substances : Pure substances have a definite chemical composition and definite physical and chemical properties. They are all homogeneous, *i.e.* their composition is uniform throughout the bulk. Elements and compounds are pure substances.

Example :

- Element gold is a pure substance because it is homogeneous and has a definite set of properties.
- Sodium chloride is a pure substance because all pure samples of this compound have the same chemical composition, *i.e.* 23 parts by mass of sodium and 35·5 parts by mass of chlorine.

Make a list of five elements and five compounds which you use often in your daily life.

Mixtures (impure substances) : Mixtures are made up of two or more pure substances mixed together in any proportion. They do not have any definite set of properties. They retain the properties of their components from which they are formed. They may be homogeneous or heterogeneous, *i.e.* their composition is not uniform throughout the bulk. *Example :* air, sea water, petroleum, a solution of sugar in water are all impure substances because they are made up of two or more different kinds of elements or compounds or both mixed in indefinite proportions.

Elements : An element is a pure substance which cannot be converted into anything simpler than itself by any physical or chemical process. *Example :* hydrogen, nitrogen, oxygen, sodium, iron, etc.

Elements are made up of only one kind of atoms.

Elements are the basic substances from which all other substances (compounds and mixtures) are made.

They are widely distributed in the earth's

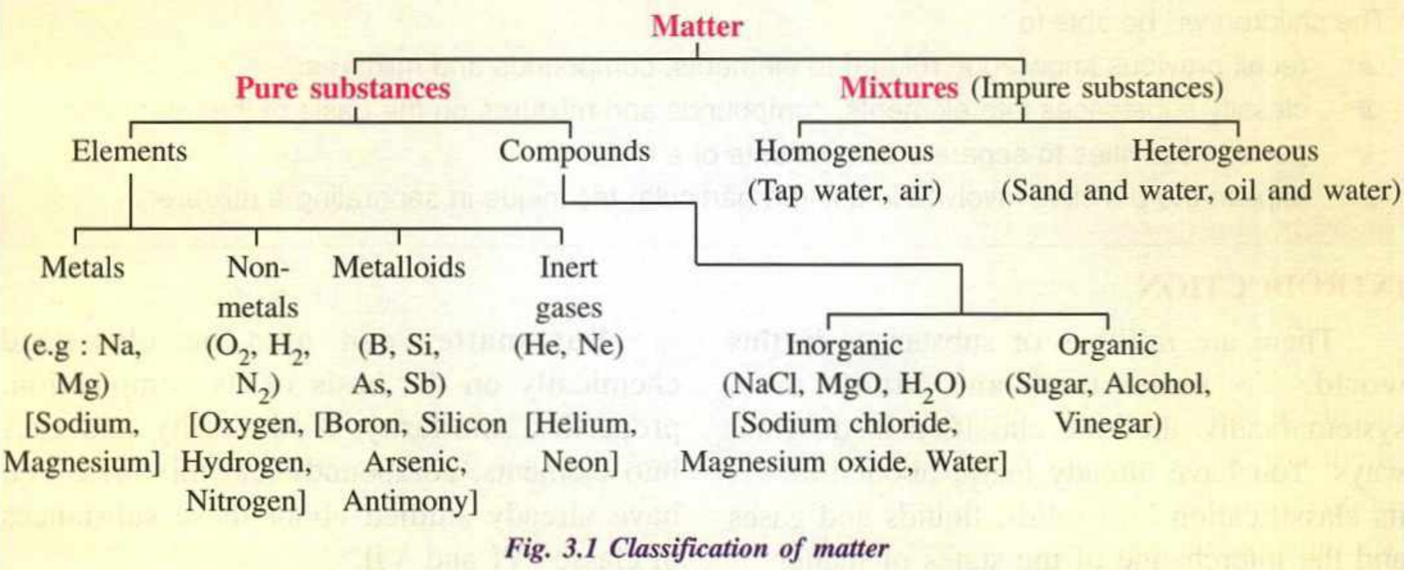


Fig. 3.1 Classification of matter

crust in free state as well as in combined states. At present 118 elements are known, out of which 92 are naturally occurring, while the rest 26 are artificially created.

Characteristics of elements :

- Each element consists of only one kind of atoms. The atoms of each element differ in properties from the atoms of other elements.
- An element is composed either of individual atoms or of molecules made up of atoms.
- Elements are pure and homogeneous.
- Elements have fixed melting and boiling points.
- Elements cannot be broken down into more simpler substances that exhibit all the properties of that element.
- Elements may react chemically with other elements and compounds.

  For example : hydrogen reacts with an element chlorine in diffused sunlight to produce a compound hydrogen chloride.

  $$H_2 + Cl_2 \rightarrow 2HCl$$

- Potassium reacts with a compound water to produce hydrogen and a compound potassium hydroxide.

  $$2K + 2H_2O \rightarrow 2KOH + H_2$$

Classification of elements : Based on their properties, elements are classified into (i) metals, (ii) non-metals, (iii) metalloids and (iv) noble or inert gases.

Metals : Most of the elements known to us are metals. They are monatomic. Most of them are hard solids. *Examples :* magnesium, copper, silver, iron, lead, aluminium, gold, sodium, potassium, etc.

Non-metals : Non-metals are mostly polyatomic in nature, *i.e.* their molecules contain two, three or more atoms. *Example :* oxygen contains two atoms in its molecule, phosphorus has four atoms, etc. They are very less in number in comparison to metals. There are only eleven non-metals (excluding noble gases) known to us. They exist in all the three physical states, *i.e.* solid, liquid and gas.

Metalloids : These are monatomic elements showing some properties of metals and some properties of non-metals. They are hard solids. *Examples :* boron, arsenic, antimony, germanium, silicon and tellurium.

Inert or noble gases : These are monatomic gaseous elements which do not react chemically with other elements or compounds, so they are known as noble or inert gases. They are found in air, in traces. They are only six in number — helium, neon, argon, krypton, xenon and radon.

Now let us recall some facts about elements that you have already studied in your previous classes.

- The most abundant element present in earth's crust is ....
- Name the element which has the highest percentage in air ....
- Name two soft metals.
- Name a liquid metal and a liquid non-metal. .... and ....
- An example of a brittle metal is ....
- Name a non-metal which is lustrous and a good conductor of electricity. ....

Write the names and symbols of the first twenty elements you have studied in class VII.

Table 3.1 : Names and symbols of some common elements

Types of elements	Name	Symbols
Metals	Iron	Fe
	Zinc	Zn
	Copper	Cu
	Magnesium	Mg
	Lead	Pb
	Calcium	Ca
	Sodium	Na
	Cobalt	Co
	Potassium	K
	Chromium	Cr
	Gold	Au
	Silver	Ag
	Platinum	Pt
	Aluminium	Al
	Mercury	Hg
	Manganese	Mn
	Nickel	Ni
Non-metals	Carbon	C
	Hydrogen	H
	Nitrogen	N
	Oxygen	O
	Fluorine	F
	Chlorine	Cl
	Bromine	Br
	Iodine	I
	Phosphorus	P
	Sulphur	S

Table 3.2 : Properties of different types of elements

Metals	Non-metals	Metalloids	Noble gases
• These elements have metallic lustre. • Mostly hard solids. • Malleable and ductile, *i.e.* can be beaten into sheets and drawn into wires. • Good conductors of heat and electricity. • They are sonorous. *Examples :* Iron, gold, silver, etc. **Exceptions :** • Sodium and potassium are soft metals. • Mercury is a liquid metal. • Zinc is brittle in nature. • Tungsten is a poor conductor of electricity.	• These elements have no lustre. • They can be soft solids, liquids or gases. • They are neither malleable nor ductile. • They are bad conductors of heat and electricity. • They are not sonorous. *Examples :* Hydrogen, nitrogen, sulphur, phosphorus, bromine, etc. **Exceptions :** • Iodine and graphite are lustrous. • Diamond is the hardest naturally occurring substance. • Carbon fibre is ductile. • Graphite is a good conductor of electricity.	These elements have properties of both metals and non-metals. *Examples :* Boron (B), Silicon (Si), Germanium (Ge), Arsenic (As) Antimony (Sb)	• Noble gases are chemically not reactive. They are inert. • They are all gases. • They are found in traces in air. *Examples :* Helium (He) Neon (Ne) Argon (Ar) Krypton (Kr) Xenon (Xe) Radon (Rn)

COMPOUNDS

A compound is a pure substance composed of two or more elements, combined chemically in a definite proportion by mass.

> Compounds are made up of different types of atoms combined chemically to form molecules.
> Hence, the smallest unit of a compound is a molecule.

The properties of a compound are entirely different from those of its constituent elements.

Example : Sodium chloride is a compound made up of sodium and chlorine when they combine chemically in the ratio of 23 : 35·5 by mass.

The properties of sodium chloride are completely different from those of sodium and chlorine. Sodium is a soft, highly reactive metal. Chlorine is a poisonous greenish yellow gas with choking smell.

But sodium chloride is a white crystalline solid which is non-poisonous and useful. It is added to our food as a mineral and also to add taste to it.

Characteristics of compounds :

- A compound contains atoms of two or more elements combined chemically in a fixed ratio by mass.
- They are pure and homogeneous. All samples of a pure compound have identical definite physical and chemical properties.
- Compounds have definite melting and boiling points.
- The constituents of a compound, from which it is made, can be separated only by chemical methods.

  Example : Mercuric oxide is a compound made up of mercury and oxygen. On heating mercuric oxide, these two elements get separated which is not possible by physical methods.

$$2HgO \xrightarrow{\text{heat}} 2Hg + O_2$$

- During the formation of a compound, energy is either absorbed or liberated.

  Example : Carbon dioxide is formed from carbon and oxygen with the liberation of heat, while nitric oxide is formed from nitrogen and oxygen with the absorption of heat energy.

(i) $C + O_2 \xrightarrow{\text{heat}} CO_2 + \text{heat}$

(heat energy is liberated)

(ii) $N_2 + O_2 + \text{heat} \xrightarrow{3000°C} 2NO$

(heat energy is absorbed in a large amount)

- Compounds have definite chemical formulae representing their molecules.

  Example : A molecule of sodium oxide is Na_2O. A molecule of sulphuric acid is H_2SO_4.

Following are some more examples to understand the characteristics of compounds:

1. **Water (H_2O) :** Water is a compound made up of two elements hydrogen and oxygen combined chemically in 1 : 8 ratio by mass. The properties of water are completely different from those of its constituent elements.

Both hydrogen and oxygen are gases but water at ordinary temperature is a liquid.

Hydrogen gas is combustible and burns, oxygen is a supporter of combustion, while water is used to extinguish fire.

These two elements cannot be separated from water by physical methods. They can be obtained only when electric current is passed through water.

2. Iron (II) sulphide (FeS) : Iron (II) sulphide is a compound formed when iron and sulphur combine chemically, on heating, in 7 : 4 ratio by mass.

Iron sulphide is a black solid, while iron is a grey black metal and sulphur is a yellow amorphous non-metallic solid.

Iron is attracted by a magnet and sulphur is soluble in carbon disulphide. Iron sulphide is neither attracted by a magnet nor it is soluble in carbon disulphide.

Table 3.3 : Names and formulae of some common compounds

Name	Formula
1. Water	H_2O
2. Carbon dioxide	CO_2
3. Sodium chloride (Common salt)	$NaCl$
4. Glucose	$C_6H_{12}O_6$
5. Sodium bicarbonate (Baking soda)	$NaHCO_3$
6. Sodium carbonate	Na_2CO_3
7. Calcium carbonate (Marble and chalk)	$CaCO_3$
8. Silicon dioxide (Sand)	SiO_2
9. Acetic acid (Vinegar)	CH_3COOH
10. Copper sulphate	$CuSO_4$
11. Magnesium oxide	MgO
12. Potassium hydroxide	KOH
13. Calcium hydroxide	$Ca(OH)_2$
14. Hydrochloric acid	HCl
15. Sulphuric acid	H_2SO_4
16. Nitric acid	HNO_3
17. Silver nitrate	$AgNO_3$
18. Barium chloride	$BaCl_2$
19. Calcium oxide	CaO
20. Ammonia	NH_3

EXERCISE 3(A)

1. Define :

(a) Elements (b) Compounds

2. Give two examples for each of the following

(a) Metals (b) Non-metals

(c) Metalloids (d) Inert gases

3. Differentiate between :

(a) Pure and impure substances

(b) Homogeneous and heterogeneous substances

4. Write the chemical name of the following and also give their molecular formulae :

(a) Baking soda (b) Vinegar

(c) Marble (d) Sand

5. Name

(a) a soft metal.

(b) a metal which is brittle.

(c) a non-metal which is lustrous.

(d) a liquid metal.

(e) a metal which is a poor conductor of electricity.

(f) a non-metal which is a good conductor of electricity.

(g) a liquid non-metal.

(h) the hardest naturally occurring substance.

(i) an inert gas.

6. How does sodium chloride differ from its constituent elements ? Explain.

7. Why is iron sulphide a compound ?

MIXTURES

Mixtures can be defined as a kind of matter which is formed by mixing two or more pure substances (elements and compounds) in any proportion, such that they do not undergo any chemical change and retain their individual properties. Therefore, they are impure substances.

Most of the substances known to us are in the form of mixtures. The substances which form mixtures are called components or constituents of mixtures.

Some common mixtures in our daily use are air, milk, tap water, honey, ice cream, etc.

Air is a mixture of oxygen, nitrogen and carbon dioxide gases. It also contains water vapour, dust particles and traces of inert gases.

Characteristics of mixtures :

- A mixture consists of two or more pure substances that exist together without any chemical combination between them.
- A mixture may be homogeneous or heterogeneous.
- The components of mixtures vary in their proportions.
- Mixtures do not have fixed melting and boiling points, they depend on the proportions of the components present in them.

  Example : Boiling point of a salt solution depends upon the amount of salt in it. The more the salt, higher is the boiling point of the solution.
- The components of mixtures can be separated by simple physical methods.
- Usually no energy change takes place during the formation of mixtures.
- Mixtures cannot be represented by any chemical formula.

Mixtures are divided into two main types on the basis of their composition :

1. **Homogeneous mixtures :** In this type of a mixture, the components are uniformly distributed throughout its volume and cannot be seen separately.

Example : A salt solution is a homogeneous mixture of salt and water in which you cannot see salt particles separately from water.

Tap water, milk, air, fruit juice, brass, bronze, etc. are some more examples of homogeneous mixtures.

2. **Heterogeneous mixtures :** In this type of a mixture, the components are not uniformly distributed throughout its volume and can be easily seen separately.

It has different composition in different parts of its bulk or mass.

Example : Soil is a mixture of many elements and compounds. Its composition varies from place to place, that is why different substances are found in the soil at different places.

Sand and stone, mud and water, kerosene and water, rice and pulses are other examples of heterogeneous mixtures.

Sand and stone

Mud and water

Fig. 3.2 Some mixtures

Table 3.4 : Differences between compounds and mixtures

Compound	Mixture
1. A compound is a pure substance.	1. A mixture is an impure substance.
2. Compounds are always homogeneous.	2. Mixtures may be homogeneous or heterogeneous.
3. A compound has a fixed composition, *i.e.* it is formed when two or more pure substances chemically combine in a definite ratio by mass.	3. A mixture has no fixed composition, *i.e.* it is formed by mixing two or more substances in any ratio without any chemical reaction.
4. Formation of a compound involves a change in energy.	4. Formation of a mixture does not involve any change in energy.
5. Compounds have a specific set of properties.	5. Mixtures do not have any specific set of properties.
6. Components of compounds can be separated only by complex chemical processes.	6. Components of mixtures can be separated by simple physical methods of separation.

FORMATION OF MIXTURES AND TYPES OF MIXTURES ON THE BASIS OF STATES OF COMPONENTS

Mixtures exist in any three states of matter, *i.e.* solid, liquid or gas depending upon the physical state of its components.

SEPARATION OF COMPONENTS OF MIXTURES

We need many substances to make our life convenient and comfortable but most of these substances are available in the form of mixtures. These mixtures contain unwanted

Table 3.5 : Different types of mixtures on the basis of their physical states

Types of mixtures	Homogeneous	Heterogeneous
1. Solid + solid	Alloys of metals, e.g. brass, bronze stainless steel, etc.	Sugar and salt, sand and stone, sodium chloride and calcium carbonate (common salt and chalk).
2 Solid + liquid	Sugar and water, salt and water, sulphur and carbon disulphide, iodine and alcohol.	Sand and water, mud and water, sugar and oil.
3. Liquid + liquid	Water and alcohol, acetone and water.	Oil and water.
4. Liquid + gas	Tap water, ammonia and water, carbon dioxide and water.	Soap lather.
5. Gas + gas	Air.	All gases are miscible.

substances which may be harmful and may degrade the properties of mixtures.

Example : Sea water is rich in common salt which is an important ingredient of our food to add taste and nutrients. But sea water cannot be directly used. It is necessary to separate salt from sea water.

The purpose of separating the components of a mixture are :

(i) to remove unwanted and harmful substances.

(ii) to obain pure and useful susbtances.

The principle of separation depends upon:

- types of mixtures and their physical states.
- size, shape and colour of the mixtures.
- their characteristic properties such as boiling point, melting point, density, volatility, solubility, magnetic properties, ability to sublime, etc.

Thus for different types of mixtures, different methods are used.

(A) Separation of solid-solid mixtures

1. Mechanical removal/hand picking : This method of separation is used only when the quantity of the mixture is small and the substance to be separated is in less amount in the mixture. The size of particles to be separated should be large and of different colours and shapes so that they could be easily recognised. *For example*, tiny stones and chaff can be separated from rice and pulses by this method.

2. Magnetic separation : This method is used when one of the components is magnetic in nature. *Example,* iron, cobalt, nickel, etc.

Mixtures of iron and sulphur, iron and sand can be easily separated by this method as iron gets attracted towards the magnet.

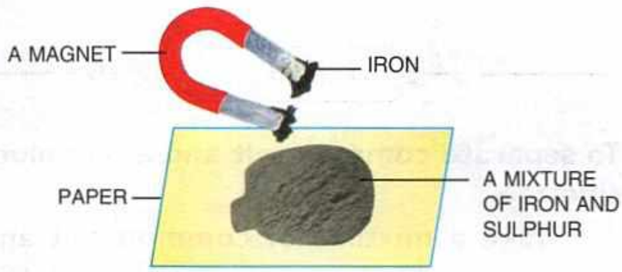


Fig. 3.3 Separation by a magnet

3. Gravitational method : This method is used only when one of the components is much heavier than water and the other component is much lighter than water.

Example : If a mixture of sand and sawdust is put in water, sawdust being lighter floats while sand settles down. Decant the water into another container to separate the sand and filter to remove the sawdust.

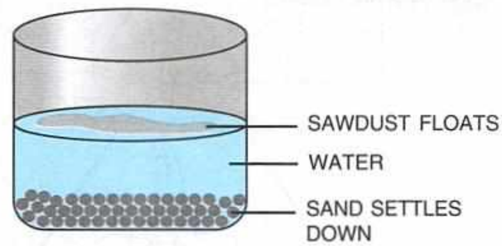


Fig. 3.4 Separation by gravitational method

4. Sublimation : *The process in which a solid changes directly into its vapours on heating is called* **sublimation**. *e.g. camphor, naphthalene, iodine, and ammonium chloride undergo sublimation.* The vapours so formed again turn into a solid upon cooling without changing into the liquid state and this process is called **deposition**.

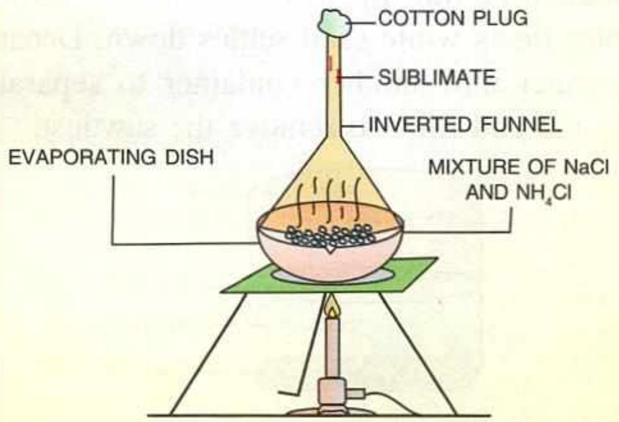
This method is used for solid mixtures in which one of the components can *sublime* on heating. The solid which sublimes escapes as vapours, while the other one is left behind.

Mixture of sand and iodine, common salt and ammonium chloride, *etc*. are separated by sublimation.

Activity 1

To separate common salt and ammonium chloride.

Take a mixture of common salt and ammonium chloride and place it in a dish and cover it with an inverted funnel as shown below and heat it. On heating, ammonium chloride changes into vapour, which condenses into a solid along the neck of the funnel (from where it may be scraped off), whereas common salt is left behind in the dish.



Sublimation : Separating a mixture of common salt and ammonium chloride

5. Solvent extraction method : This method is used when one of the components of a mixture is soluble in a particular liquid, either water or any other solvent, and the other component, which is insoluble, is separated as a residue by filtration or is decanted.

Example : A mixture of **sodium chloride** and **calcium carbonate** can be separated by this method. Salt gets dissolved in water while calcium carbonate being insoluble settles down in the container. The salt solution is then poured out (decanted), leaving behind calcium carbonate. Salt is obtained from the solution by evaporation as water evoporates leaving behind salt.

Table 3.6 : Some substances and their solvents

Substance	Solvent
1. Sulphur	Carbon disulphide
2. Paint	Turpentine oil
3. Rust	Oxalic acid
4. Rubber	Benzene
5. Nail polish	Acetone
6. Nitre (potassium nitrate)	Water
7. Grease	Petrol
8. Iodine	Ethyl alcohol
9. Chlorophyll	Methylated sprit

6. By fractional crystallisation : This method is used when the solubility of solid components of a mixture is different in the same solvent.

For example : In a mixture of common salt and potassium nitrate, both are soluble in water but solubility of potassium nitrate is more than that of sodium chloride in water.

When the hot supersaturated solution containing the mixture is cooled, potassium nitrate crystallises out first leaving behind sodium chloride. The solution left behind is boiled again and cooled to separate more potassium nitrate. The process is repeated for complete separation.

(B) Separation of solid-liquid mixtures

Such mixtures can be homogeneous or heterogeneous. Different methods are used to separate them depending upon the type of mixture.

1. Sedimentation and decantation : *The settling down of suspended, insoluble, heavy solid particles in a solid-liquid mixture when left undisturbed is called* ***sedimentation****.*

The solid which settles at the bottom is called the ***sediment****, while the clear liquid above it is called the* ***supernatant liquid****.*

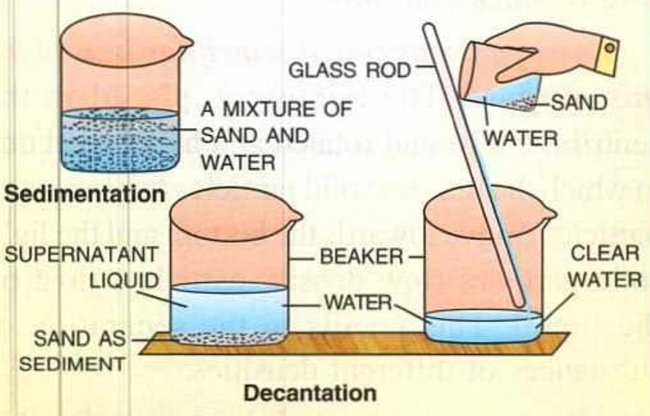
The process of pouring out the clear liquid, without disturbing the ***sediment****, is called* ***decantation****.*

This method is used for a heterogeneous mixture of a solid and a liquid, where the solid component is insoluble and heavier than the liquid component.

Example : A mixture of sand and water.

Activity 2

Take some sand and water in a beaker and stir it. Now allow the mixture to stand for some time. You will see that the sand settles at the bottom of the beaker. This is called *sedimentation*. Now pour the water gently into another vessel without disturbing the sand. This process is called ***decantation***.



2. Filtration : The process of separating insoluble solid particles from a liquid-solid mixture by allowing it to pass through a filter is called filtration. The insoluble solid left on the filter is called the **residue** while the liquid which passes through the filter is called the **filtrate**.

This method is used for separating the components of solid-liquid mixtures in which solids are insoluble in the liquid.

Example : Chalk and water, clay and water, sawdust and water, etc.

The common filters used are filter paper, charcoal, glass wool, layer of sand, etc.

In nature, suspended solid particles are removed from ground water by getting filtered through a bed of sand or gravel.

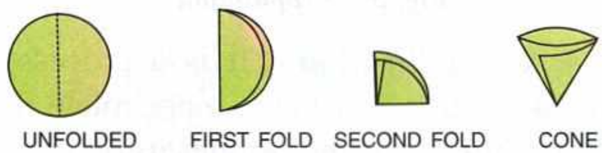


Fig. 3.5 Folding of a filter paper

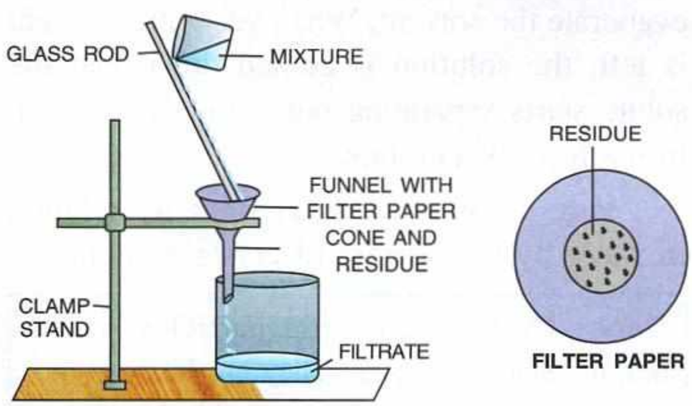


Fig. 3.6 Filtration using a funnel and filter paper

3. Evaporation : Evaporation is the process of converting a liquid into its vapour state, either by exposing it to air or by heating.

This method is used to separate the components of a homogeneous solid-liquid mixture, in which only the solid is recovered, while the liquid escapes in the form of vapour.

Example : From a mixture of common salt and water, salt is obtained easily by evaporating the solution.

Salt from sea water is separated out by this method.

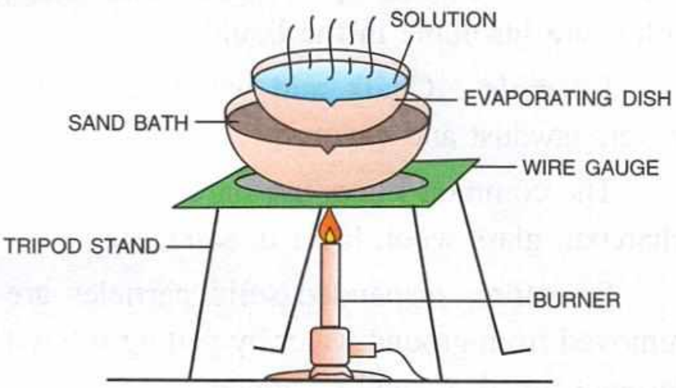


Fig. 3.7 Evaporation

4. Crystallisation : It is a process in which a solution containing more solute than it can hold at room temperature (*i.e.* a supersaturated solution) is heated in order to evaporate the solvent. When very little solvent is left, the solution is cooled down and the solute starts separating out from the solution in the form of **crystals**.

Pure sugar is obtained from its solution in water by the process of crystallisation.

> *Note :* Crystals are solid particles with a definite shape and size. They are lustrous too. *Example :* Sugar crystals are cubical and they shine.

5. Distillation : Distillation is the process of converting a liquid into vapour by heating and the subsequent condensation of the vapour back into liquid.

This method is used to separate the components of a solid-liquid mixture in which both solid and liquid are recovered. When the solution is heated, the liquid evaporates in the form of vapours, which pass through the condenser, and get condensed into pure liquid again which is called as the **distillate**, while the solid is left behind in the distilling flask.

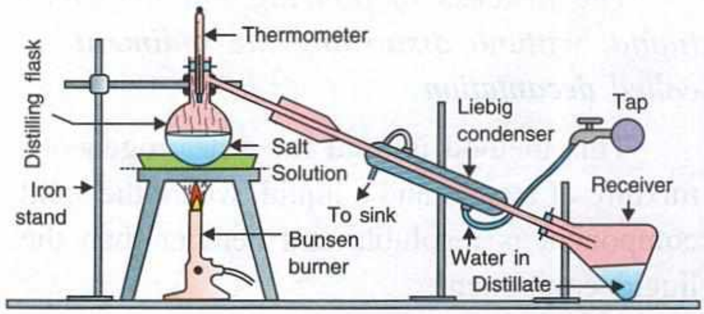


Fig. 3.8 Distillation

Tap water, a mixture containing dissolved salt is purified by this method. The pure water so obtained is called **distilled water**.

Separation of a mixture of iodine and alcohol is also done by distillation.

6. Centrifugation : *Centrifugation is the method of separating suspended solids from liquids where the mixture is homogeneous. This is also called churning.*

An apparatus called centrifuge is used for this purpose. The mixture is placed in the centrifuge tube and rotated at a high speed due to which the heavier solid particles (high density particles) move towards the bottom and the light solid particles (low density particles) float on the liquid. This results in the separation of substances of different densities.

Cream is separated from milk by this method. At home, we use mixers or traditional churners to separate cream from milk. This process is used even now *in dairies.* In washing machines, this principle is used to squeeze out water from wet clothes.

This method is also used in diagnostic laboratories for testing blood and urine.

(C) Separation of liquid-liquid mixtures

1. By separating funnel : It is a simple device used to separate the components of a liquid-liquid immiscible mixture, in which liquids have different densities.

Example : The mixture of kerosene oil and water is placed in a separating funnel and allowed to stand for sometime. The components form two clear layers. Water being heavier forms the lower layer and kerosene oil being lighter forms the upper layer. When the stopper of the funnel is opened, the heavier liquid trickles out slowly and is collected in a vessel. The stopper is closed when the lower layer is entirely removed from the funnel. In this way, the two liquids kerosene and water are separated.

Mixture of carbon tetrachloride and water can also be separated by this method in which water forms the upper layer.

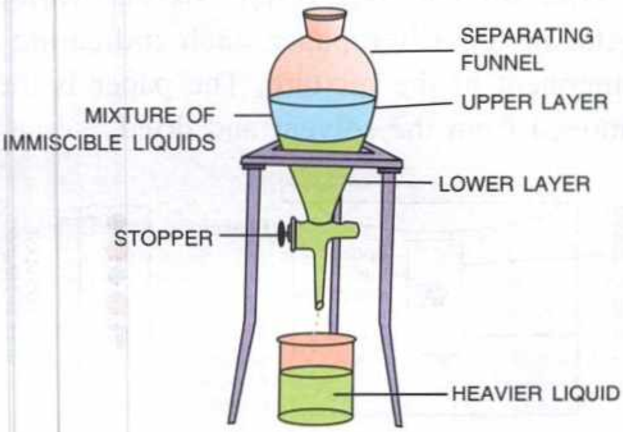


Fig. 3.9 Separation of immiscible liquids using separating funnel

2. Fractional distillation : This process of distillation is used for separating the liquid-liquid homogeneous (*miscible*) mixtures, in which the liquids have different boiling points.

On heating the mixture in a distilling flask, the liquid with the lower boiling point converts into vapour first and then gets condensed and collected in a receiver. The temperature remains constant till whole of that liquid distils over. The heating is now continued, the temperature rises and the second liquid starts vaporising at its boiling point, the vapours then get condensed and collected in another receiver.

Complete separation is possible only when the difference in boiling points of the different liquids in a mixture is 30°C or more. If the difference is less than 30°C, a fractionating column is fitted over the distilling flask. This process is called **fractional distillation**. Water and alcohol are separated by this method, boiling point of alcohol is 78°C and that of water is 100°C. 95·5% pure alcohol is obtained by this method.

Petrol, kerosene, diesel, etc. are obtained from crude petroleum oil in a similar way.

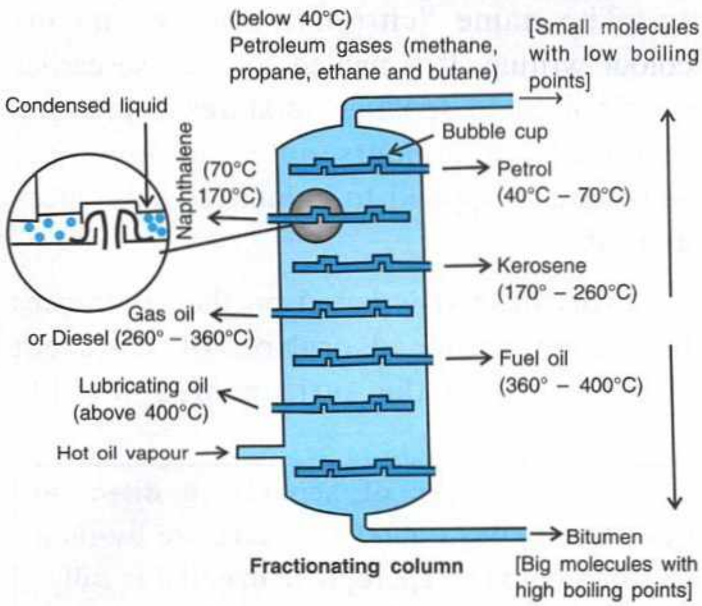


Fig. 3.10 Fractional distillation of petroleum

Fractional distillation is a process which involves distillation and the collection of fractions of different liquids which are boiling at different temperatures.

Note : Homogeneous liquid-liquid mixtures are called *miscible* liquids.

- Liquids which dissolve in each other completely in all proportions are called miscible liquids. *Example* – alcohol is miscible with water.
- Liquids which do not dissolve in each other are called immiscible liquids. They are heterogeneous liquid-liquid mixtures. *Example* – oils are immiscible with water.

3. Chromatography : This is one of the latest techniques to separate the components of a mixture when all the components are very similar in their properties.

Example : Components of ink are separated by this method. Ink is a mixture of different dyes, which are separated by chromatography because some of the dyes are less soluble and some are more soluble in a solvent.

The name "chromatography" means colour writing. It is named so because earlier it was used to separate mixtures containing coloured components only but now this technique is applied to colourless substances as well.

This method is based on the differences in the rates of adsorption of different components on the surface of a suitable adsorbent.

The process of separating different dissolved constituents of a mixture by their adsorption on an appropriate material is called chromatography.

Common adsorbents used are filter paper, silica gel, etc.

Common solvents used are water, ethyl alcohol, acetic acid, etc.

Principle involved in chromatography

Chromatography separates the components of a mixture on the basis of differences between two phases, one of which is stationary while the other is mobile.

The simplest type of chromatography is **"Paper chromatography".**

In this method, a special type of paper called chromatographic paper or Whatman filter paper is taken. A line is drawn with the pencil near the bottom edge of the paper. A drop of the mixture is placed on the filter paper above this line. The paper is then dipped in a solvent, taken in a beaker, such that the line drawn on the paper is above the level of the solvent.

The filter paper acts as "stationary phase" while the solvent acts as the "mobile phase".

As the solvent rises on the paper, it takes along with it the constituent substances of the mixture. The component of the mixture which is more soluble rises faster. We see various spots on the filter paper each indicating a component of the mixture. The paper is then removed from the solvent and dried.

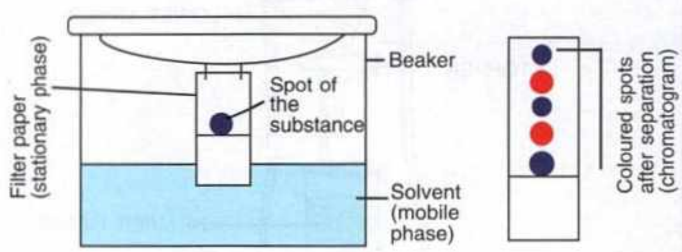


Fig. 3.13 Paper chromatography

Advantages of chromatography

- A very small quantity of the substance can be separated.
- Components with very similar physical and chemical properties can be separated.
- It identifies the different constituents of a mixture.
- It also helps in quantitative estimation of the components of a mixture.

Applications of chromatography

Chromatography can be used

1. to separate
   (a) colours in a dye
   (b) drugs from blood
   (c) pigments from natural colours.
2. to purify many industrial products.

(D) Separation of liquid-gas mixtures

A liquid-gas mixture is separated by boiling it. Dissolved gas escapes from the liquid on heating or boiling.

"The principle is based on the fact that, the solubility of a gas in a liquid decreases with an increase in temperature."

Example : Drinking water contains air dissolved in it. On boiling, air escapes from it, and the boiled water becomes tasteless.

(D) Separation of gas-gas mixtures

1. Diffusion : This method depends upon the differences in the densities of the gases present in the mixture. The lighter gas diffuses more rapidly compared to the heavier one.

A mixture of hydrogen and oxygen can be separated by diffusion as hydrogen is lighter than oxygen and diffuses first.

> The spreading out and intermixing of one substance with another substance due to the motion of its particles is called diffusion.

2. Solvent extraction : This method depends upon the fact that some gases dissolve in water or in some other solvent, while some gases do not or are less soluble.

Example : a mixture of carbon dioxide and carbon monoxide can be separated by this method as carbon dioxide is highly soluble in water while carbon monoxide is sparingly soluble.

3. Liquefaction : This method is based upon the fact that some gases like ammonia, carbon dioxide, etc. liquify easily at high pressure and low temperature while others are not easily liquefied.

Example : A mixture of ammonia and nitrogen can be separated by this method as ammonia is liquefied easily, leaving behind nitrogen gas.

Table 3.7 : Some mixtures, their methods of separation and the principle of separation

Types of mixture	Nature of mixture	Example of mixture	Method	Principle of separation
Solid + solid	Heterogeneous	Iron + Sand	Magnetic separation	Iron being magnetic in nature gets attracted to a magnet.
Solid + solid	Heterogeneous	Iodine + Common salt	Sublimation	Iodine sublimes.

Solid + solid	Heterogeneous	Common salt + sand	Solvent extraction	Common salt dissolves in water.
Solid + solid	Heterogeneous	Potassium nitrate + common salt	Fractional crystallisation	Potassium nitrate is more soluble than common salt in water.
Solid + liquid	Heterogeneous	Clay + water	Sedimentation and decantation	Clay settles down as a sediment.
Solid + liquid	Heterogeneous	Chalk + water	Filtration	Chalk is obtained as a residue and water as a filtrate.
Solid + liquid	Homogeneous	Common salt and water	Evaporation	Common salt is non-volatile, while water evaporates.
Solid + liquid	Homogeneous	Iodine + ethyl alcohol	Distillation	Ethyl alcohol vaporises and is obtained as a distillate.
Liquid + liquid	Heterogeneous	Kerosene oil + water	Separating funnel	They are immiscible liquids forming two layers, water forming the lower layer.
Liquid + liquid	Homogeneous	Ethyl alcohol + water	Fractional distillation	They differ in their boiling points.
Liquid + gas	Homogeneous	Water + carbon dioxide	Boiling	Solubility of carbon dioxide decreases on heating.
Gas + gas	Homogeneous	Nitrogen + oxygen	Liquefaction	The two gases liquify under high pressure at different temperatures.

RECAPITULATION

- There are various kinds of substances which can be pure and impure.
- A pure substance is homogeneous with a definite composition and definite physical and chemical properties.
- Elements and compounds are pure substances.
- Elements are made up of only one kind of atoms. They cannot be broken into more simpler substances.
- There are four types of elements, *i.e.* metals, non-metals, metalloids and inert gases.
- Most of the elements are metallic in nature, non-metals are very less in number.
- Compounds are pure substances formed by the chemical combination of atoms of different elements in a definite proportion. Hence, compounds are made up of different types of atoms.
- Compounds can be broken into their component elements or more simpler compounds by chemical methods.
- A mixture is an impure substance.
- A mixture can be homogeneous or heterogeneous.

- Separation of components of mixtures is necessary for their purposeful use.
- Separation of components of mixtures depends upon the type of mixtures and their characteristic properties.
- Mixtures can be solid-solid, solid-liquid, liquid-liquid, liquid-gas and gas-gas mixtures.
- Various methods are applied to separate the components of mixtures. Some of them are : magnetic separation, solvent extraction, distillation, crystallisation, sedimenation and decantation, filtration, sublimation, by separating funnel, chromatography, etc.
- Chromatography is one of the latest techniques used to separate the components which are very similar in their properties.
- Gas-gas mixtures are separated by diffusion, solvent extraction and liquefaction.

EXERCISE 3(B)

1. Classify the following substances into compounds and mixtures :

   Carbon dioxide, air, water, milk, common salt, blood, fruit juice, iron sulphide.

2. Give one example for each of the following types of mixtures.

   (a) solid-solid homogeneous mixture

   (b) solid-liquid heterogeneous mixture

   (c) miscible liquids

   (d) liquid-gas homogeneous mixture

3. Suggest a suitable technique to separate the constituents of the following mixtures. Also give the reason for selecting the particular method.

   (a) Salt from sea water

   (b) Ammonium chloride from sand

   (c) Chalk powder from water

   (d) Iron from sulphur

   (e) Water and alcohol

   (f) Sodium chloride and potassium nitrate

   (g) Calcium carbonate and sodium chloride

4. *(a)* Define 'mixture'.

   (b) Why is it necessary to separate the constituents of a mixture ?

   (c) State four differences between compounds and mixtures.

5. *(a)* What is chromatography ? For which type of mixture is it used ?

   (b) What are the advantages of chromatography ?

   (c) Give two applications of chromatography.

6. Choose the most appropriate answer from the options given below :

   (a) a mixture of sand and ammonium chloride can be separated by

   (i) filtration *(ii)* distillation
   (iii) sublimation *(iv)* crystallisation

   (b) A pair of metalloids are

   (i) Na and Mg *(ii)* B and Si
   (iii) C and P *(iv)* He and Ar

   (c) Which of the following property is not shown by compounds ?

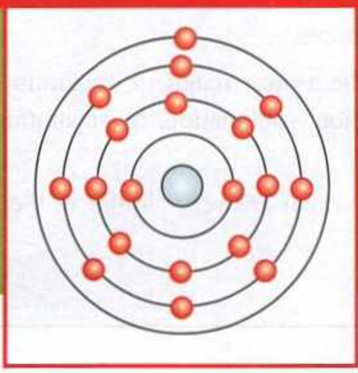
   (i) They are heterogeneous.
   (ii) They are homogeneous.
   (iii) They have definite molecular formulae.
   (iv) They have fixed melting and boiling points.

   (d) A solvent of iodine is

   (i) Water *(ii)* Kerosene oil
   (iii) Alcohol
   (iv) Petrol

   (e) This gas is highly soluble in water

   (i) Ammonia *(ii)* Nitrogen
   (iii) Carbon monoxide
   (iv) Oxygen



Atomic Structure

4

Theme : An atom is the building block of all types of matter. In fact everything on this earth is made up of atoms. It is the atom of an element that takes part in chemical reactions. Therefore, in science, it becomes important to know about the atom and its structure.

In this chapter you will learn :

- *Fundamental subatomic particles present in an atom : electrons, protons and neutrons.*
- *Nucleus and extra nuclear parts of an atom.*
- *Atomic number and mass number.*
- *The combining capacity of elements.*

LEARNING OUTCOMES

The children will be able to :

- describe that an atom consists of electrons, protons and neutrons.
- define atomic number and mass number.
- discuss valency of elements and radicals with respect to the number of hydrogen atoms combining with one atom of the element.

ATOMS : BUILDING BLOCKS OF MATTER

You are aware that anything which occupies space and has mass is matter. But do you know the smallest particle of matter is the **atom**? Atoms are extremely minute particles. They cannot be seen through the naked eye. However, there are experimental proofs about the existence of atoms, and they can even be seen through very powerful electron microscopes.

In ancient times, Indian and Greek philosophers were puzzled about the nature of matter. Gradually, the idea developed that all matter must be made of some basic elements.

In ancient times water, earth, fire, air and sky were thought to be the five fundamental elements. But we know now that an element is a pure substance made up of only one kind of atoms and has a definite set of properties.

MAHARISHI KANADA'S VIEWS ON ATOM

Maharishi Kanada was a great Indian philosopher (600 BC). According to him, "*matter consisted of indestructible particles called* ***paramanus*** (*param* means ultimate and *anu* means particle) (now called *atoms*)". A *paramanu* does not exist in free state, rather it combines with other *paramanus* to form a bigger particle called the *anu* (now known as a *molecule*). There are different types of *paramanus*. Each one of them exhibits specific properties.

The Greek philosopher **Democritus** (460 BC – 370 BC) called the *paramanu* as 'atom', which comes from the Greek word *atomos*, meaning *indivisible*.

DALTON'S ATOMIC THEORY

In 1808, **John Dalton**, an English scientist, described the *atom* as the smallest particle exhibiting all the properties of a particular element.

The main features of Dalton's atomic theory are :

1. Matter consists of very small and indivisible particles called atoms.
2. Atoms can neither be created nor be destroyed.
3. The atoms of an element are identical in all respects, *i.e.* size, mass, density, chemical properties, but they differ from the atoms of other elements.
4. Atoms of an element combine in small numbers to form molecules of that element.
5. Atoms of an element combine with the atoms of another element in a simple whole number ratio to form molecules of a compound.
6. Atoms are the smallest unit of matter that take part in chemical reactions during which only rearrangement of atoms takes place.

Note : *The latest research about atoms has proved that most of the features of Dalton's atomic theory are incorrect. But Dalton was right that atoms take part in chemical reactions.*

SUB-ATOMIC (FUNDAMENTAL) PARTICLES OF ATOMS AND EARLY MODELS OF MATTER

Studies and discoveries in the late nineteenth and the early twentieth centuries showed that atoms are divisible, *i.e.* they are composed of still smaller particles. The three main particles present in an atom are *electrons, protons* and *neutrons*. These particles are also called *fundamental* particles or *sub-atomic* particles.

The existence of the sub-atomic particles was proved by the fact that an atom is electrically neutral but it can be made to gain a positive or a negative charge. This means that an atom must contain tiny particles, each carrying either a positive or a negative charge. These opposing charges balance each other under ordinary conditions to make an atom electrically neutral.

DISCOVERY OF ELECTRONS (e^-)

Electrons were discovered in 1897 by **J.J. Thomson** when he was studying the properties of cathode rays.

Earlier, William Crooks, another British Scientist, had performed an experiment to study the phenomenon of electric discharge through gases. He observed that when an electric current of high voltage was passed through a ***discharge tube*** (a glass tube sealed at both ends with metal plates) containing a gas at very low pressure (0·01 mm of mercury), rays were emitted from the negative terminal called cathode. He called these rays '*cathode rays*'.

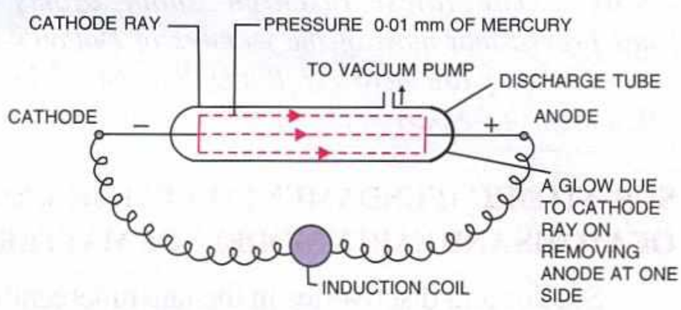


Fig. 4.1 Discharge tube in which electrons are flowing.

J.J. Thomson's work on cathode rays

J.J. Thomson studied the characteristics and the constituents of the cathode rays and concluded that : *Cathode rays consist of negatively charged particles (now called electrons),* present in atoms of all the elements.

J.J. Thomson's Experiment : An electric field was applied in the path of cathode rays in the discharge tube. It was observed that cathode rays were deflected towards the positive plate of the electric field. This showed that cathode rays were negatively charged.

When a magnetic field was applied in the path of cathode rays, they were again deflected in a direction in which moving negative charge would be deflected.

This proved that cathode rays contained negatively charged particles called electrons.

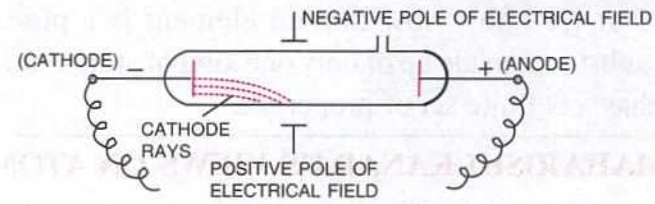


Fig. 4.2 Diagram showing deflection of cathode rays in an electric field

Properties of electrons

1. Electrons are an integral part of all atoms.
2. Its properties are independent of the nature of the gas in the discharge tube.
3. An electron has a definite mass and it carries a definite electric charge.
4. The mass of an electron has been found to be 1/1837 of the mass of a hydrogen atom, *i,e.* $9{\cdot}108 \times 10^{-28}$ g.
5. Its charge is one (1) unit negative charge, *i.e.* $1{\cdot}602 \times 10^{-19}$ coulomb.

An ***electron*** is denoted by the symbol $_{-1}e^0$. The superscript 0 represents its mass and the subscript -1 represents its one unit negative electrical charge.

DISCOVERY OF PROTONS (p^+)

The presence of the negatively charged electrons in an atom suggests that it must contain positively charged particles as well, otherwise an atom would not be electrically neutral. These positively charged particles were discovered by ***E. Goldstein***, a

German scientist, while he was performing an experiment with a discharge tube fitted with a perforated cathode with small holes to allow passage of positive rays (called canal rays) (Fig 4.3). A ray, which was just the opposite to the cathode ray in all respects, was emitted from the anode. This ray was named the *anode ray.* The anode ray consisted of the positively charged particles (now called **protons**).

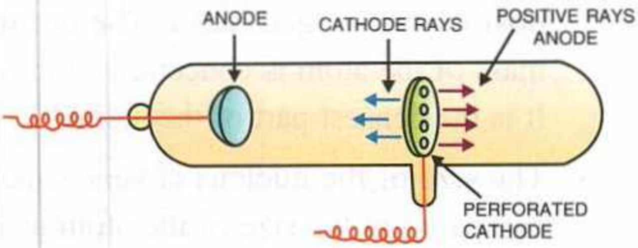


Fig. 4.3 Rays travelling in opposite directions.

Properties of protons

1. The mass of a proton was calculated as being equal to the mass of an atom of hydrogen, *i.e.* $1{\cdot}672 \times 10^{-24}$ g.
2. The positive charge on a proton is equal to the negative charge on an electron, *i.e.* $1{\cdot}602 \times 10^{-19}$ coulomb.

Further experiments proved that all elements are composed of electrons and protons. However, no two elements contain the same number of protons in their respective nuclei. *For example*, the atoms of hydrogen, helium, lithium, carbon, nitrogen and oxygen contain 1, 2, 3, 6, 7 and 8 protons respectively. Since an atom is electrically neutral, the number of electrons in an atom is equal to the number of protons in that atom.

Protons are denoted as $_{+1}p^1$, where the superscript 1 represents 1 amu (atomic mass unit) mass and the subscript +1 represents one unit positive charge.

THOMSON'S MODEL OF THE ATOM

Now the question arose as to how protons and electrons were arranged in an atom. The first model for an atom was worked out by **J.J. Thomson**. It is known as the *Plum Pudding Model* (Fig. 4.4).

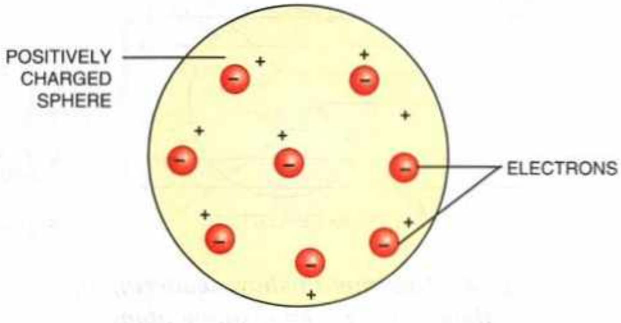


Fig. 4.4 Thomson's Plum Pudding model of the atom

According to this model, an atom is a positively charged sphere in which electrons are embedded just like dry fruits are distributed in a pudding. Therefore, it is known as the **Plum Pudding Model**.

Since the total positive charge of the atom was equal to the total negative charge of its electrons, it followed that an atom would become negatively charged if it gained electrons and positively charged if it lost electrons. However, this model failed to explain many experimental observations about atoms. Hence, Thomson's model was not accepted.

DISCOVERY OF THE NUCLEUS

In 1911, **Lord Rutherford**, a scientist from New Zealand, conducted an experiment in order to find the arrangement of electrons and protons in an atom. His experiment led to the discovery of a small, positively charged *nucleus* in the centre of the atom.

Rutherford's alpha particles scattering experiment : Rutherford bombarded a thin sheet of gold (of 0·00004 cm thickness) with alpha particles* in an evacuated chamber.

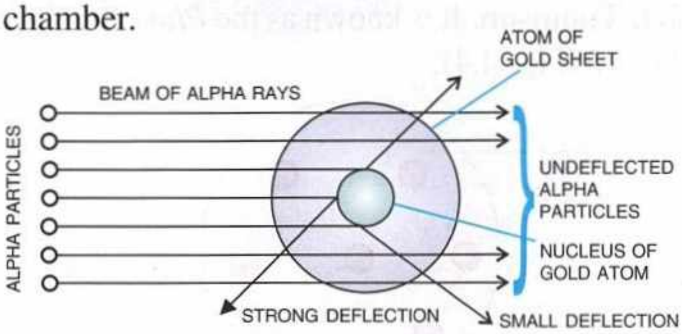


Fig. 4.5 Diagram to show scattering of alpha particles by a single atom

Following were his observations :

- Most of the alpha particles passed straight through the foil without any deflection from their path.
- A small fraction of them were deflected from their original path by small angles.
- Only a few particles bounced back.

On the basis of the above observations, Rutherford made the following conclusions :

- *Most of the space in an atom was empty* because alpha particles went straight.
- *There was a heavy positively charged mass in the atom* which caused deflection of a small fraction of alpha particles.
- The positively charged mass is very small and is centrally located because only few particles bounced back. It was named as the nucleus of an atom.

****Alpha particles are positively charged particles with two units of positive charge and four units of mass. They are formed by the removal of two electrons from a helium atom.***

Based on his experiment, Rutherford suggested a model for the structure of the atom which is known as *Rutherford's Atomic Model*.

RUTHERFORD'S ATOMIC MODEL

According to this model, an atom consists of mainly two parts :

1. *The centrally located nucleus*
   - The nucleus is a centrally located positively charged mass. The entire mass of the atom is concentrated in it. It is the densest part of the atom.
   - The size of the nucleus is very small compared to the size of the atom as a whole.

> If we consider a circular stadium as an atom, then its nucleus is no more than a cricket ball placed at the centre of the stadium.

2. *The outer circular orbits*
   - Electrons revolve in circular orbits called shells in the space available around the nucleus.
   - An atom is electrically neutral, *i.e.* the number of protons and the number of electrons present in an atom are equal.

Thus, a model similar to that of the solar system was proposed by Rutherford (Fig. 4.6). Just as in the solar system, the sun is at the centre and the planets revolve around it, in an atom the electrons revolve around the centrally located nucleus containing protons.

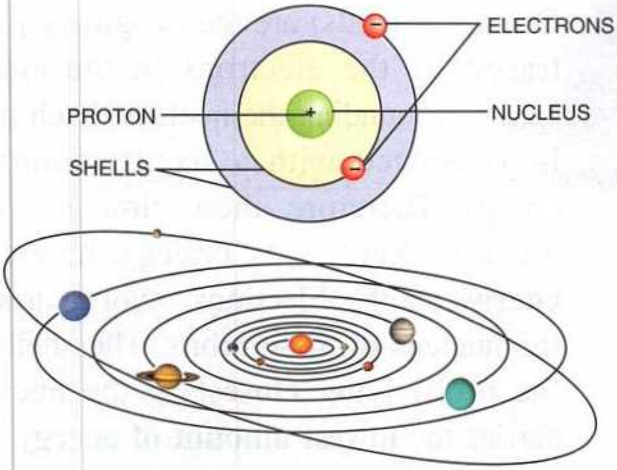


Fig. 4.6 Rutherford's model of the atom was somewhat like the solar system

DISCOVERY OF NEUTRONS ($_0n^1$)

The mass of an atom was in fact considered to be entirely concentrated within the nucleus in the form of protons, since electrons were rightly thought to have negligible mass. But when it was discovered that the nucleus had a greater mass than what could be accounted for by protons alone, it was realized that there must be a third type of sub-atomic particle, which was present in the nucleus and had neither positive nor negative charge.

In 1932, **James Chadwick** discovered this subatomic particle and called it *neutron*, since it had no charge. Its mass was found to be almost equal to that of a proton, *i.e.* $1{\cdot}672 \times 10^{-24}$ g.

Properties of neutrons

1. The mass of a neutron is slightly more than that of a proton, *i.e.* $1{\cdot}676 \times 10^{-24}$ g compared to $1{\cdot}672 \times 10^{-24}$ g of the proton.
2. Electrically a neutron is neutral, *i.e.* it has no charge.
3. Atoms of same element may differ in the number of neutrons leading to the formation of isotopes.

Table 3.1 : Properties of sub-atomic particles

Particle	*Symbol*	*Charge* ($1{\cdot}602\times10^{-19}$C)	*Atomic mass grams*
Electron	$_{-1}e^0$ or e^-	−1	$9{\cdot}1 \times 10^{-28}$ g
Proton	$_{+1}p^1$ or p^+	+1	$1{\cdot}6 \times 10^{-24}$ g
Neutron	$_0n^1$ or n	0	$1{\cdot}6 \times 10^{-24}$ g

An atom of hydrogen contains only one proton and one electron but no neutron. All other atoms have all the three particles.

STRUCTURAL STABILITY OF AN ATOM

We know that there exists a force of attraction between particles with opposite electrical charges. Thus, there is a force of attraction between the electrons and the protons present within an atom. It is expected that electrons being lighter, charged and in constant motion, would gradually lose energy and come closer to the nucleus and eventually fall into it, thus resulting in the structural collapse of the atom. But this does not happen.

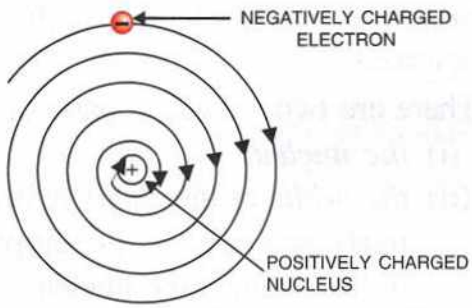


Fig. 4.7 Showing an electron losing energy and eventually falling into the nucleus, but this does not happen.

Rutherford could not explain the stability of atom. It was Neil Bohr who could explain the atomic stability.

According to Bohr's theory, the electrons revolve in fixed orbits or shells around the nucleus at a very high speed*, with each orbit

* *Velocity of light is 3×10^8 m/sec. The speed of an electron is about 1/10th the speed of light.*

associated with a fixed amount of energy. The electrons present in these shells neither lose nor gain energy until some external force is applied on it. Thus, they maintain their position. As a result the inward force exerted by the nucleus is counter balanced by the outward force of the moving electrons, thus preventing the electrons from falling into the nucleus thereby making the atom structurally stable.

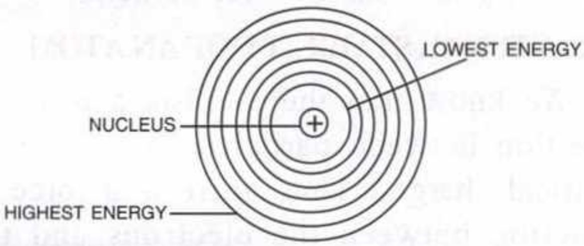


Fig. 4.8 Showing shells or orbits in an atom each with a fixed amount of energy.

MODERN MODEL OF AN ATOM

According to the modern, *standard model of atom* :

1. An atom consists of the sub-atomic particles called electrons, protons and neutrons.
2. There are two structural parts of an atom.
   (i) *the nucleus*
   (ii) *the orbits* or the *shells* (extra nuclear part) present in the empty space that surrounds the nucleus.
3. The *nucleus* is the positively charged, central part of an atom. It contains protons and neutrons. The protons and neutrons (collectively known as **nucleons**) are held firmly in the nucleus by strong nuclear forces. The entire mass of an atom lies in its nucleus, since electrons have negligible mass. The positive charge of the nucleus is due to the protons present in it. The protons remain unaffected by the neutrons since the latter have no electrical charge.
4. Orbits (or shells) are the imaginary paths traced by the electrons in the empty space surrounding the nucleus. Each orbit is associated with a fixed amount of energy. Therefore, these circular orbits are also known as **energy levels** or **energy shells**. Electrons revolve around the nucleus in these orbits. The shell (or the orbit) lying **closest** to the nucleus carries the **lowest amount of energy** and the shell that lies **farthest** from it carries the **highest amount of energy**.
5. An atom is electrically neutral because the number of protons and the number of electrons present in it are the same, thus balancing the total charge of the atom.

While the neutral atom of an element has an equal number of protons and electrons, this number itself varies from one element to another. In fact, no two elements contain the same number of protons (or electrons), that is why each element differs from the other in its properties.

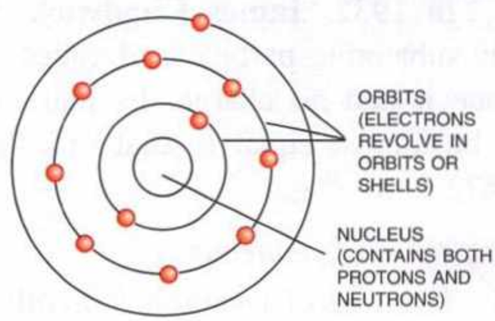


Fig. 4.9 Structure of an atom (sodium atom)

Do You Know ?

The electrons, protons and neutrons that make up an atom are just three of more than 200 subatomic particles known till date.

ATOMIC NUMBER (Z)

The number of protons present in the nucleus of the atom of an element is called its atomic number, which is denoted by the letter Z.

Since an electrically neutral atom has an equal number of protons and electrons, in such an atom :

Atomic number (Z) = Number of protons
= Number of electrons.

For example : An atom of oxygen contains 8 protons. In its neutral atom, the number of electrons is also 8. Therefore, its atomic number is 8.

> We have already noted that the atom of an element has a characteristic number of protons in its nucleus, which distinguishes it from the atoms of the other elements. The atomic number of an element never varies and it is a fixed value for a particular element.

MASS NUMBER (A)

The sum of the number of protons and the number of neutrons present in the nucleus of the atom of an element is called the mass number of that element. It is denoted by the letter A.

Mass number (A) = Number of protons
+
Number of neutrons

Thus, if the atomic number and the mass number of an element are known, one can easily calculate the number of neutrons present in the nucleus of that element.

Number of neutrons = Mass number (A) – Atomic number (Z)

For example, consider *sodium atom.* Its atomic number is 11, and its mass number is 23.

∴ Number of neutrons = A – Z
= 23 – 11 = 12.

Symbolic representation of an element with its atomic number and mass number

Suppose there is an element X with mass number A and atomic number Z. To symbolise this element, its mass number should be written as superscript (on top) and its atomic number as subscript (at the bottom), *i.e.* ${}^{A}_{Z}X$ or ${}_{Z}X^{A}$.

Example : The symbol of oxygen is O. Its mass number is 16 and its atomic number is 8. Therefore, oxygen is denoted as ${}^{16}_{8}O$ or ${}_{8}O^{16}$.

> - An atom X has atomic number 12 and mass number 24.
>
>   Find the number of protons, electrons and neutrons in the atom. Also give its symbolic representation showing its atomic number and mass number.

Table 4.2 Atomic number and mass number of some common elements

Element	*Atomic number*	*Mass number*
Hydrogen	1	1
Helium	2	4
Lithium	3	7
Beryllium	4	9
Boron	5	11
Carbon	6	12
Nitrogen	7	14
Oxygen	8	16

Fluorine	9	19
Neon	10	20
Sodium	11	23
Magnesium	12	24
Aluminium	13	27
Silicon	14	28
Phosphorus	15	31
Sulphur	16	32
Chlorine	17	35
Argon	18	40
Potassium	19	39
Calcium	20	40

ISOTOPES

By now you have learnt that each element has its unique atomic number, *i.e.* each atom of an element has the same number of protons in its nucleus. But the mass number of all the atoms of an element may or may not be the same. In other words, the nuclei of all the atoms of an element have the same number of protons but some of them have a different number of neutrons.

Therefore, we can say that there are some atoms of an element that differ from the other atoms of that element with respect to mass number, not atomic number. **Such atoms are called the isotopes of that element**.

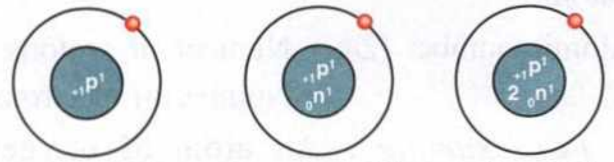
> **Isotopes** are the atoms of the same element with the same atomic number but a different mass number due to the difference in the number of neutrons in their nucleus.

Example : Element hydrogen has *three* isotopes.

Protium (Ordinary hydrogen $_1H^1$) **:** It has one proton, one electron but no neutron.

Deuterium (Heavy hydrogen $_1H^2$) **:** It has one proton, one electron and one neutron.

Tritium (Very heavy hydrogen $_1H^3$) **:** It has one proton, one electron and two neutrons.



Protium ($_1H^1$) $p^+ = 1, e^- = 1, n^0 = 0$ | Deuterium ($_1H^2$) $p^+ = 1, e^- = 1, n^0 = 1$ | Tritium ($_1H^3$) $p^+ = 1, e^- = 1, n^0 = 2$

Fig. 4.10

Similarly, **carbon** has three isotopes with mass numbers 12, 13 and 14, known as

- Carbon – 12 [$_6C^{12}$],
- Carbon – 13 [$_6C^{13}$] and
- Carbon – 14 [$_6C^{14}$]

Chlorine has *two* isotopes

- Chlorine 35 [$_{17}Cl^{35}$] and
- Chlorine 37 [$_{17}Cl^{37}$]

Properties of isotopes

1. The isotopes of an element have the same chemical properties since they all have the same atomic number.
2. Due to the same atomic number, all the isotopes of an element have the same electronic configuration.
3. Isotopes differ from each other only in their physical properties such as density, melting point, boiling point, etc. due to the difference in their mass number.

Contradictions of Dalton's Atomic Theory

It is now clear that Dalton's atomic theory was not fully correct regarding the structure of an atom. Latest research has proved this.

The drawbacks are :

1. Atoms are not indivisible, they can be further divided into fundamental particles electrons, protons and neutrons.
2. Every atom of an element is not identical. Isotopes are the atoms of the same element with same atomic number but different mass number, that means they have same number of protons and electrons but differ in their number of neutrons.

Note : Although most of the characteristic features are contradicted with new discoveries, the essence of the Dalton's atomic theory is still in use.

ARRANGEMENT OF ELECTRONS AROUND THE NUCLEUS OF AN ATOM [ELECTRONIC CONFIGURATION]

Electrons revolve around the nucleus in imaginary paths called **orbits** or **shells**. The orbit closest to the nucleus is called the *first orbit*, the next orbit is called the *second orbit* and so on. They are labelled as :

K, L, M, N, ... or

1, 2, 3, 4, ...

Each of these shells contain different number of electrons depending upon the

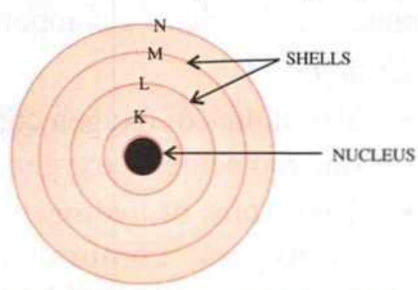


Fig. 4.11 The shells of an atom in which electrons revolve around the nucleus

amount of energy associated with them. *The shell closest to the nucleus has the lowest amount of energy and hence contains the least number of electrons.*

The following rules are followed for writing the number of electrons in different energy levels or shells :

(1) The maximum number of electrons in each shell or orbit is determined by a formula $2n^2$, where n is the number of shell. Therefore :

K shell, $n = 1$, no. of electrons $= 2 \times 1^2 = 2$

L shell, $n = 2$, no. of electrons $= 2 \times 2^2 = 8$

M shell, $n = 3$, no. of electrons $= 2 \times 3^2 = 18$

N shell, $n = 4$, no. of electrons $= 2 \times 4^2 = 32$

(2) Electrons are not accommodated in a given shell, unless the inner shells are filled. That is, the shells are filled in a step-wise manner.

Note : $2n^2$ is known as 'Bohr-Bury scheme' as it was given by scientists Bohr and Bury.

Table 4.3 : Electronic configuration of elements with atomic numbers 1-10

Element	*At. number*	*Electronic configuration*			
		K	*L*	*M*	*N*
Hydrogen	1	1			
Helium	2	2			
Lithium	3	2	1		
Beryllium	4	2	2		
Boron	5	2	3		
Carbon	6	2	4		
Nitrogen	7	2	5		
Oxygen	8	2	6		
Fluorine	9	2	7		
Neon	10	2	8		

Thus, the maximum number of electrons in the first orbit is 2, in the second 8, in the third 18 and in the fourth 32. ***However, the outermost orbit of an electrically neutral atom cannot have more than 8 electrons.***

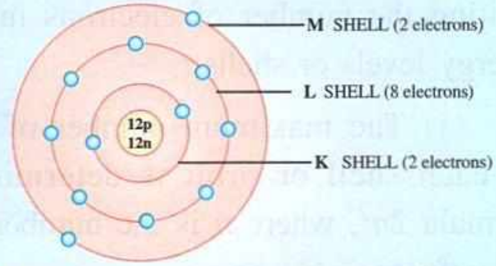


Fig. 4.12 The arrangement of electrons in a magnesium atom with atomic number = 12.

This is called the **octet rule**. If the atom has only one shell, as in case of hydrogen and helium, the outermost (single shell) can have only 2 electrons, called as the **duplet rule**.

The rules for electronic distribution can be better understood by the following examples.

- A helium atom has 2 electrons. They occupy the first and the only shell of the atom.
- A lithium atom has 3 electrons, 2 of which occupy the first shell and the third electron occupy the second shell.
- A neon atom has 10 electrons. The first shell takes up 2 electrons, while the second shell 8 electrons.
- A potassium atom has 19 electrons. The first shell takes up 2 electrons, second shell takes up 8 electrons, that leaves 9 electrons, which should occupy the third shell. Since according to the scheme $2n^2$, third shell can accomodate upto 18 electrons. But according to the octet rule, the outermost shell cannot have more than 8 electrons, hence the third shell takes up only 8 electrons and the fourth shell takes up the last electron. Therefore the electronic configuration of potassium is

  K L M N

  2 8 8 1.

Taking help from the above examples, write the electronic configuration of the following elements. Sodium (Na), Magnesium (Mg), Aluminium (Al), Silicon (Si), Phosphorus (P), Sulphur (S), Chlorine (Cl), Argon (Ar) and Calcium (Ca).

VALENCE SHELL, VALENCE ELECTRONS AND VALENCY

Valence shell : The outermost shell of an atom is known as its valence shell or valence orbit.

Valence electrons : The electrons present in the valence shell of an atom are called valence electrons. The number of valence electrons varies from 1 to 8 for the atoms of different elements. The valence electrons of an atom determine the valency of that element.

The knowledge of atom can be applied to understand how molecules of elements and compounds are formed.

You have already studied that when the atoms of the same element combine with one another, a molecule of that element is formed and when the atoms of different elements combine, a molecule of a compound is formed.

Example :

- Two atoms of oxygen combine to form one molecule of oxygen [O_2].
- Two atoms of hydrogen and one atom of oxygen combine to form one molecule of water which is a compound [H_2O].

Atomic diagrams of the atoms of the first twenty elements

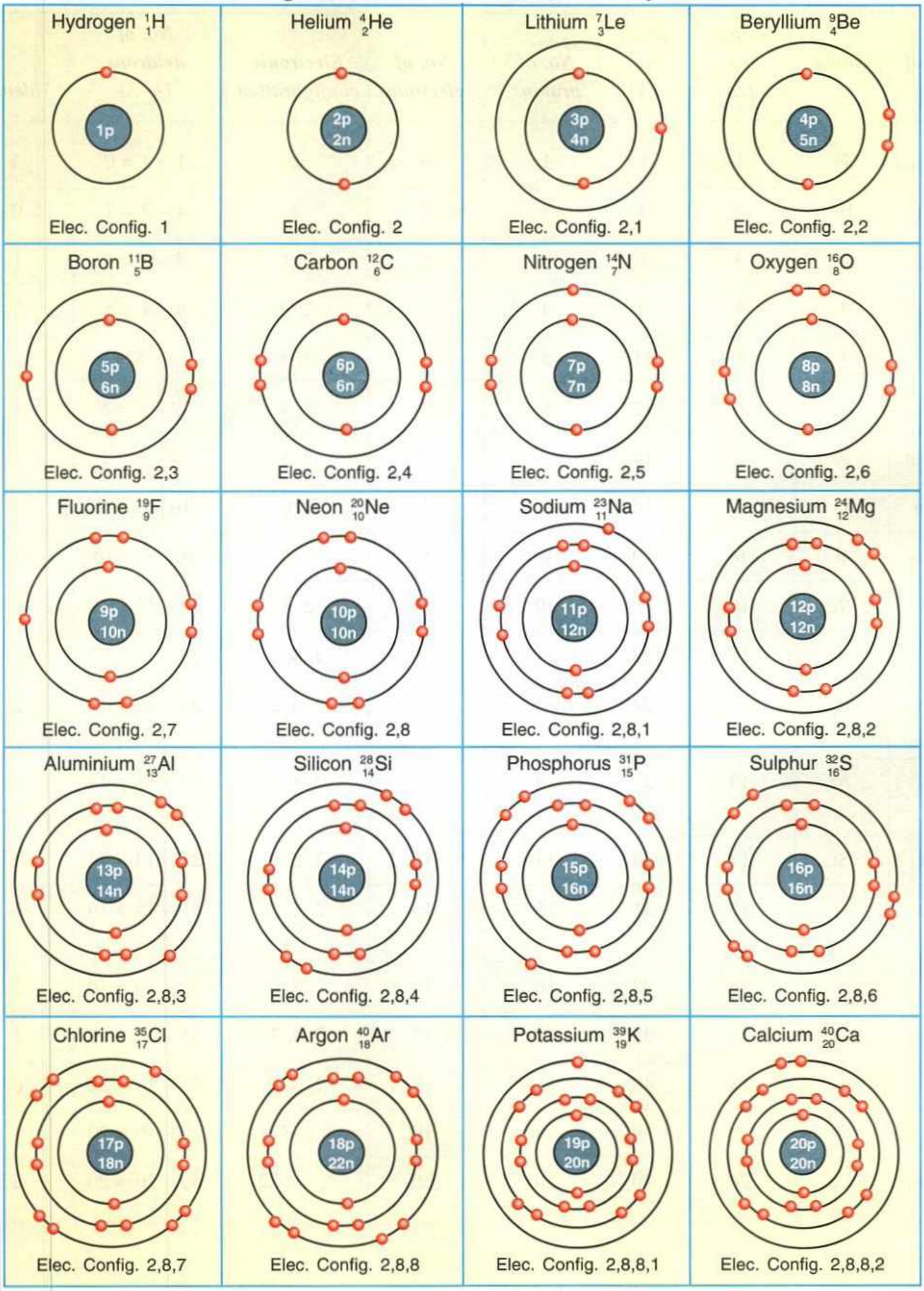


Table 4.4 : Atomic structure of the first twenty elements.

Name of element	*Symbol*	*Atomic No. (Z)*	*Mass No. (A)*	*No. of protons*	*No. of electrons*	*Electronic configuration*	*No. of neutrons (A–Z)*	*Valency*
Hydrogen	H	1	1	1	1	1	1 – 1 = 0	1
Helium	He	2	4	2	2	2	4 – 2 = 2	0
Lithium	Li	3	7	3	3	2, 1	7 – 3 = 4	1
Beryllium	Be	4	9	4	4	2, 2	9 – 4 = 5	2
Boron	B	5	11	5	5	2, 3	11 – 5 = 6	3
Carbon	C	6	12	6	6	2, 4	12 – 6 = 6	4
Nitrogen	N	7	14	7	7	2, 5	14 – 7 = 7	3
Oxygen	O	8	16	8	8	2, 6	16 – 8 = 8	2
Fluorine	F	9	19	9	9	2, 7	19 – 9 = 10	1
Neon	Ne	10	20	10	10	2, 8	20 – 10 = 10	0
Sodium	Na	11	23	11	11	2, 8, 1	23 – 11 = 12	1
Magnes-ium	Mg	12	24	12	12	2, 8, 2	24 – 12 = 12	2
Alumi-nium	Al	13	27	13	13	2, 8, 3	27 – 13 = 14	3
Silicon	Si	14	28	14	14	2, 8, 4	28 – 14 = 14	4
Phosph-orus	P	15	31	15	15	2, 8, 5	31 – 15 = 16	3
Sulphur	S	16	32	16	16	2, 8, 6	32 – 16 = 16	2
Chlorine	Cl	17	35	17	17	2, 8, 7	35 – 17 = 18	1
Argon	Ar	18	40	18	18	2, 8, 8	40 – 18 = 22	0
Potassium	K	19	39	19	19	2, 8, 8, 1	39 – 19 = 20	1
Calcium	Ca	20	40	20	20	2, 8, 8, 2	40 – 20 = 20	2

Why do atoms combine to form molecules ?

Atoms combine to form molecules in order to attain chemical stability because they have incomplete valence shell. Elements try to attain a stable electronic configuration of the nearest inert gas [Octet or Duplet] by either gaining or losing electrons or sharing electrons with other atoms.

Ion formation : When an atom loses or gains electron, it develops either a positive or a negative charge because the balance of positive and negative charges becomes unequal. In this way, ions are formed.

Hence ions are the charged particles formed due to the loss or gain of one or more electrons by an atom.

The positively charged ions are called **cations**. e.g. Na^+, Mg^{2+}, etc.

The negatively charged ions are called **anions**. *e.g.* Cl^-, O^{2-}, etc.

$$\underset{2,\,8,\,1}{Na} \xrightarrow{-e^-} \underset{2,8}{Na^+}$$

$$\underset{2,\,8,\,7}{Cl} + e^- \longrightarrow \underset{2,8,\,8}{Cl^-}$$

How do atoms combine ?

Atoms combine with one another according to their combining capacity known as **valency**.

Valency : Valency is the combining capacity of the atoms of an element with the atoms of other elements to form molecules.

The valency of an element or a radical is the number of hydrogen atoms that will combine with or displace one atom of that element or radical.

e.g. The valency of hydrogen is 1; then

(a) In the compound hydrogen chloride (HCl), one atom of chlorine combines with one atom of hydrogen, hence valency of chlorine is 1.

(b) In the compound hydrogen sulphide (H_2S), one atom of sulphur combines with two atoms of hydrogen, hence valency of sulphur is 2.

(c) In the compound hydrogen sulphate (H_2SO_4), one sulphate radical combines with two atoms of hydrogen, hence valency of sulphate is 2.

(d) In a compound hydrogen nitrate (HNO_3) *i.e.* nitric acid, one nitrate radical combines with one atom of hydrogen, hence valency of nitrate is 1.

VARIABLE VALENCY

Some elements exhibit more than one valency. They are said to have *variable valency*.

Examples : Iron, copper, tin, lead, sulphur, phosphorus, *etc*.

(a) In the case of metals exhibiting variable valency, we represent the lower valency by adding the suffix *ous* to the name of the metal; to represent the higher valency, the suffix *ic* is attached to the name of the metal.

For example, the metal iron has valencies +2 and +3. For the lower valency (+2) we write *ferrous* (Fe^{2+}) and for the higher valency (+3) we write *ferric* (Fe^{3+}). Note that the symbol remains the same but the name changes.

However in the modern method, the variable valency of the element is represented by Roman numbers. Thus ferrous ion is represented as Fe (II) and ferric ion as Fe (III).

The advantage of the modern convention is that neither the name of the element nor its symbol changes.

(b) In the case of a non-metallic atom, the number of the other types of atoms attached to it determines its valency.

Example : Phosphorus has valencies 3 and 5. With chlorine it forms two compounds, PCl_3 and PCl_5. Therefore the molecule of phosphorus trichloride, which has three chlorine atoms in it, has the lower valency (3), and the molecule of phosphorus pentachloride, which has five chlorine atoms in it, has the higher valency (5) for phosphorus atom.

Table 5.5 : Variable positive valency

Metal	*Radicals*	*Valency*
Iron (Fe)	Ferrous [Iron (II)]	2
	Ferric [Iron (III)]	3
Copper (Cu)	Cuprous [Copper (I)]	1
	Cupric [Copper (II)]	2

RADICALS

"A radical is an atom of an element or a group of atoms of different elements that behaves as a single unit with a positive or negative charge on it."

- Positively charged radicals are called basic radicals such as Na^+, NH_4^+, etc.
- Negatively charged radicals are called acid radicals such as NO_3^-, Cl^-, etc.

Basic radicals are also known as cations and acid radicals are known as anions.

Table 4.6 : Representation of some radicals

Name of radical	*Representation*	*Valency*
Ammonium	NH_4^+	1
Nitrate	NO_3^-	1
Nitrite	NO_2^-	1
Bisulphate	HSO_4^-	1
Bisulphite	HSO_3^-	1
Bicarbonate	HCO_3^-	1
Hydroxide	OH^-	1
Acetate	CH_3COO^-	1
Sulphate	SO_4^{2-}	2
Sulphite	SO_3^{2-}	2
Carbonate	CO_3^{2-}	2
Dichromate	$Cr_2O_7^{2-}$	2
Phosphate	PO_4^{3-}	3

RECAPITULATION

- All matter is made up of elements, which in turn are made of atoms.
- Maharishi Kanada was the first man to give the idea of *paramanu* and *anu*.
- Dalton's atomic theory states that atoms are indivisible.
- Later discoveries led to the modification of Dalton's theory and now it is known that an atom consists of three fundamental particles : electrons, protons and neutrons, *i.e.* atoms are divisible.
- J.J. Thomson put forth the Plum Pudding Model of the atom.
- Rutherford suggested an atomic model on the lines of the solar system.

- According to the modern model of the atom, protons and neutrons are contained in the nucleus while electrons move along only in certain fixed orbits or shells each of which is associated with a fixed amount of energy. The energy associated with a shell increases with distance from the nucleus.
- An atom is electrically neutral, *i.e.* the number of protons in an atom is equal to the number of electrons in it.
- The distribution of electrons in the various orbits of the atom of an element is called the electronic configuration of that element. The maximum number of electrons in a shell is given by the formula $2n^2$, where n is the serial number of the shell. But the outermost shell cannot have more than 8 electrons.
- The number of protons in the nucleus of an atom is called its atomic number (Z), which is also equal to the number of electrons in a neutral atom.
- The sum of the number of protons and the number of neutrons present in the nucleus of an atom is called its mass number (A).
- The charged atoms formed due to the transfer of electrons are called ions. The positively charged ion is called a cation and the negatively charged ion is called an anion.
- The atoms of an element having the same atomic number but different mass number are called the isotopes of that element.

EXERCISE

1. Fill in the blanks :

(*a*) Dalton said that .... could not be divided.

(*b*) An ion which has a positive charge is called a ....

(*c*) The outermost shell of an atom is known as ....

(*d*) The .... of an atom is very hard and dense.

(*e*) Neutrons are .... particles having mass equal to that of protons.

(*f*) Isotopes are the atoms of .... element having the .... atomic number but .... mass number.

2. Write '*true*' or '*false*' for the following statements:

(*a*) An atom on the whole has a positive charge.

(*b*) The maximum number of electrons in the first shell can be 8.

(*c*) The central part of an atom is called nucleus.

3. Give the following a suitable word/phrase.

(*a*) The subatomic particle with negative charge and negligible mass.

(*b*) Protons and neutrons present in the nucleus.

(*c*) The electrons present in the outermost shell.

(*d*) Arrangement of electrons in the shells of an atom.

(*e*) The number of protons present in the nucleus of an atom.

(*f*) The sum of the number of protons and neutrons of an atom.

(*g*) Atoms of same element with same atomic number but a different mass number.

(*h*) The smallest unit of an element which takes part in a chemical reaction.

4. Multiple Choice Questions

(*a*) The outermost shell of an atom is known as

(*i*) valency (*ii*) valence electrons
(*iii*) nucleus (*iv*) valence shell

(*b*) The number of valence electrons present in magnesium is

(*i*) two (*ii*) three
(*iii*) four (*iv*) five

(*c*) The subatomic particle with positive charge is

(*i*) proton (*ii*) neutron
(*iii*) electron (*iv*) nucleon

(d) If the atomic number of an atom is 17 and mass number is 35, then number of neutrons will be

(i) 35 (ii) 17

(iii) 18 (iv) 52

(e) The number of electrons in an atom is equal to the number of

(i) protons in a neutral atom

(ii) neutrons in a neutral atom

(iii) nucleons in a neutral atom

(iv) none of the above

(f) The sum of number of protons and number of neutrons present in the nucleus of an atom is called its

(i) mass number (ii) atomic number

(iii) number of electrons (iv) all of the above

5. Name three fundamental particles of an atom. Give the symbol with charge on each particle.

6. Define the following terms :

(a) Atomic number (b) Mass number

(c) Nucleons (d) Valence shell

7. Mention briefly the salient features of Dalton's atomic theory (five points).

8. (a) What are the two main features of Rutherford's atomic model ?

(b) State its one drawback.

9. What are the observations of the experiment done by Rutherford in order to determine the structure of an atom ?

10. State the mass number, the atomic number, number of neutrons and electronic configuration of the following atoms.

$^{12}_{6}C$ $^{16}_{8}O$ $^{19}_{9}F$ $^{20}_{10}Ne$ $_{13}Al^{27}$ $_{17}Cl^{35}$

Also, draw atomic diagrams for each of them.

11. What is variable valency ? Name two elements having variable valency and state their valencies.

12. The atomic number and the mass number of sodium are 11 and 23 respectively. What information is conveyed by this statement ?

13. Draw the diagrams representing the atomic structures of the following :

(a) Nitrogen (b) Neon

14. Explain the rule with example according to which electrons are filled in various energy levels.

15. The atom of an element is made up of 4 protons, 5 neutrons and 4 electrons. What is its atomic number and mass number ?

16. (a) What are the two main parts of which an atom is made of ?

(b) Where is the nucleus of an atom situated ?

(c) What are orbits or shells of an atom ?

17. What are isotopes ? How does the existence of isotopes contradict Dalton's atomic theory ?

18. Complete the table below by identifying A, B, C, D, E and F.

Element	Symbol	Number of protons	Number of neutrons	Number of electrons
Fluorine	$_{9}F^{19}$	9	A	**B**
Aluminium	**C**	**D**	14	13
Potassium	$_{19}K^{39}$	**E**	**F**	19

Language Of Chemistry

5

Theme : In previous classes, discussions about the symbols of elements and the formulae of compounds help in expressing their long names as short hand notations which forms the language of chemistry. Formulae of compounds can be assigned if symbols of elements/radicals forming the compound and their valencies are known. Chemical equations can also be written if the reactants and products and their symbols/formulae are known.

In this chapter you will learn :

- *Symbols of elements and their valency.*
- *Symbols of radicals and their valency.*
- *Formulae of compounds on the basis of symbols and valency of elements and radicals.*
- *To write chemical equations from the word equations by knowing the formulae of reactants and products.*
- *Law of conservation of mass.*
- *Balancing chemical equations.*
- *Information obtained from chemical equations.*
- *Limitations of chemical equations.*

LEARNING OUTCOMES

The children will be able to :

- recall the symbols of different elements.
- derive the formulae of compounds on the basis of valencies of elements and radicals.
- write chemical equations of a reaction.
- balance chemical equations by applying the law of conservation of mass.

CHEMICAL SYMBOLS OF ELEMENTS

"A symbol is the short hand representation for the atom of a specific element. Each chemical element has its own unique symbol."

But why is it necessary to represent the elements with their symbols ?

How were the symbols derived ?

The use of symbols for the elements existed long before a systematic method was developed. The alchemists associated the symbols of the planets not only with the days of a week but also with the seven metals known at that time — gold, silver, iron, mercury, tin, copper and lead.

By the beginning of nineteenth century, 26 elements were known which increased upto 81 in the twentieth century

As more elements were discovered, the need for symbolic representation for these elements became more evident.

In the 19^{th} century, a Swedish chemist John Jacob Berzelius systematically assigned the first letter of the name of an element, in capital as the symbol for that element. Such as, O for oxygen, H^- for hydrogen and so on, which was soon accepted by chemists everywhere. This later on formed the basis of IUPAC (International Union of Pure and Applied Chemistry) system of chemical symbols and formulae.

> **Suggestion for the teacher**
>
> Since students have already studied the topic, it is advised to introduce the lesson with a question-answer session for the symbols and names of the elements / radicals.

1. Each element is denoted by a symbol which is usually the first letter of its name in English or Latin [written in capital].

   Example : Oxygen is an element. It is denoted by the symbol 'O'. Similarly, hydrogen is denoted by the symbol 'H'. Now-a-days, IUPAC (International Union of Pure and Applied Chemistry) approves the names and symbols of elements.

2. However, when the first letter of more than one element is the same, the symbol is denoted by two letters. The first letter is written in capital while the second is written in small letter.

   Example : Carbon, cobalt, calcium, chromium and chlorine are the elements whose first letter is 'C'. *Carbon* is denoted by the symbol 'C'. *Cobalt* is denoted by two letters 'Co'. *Calcium* is denoted by the symbol 'Ca'. Chromium is denoted by Cr and Chlorine by Cl.

3. These symbols also represent an atom of that element.

   Example :

   (i) 'H' represents the element hydrogen as well as one atom of hydrogen.

   (ii) 'C' represents the element carbon as well as one atom of carbon.

4. Some symbols have been taken from the names of elements in Latin, German or Greek.

   Example : The symbol of iron is Fe from its Latin name Ferrum, sodium is Na from Natrium, potassium is K from Kalium. Copper is Cu from Cuprum. Therefore, each element has a name and a unique chemical symbol.

Table 5.1 : Names and symbols of some elements derived from Latin and Greek languages.

Name in English	*Name in Latin/Greek*	*Symbol*
Sodium	Natrium	Na
Potassium	Kalium	K
Magnesium	Magnesia	Mg
Aluminium	Alumen	Al
Calcium	Calx	Ca
Iron	Ferrum	Fe
Copper	Cuprum	Cu
Zinc	Zinke	Zn
Silver	Argentum	Ag
Gold	Aurum	Au
Mercury	Hydrargyrum	Hg
Lead	Plumbum	Pb
Tin	Stannum	Sn
Antimony	Stibium	Sb

Table 5.2 : Names and symbols of elements derived from English alphabets

Name in English	*Symbol*	*Name in English*	*Symbol*
Hydrogen	H	Helium	He
Nitrogen	N	Neon	Ne
Oxygen	O	Argon	Ar
Carbon	C	Krypton	Kr
Sulphur	S	Radon	Rn
Phosphorus	P	Xenon	Xe
Boron	B	Chromium	Cr
Chlorine	Cl	Cobalt	Co
Fluorine	F	Radium	Ra
Bromine	Br	Manganese	Mn
Iodine	I	Nickel	Ni
Arsenic	As	Barium	Ba
Platinum	Pt	Uranium	U
Germanium	Ge	Silicon	Si

Activity 1

Write the names and symbols of the first twenty elements that you have studied in class VI & VII.

Valency

Valency is the combining capacity of an atom of an element or of a radical with the atoms of other elements or radicals to form molecules.

The valency of an element or of a radical is the number of hydrogen atoms that will combine with or displace one atom of that element or a radical.

Examples **:** The valency of hydrogen is taken as 1.

(a) In hydrogen chloride molecule (HCl), one atom of chlorine combines with one atom of hydrogen; hence the valency of chlorine is 1.

(b) In water (H_2O), one atom of oxygen combines with two atoms of hydrogen; hence the valency of oxygen is 2.

(c) In ammonia (NH_3) gas, one atom of nitrogen combines with three atoms of hydrogen; hence the valency of nitrogen is 3.

(d) In a methane (CH_4) molecule, one carbon atom combines with four atoms of hydrogen; hence the valency of carbon is 4.

It is noticed that the valency [combining capacity] of elements increases from 1 to 4 and then again decreases to 1. For some of the elements, like Helium, Neon and Argon, the combining capacity is zero because they are inert gases and their outermost shell is complete.

Table 5.3 : Name, symbol and valency of the first twenty elements

Name of the elements	*Symbol*	*Valency*
1. Hydrogen	H	1
2. Helium	He	0
3. Lithium	Li	1
4. Beryllium	Be	2
5. Boron	B	3
6. Carbon	C	4
7. Nitrogen	N	3
8. Oxygen	O	2
9. Fluorine	F	1
10. Neon	Ne	0
11. Sodium	Na	1
12. Magnesium	Mg	2
13. Aluminium	Al	3
14. Silicon	Si	4
15. Phosphorus	P	3
16. Sulphur	S	2
17. Chlorine	Cl	1
18. Argon	Ar	0
19. Potassium	K	1
20. Calcium	Ca	2

Variable valency

Certain elements exhibit more than one valency, that means they show variable valency.

If an element exhibits two different positive valencies, then suffix ***ous*** is used for lower valency and suffix ***ic*** is used for higher valency. Now a days, instead of using suffix ***ous*** and ***ic***, the valencies are written in Roman numbers under brackets beside the name of the element. *Example :* Ferrous is written as Iron (II) and Ferric is written as Iron (III).

Non-metals like nitrogen, phosphorus and sulphur also show variable valency. Sulphur also exhibits a valency of 2, 4 and 6.

RADICALS

"A radical is an atom of an element or a group of atoms of different elements that behaves as a single unit with a positive or negative charge on it".

Radicals are of two types :

(i) **Basic radicals :** They have positive charge and are also called cations. All metallic ions and ammonium ion are basic radicals.

(ii) **Acid radicals :** They have negative charge and are also called anions. Most of the non-metallic ions and groups of non-metallic atoms with negative charge are acid radicals.

A molecule of a compound is usually made up of two parts, a positive and a negative radical. *Example :* Sodium chloride molecule is made up of sodium ion and chloride ion. Sodium ion is positively charged while chloride ion is negatively charged, represented by the symbols Na^+ and Cl^- respectively. The charges on these radicals show their combining capacity, *i.e.* valency. These radicals keep their identity in many reactions.

Table 5.4 : Metals with variable valency

Element	*Symbol*	*Lower valency state*	*Higher valency state*
Copper	Cu	Cu^+ or Cu(I) [Cuprous]	Cu^{2+} or Cu (II) [Cupric]
Iron	Fe	Fe^{2+} or Fe (II) [Ferrous]	Fe^{3+} or Fe (III) [Ferric]
Silver	Ag	Ag^+ or Ag (I) [Argentous]	Ag^{2+} or Ag (II) [Argentic]
Lead	Pb	Pb^{2+} or Pb (II) [Plumbous]	Pb^{4+} or Pb (IV) [Plumbic]
Tin	Sn	Sn^{2+} or Sn (II) [Stannous]	Sn^{4+} or Sn (IV) [Stannic]
Mercury	Hg	Hg^+ or Hg (I) [Mercurous]	Hg^{2+} or Hg (II) [Mercuric]

Table 5.5 : Valency and symbol of some common cations (Positive radicals)

Valency	*Symbol*	*Name*
Monovalent 1	Na^+	Sodium
	K^+	Potassium
	Cu^+	Copper (I) / Cuprous
	Ag^+	Silver (I) / Argentous
	Hg^+	Mercury (I) / Mercurous
	H^+	Hydrogen
	NH_4^+	Ammonium
Bivalent 2	Mg^{2+}	Magnesium
	Ca^{2+}	Calcium
	Ba^{2+}	Barium
	Zn^{2+}	Zinc
	Ni^{2+}	Nickel
	Cu^{2+}	Copper (II) / Cupric
	Fe^{2+}	Iron (II) / Ferrous
	Pb^{2+}	Lead (II) / Plumbous
	Sn^{2+}	Tin (II) / Stannous
	Ag^{2+}	Silver (II) / Argentite
	Hg^{2+}	Mercury (II) / Mercuric
	Mn^{2+}	Manganese (II) / Manganous
Trivalent 3	Al^{3+}	Aluminium
	Fe^{3+}	Iron (III) / Ferric
	Cr^{3+}	Chromium (III)
Tetravalent 4	Sn^{4+}	Tin (IV)/Stannic
	Pt^{4+}	Platinum (IV)

Table 5.6 : Valency and symbol of some common anions (Negative radicals)

Valency	*Symbol*	*Name*
Monovalent 1	Cl^-	Chloride
	Br^-	Bromide
	I^-	Iodide
	OH^-	Hydroxide
	NO_2^-	Nitrite
	NO_3^-	Nitrate
	HCO_3^-	Bicarbonate/Hydrogen carbonate
	HSO_3^-	Bisulphite/Hydrogen sulphite
	HSO_4^-	Bisulphate/Hydrogen sulphate
	CH_3COO^-	Acetate
	MnO_4^-	Permanganate
Bivalent 2	O^{2-}	Oxide
	S^{2-}	Sulphide
	SO_3^{2-}	Sulphite
	SO_4^{2-}	Sulphate
	CO_3^{2-}	Carbonate
	CrO_4^{2-}	Chromate
	$Cr_2O_7^{2-}$	Dichromate
	MnO_4^{2-}	Manganate
Trivalent 3	N^{3-}	Nitride
	PO_4^{3-}	Phosphate

MOLECULAR FORMULA OF COMPOUNDS

For convenience, each compound is represented by a chemical formula with the help of symbols and valencies of different atoms present in it. Since the formula represents the molecule of a compound, therefore, it is more commonly known as its molecular formula.

A molecular formula of a compound is the symbolic representation of its (one) molecule. It shows the number of atoms of each element present in it. These atoms combine in whole numbers to form the molecule.

For example : A molecule of sulphur dioxide gas is represented as SO_2. It indicates that one molecule of SO_2 is formed by one atom of element sulphur and two atoms of element oxygen.

WRITING THE CHEMICAL FORMULA OF A COMPOUND

To write the chemical formula of a compound, the following information should be available :

(i) Symbols of the elements or the radicals that constitute the compound.

(ii) Valencies (combining capacities) of the elements or the radicals.

The method applied for writing the chemical formula is called criss-cross method.

Step-by-step method for writing the formulae of chemical compounds.

1. *Example* – Magnesium chloride

I. Write the symbols

On the left hand side	*On the right hand side*
Magnesium Mg	Chloride Cl

II. Write the valency of the symbols

At the top right corner	*At the top right corner*
Positive ion Mg^2	Negative ion Cl^1

III. Interchange the valency number

Mg^2 Cl^1

1 2

IV. Write the interchanged numbers at the base

Mg_1 Cl_2

V. Write the formula of the compound

(ignore base number if it is 1)

$MgCl_2$

Note : The number 1 with Mg is not written in the formula because the symbol itself represents one atom.

One more example of writing molecular formula is in which valency number of positive and negative ions are divided by a common number.

2. *Example* – Calcium oxide

Step I : Write the symbols and valencies.

Symbols	
Calcium Ca	Oxide O

Valencies	
Calcium 2	Oxide 2

Step II : Ca^2 O^2

Step III : Ca_2 O_2

Step IV : Ca_2O_2

Step V : Reduce the valency numbers to the lowest ratio, if possible.
Therefore, the formula is **CaO**.

3. *Example* – Zinc hydroxide

Step I : Write the symbols and valency numbers.

Symbols	
Zinc Zn	Hydroxide OH

Valencies	
Zinc 2	Hydroxide 1

Step II : Zn^2 OH^1

Step III : Interchange the valency number. When the radical bears any number in its symbol, or if it is OH radical and more than 1 number comes after exchanging the valency number, put parenthesis (round bracket) sign enclosing the radical symbol.

Zn^2 $\quad$ OH^1

Step IV : Zn_1 $\quad$ $(OH)_2$

Step V : (Ignore the base number of Zn as it is one). Therefore, the formula is $\mathbf{Zn(OH)_2}$.

4. *Example* – Ammonium sulphate

Step I : Write the symbols and valency numbers.

Symbols	
Ammonium	Sulphate
NH_4	SO_4

Valencies	
Ammonium	Sulphate
1	2

Step II : NH_4^1 $\quad$ SO_4^2

Step III : NH_{4_2} $\quad$ SO_{4_1}

Step IV : Since radicals already bear some number in their formulae, parenthesis (round brackets) are required to be put around the radical formulae.

$(NH_4)_2 \longleftarrow (SO_4)_1$

Step V : (Ignore the base number of sulphate, as it is one). Therefore, the formula is $\mathbf{(NH_4)_2SO_4}$.

Table 5.7 : Molecular formulae of some common compounds

Name of the compound	*Molecular formula*
1. Carbon dioxide	CO_2
2. Sulphur trioxide	SO_3
3. Nitrogen dioxide	NO_2
4. Phosphorus pentoxide	P_2O_5
5. Sodium oxide	Na_2O
6. Aluminium oxide	Al_2O_3
7. Silicon dioxide (Sand)	SiO_2
8. Ammonia	NH_3
9. Iron (II) sulphide	FeS
10. Hydrogen sulphide	H_2S
11. Sodium chloride (common salt)	NaCl
12. Sodium bicarbonate	$NaHCO_3$
13. Potassium carbonate	K_2CO_3
14. Calcium carbonate (marble)	$CaCO_3$
15. Copper sulphate	$CuSO_4$
16. Iron (II) sulphate	$FeSO_4$
17. Sodium hydroxide	NaOH
18. Calcium hydroxide	$Ca(OH)_2$
19. Ammonium hydroxide	NH_4OH
20. Hydrochloric acid	HCl
21. Nitric acid	HNO_3
22. Sulphuric acid	H_2SO_4
23. Acetic acid	CH_3COOH
24. Sugar	$C_{12}H_{22}O_{11}$
25. Glucose	$C_6H_{12}O_6$

Activity 2

Write the molecular formulae of :

1. Copper oxide
2. Iron (III) chloride
3. Sodium hydroxide
4. Iron (II) sulphide
5. Lead (II) oxide
6. Hydrogen nitrate (nitric acid)
7. Hydrogen sulphate (sulphuric acid)
8. Calcium hydroxide
9. Magnesium carbonate
10. Ammonium carbonate

SIGNIFICANCE OF MOLECULAR/ CHEMICAL FORMULAE

The molecular formula of a compound not only saves time, space and energy but also gives the following information :

(i) It represents one molecule of a compound.

(ii) The number of each kind of atoms present, *i.e.* what atoms are present in one molecule of a compound and in what ratio.

(iii) The mass of one molecule of the compound can be calculated if the atomic mass of each atom present in the molecule is known. Molecular mass is the algebraic sum of the masses of all the atoms present in a given molecule.

Example :

- A molecule of sulphur dioxide is represented by the formula SO_2.
- The elements present in it are sulphur and oxygen.
- One molecule of sulphur dioxide consists of one atom of sulphur and two atoms of oxygen.
- Molecular mass of sulphur dioxide is obtained by adding the masses of each atom of sulphur and oxygen. It is found that atomic mass of sulphur is 32 amu and that of oxygen is 16 amu. Therefore, the molecular mass of sulphur dioxide $(SO_2) = 32 + 2 \times 16 = 64$ amu.

Atomic mass : The mass of an atom is called its atomic mass. Since an atom is very small, it is not possible to measure its mass in grams. Therefore, mass of a light atom like that of hydrogen or 1/12th of the mass of carbon or 1/16th of the mass of oxygen is taken as a standard unit with which mass of an atom of any other element is compared. Carbon is the most suitable and most widely accepted standard unit for the measurement of atomic mass.

This is known as **"atomic mass unit** or **amu."**

Atomic mass is also known as relative atomic mass since it is determined for an atom by comparing it with a standard unit.

Note : Atomic mass can be a whole number or fraction as it is the average mass taken for the different isotopes of an element compared to the standard unit. Usually we consider the values by rounding them off. Accordingly, from table 5.8, the atomic mass of Na is taken as 23. Atomic mass of K = 39, Al = 27 and so on.

Table 5.8 : Atomic masses of the first twenty elements

Element	*Atomic mass*	*Element*	*Atomic mass*
Hydrogen	1	Sodium	22·99
Helium	4	Magnesium	24·31
Lithium	6·9	Aluminium	26·98
Beryllium	9	Silicon	28
Boron	10·81	Phosphorus	30·97
Carbon	12	Sulphur	32
Nitrogen	14	Chlorine	35·5
Oxygen	16	Argon	39·95
Fluorine	19	Potassium	39·1
Neon	20·18	Calcium	40

Activity 3

Write the molecular formula for each of the following compounds :

1. Sulphur trioxide 2. Iron (II) sulphide and
3. Ammonia

Find the number and names of elements present in them and calculate their molecular masses.

CHEMICAL EQUATIONS

A chemical equation is the symbolic representation of a chemical reaction using symbols and formulae of the substances involved in the reaction.

The substances that are used as the starting material and which react with one another are called **reactants** and the substances which are formed as a result of the reaction are called **products**.

$$\text{Reactants} \longrightarrow \text{Products}$$

Example : Burning of coal in air is a chemical reaction in which a new substance, carbon dioxide is formed.

This reaction can be represented by either a word equation or a chemical equation (using formulae and symbols) as shown below :

$$\text{Carbon} + \text{Oygen} \xrightarrow{\text{heat}} \text{Carbon dioxide}$$

$$\underset{\text{(reactants)}}{C + O_2} \xrightarrow{\text{heat}} \underset{\text{(product)}}{CO_2}$$

In fact, a chemical equation is a short hand description of a chemical reaction. It saves space, time and energy.

The steps involved in writing a chemical equation :

(i) Write the symbols or the formulae of the reactant(s) on the left hand side, with a (+) sign between them if more than one.

(ii) Write the symbols or the formulae of the products on the right hand side, with a (+) sign between them, if more than one

(iii) Put the sign of an arrow (→) in between the reactants and the products.

(iv) Represent the reactants and the products in their molecular forms because atomic forms are usually neither stable nor capable of independent existence.

Consider the following chemical reactions

1. Reaction of lead monoxide with carbon to form lead and carbon monoxide.

$$\text{Lead monoxide} + \text{Carbon} \longrightarrow \text{Lead} + \text{Carbon monoxide}$$

$$PbO + C \longrightarrow Pb + CO$$

2. Reaction between iron and sulphur to produce iron (II) sulphide on heating.

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

$$Fe + S \xrightarrow{\text{heat}} FeS$$

3. Reaction between carbon dioxide and water to produce hydrogen carbonate (carbonic acid).

Water	+ Carbon dioxide	$\longrightarrow$	Hydrogen carbonate
H_2O	+ CO_2	$\longrightarrow$	H_2CO_3

4. Reaction between calcium oxide and water to produce calcium hydroxide.

Calcium oxide	+ Water	$\longrightarrow$	Calcium hydroxide
CaO	+ H_2O	$\longrightarrow$	$Ca(OH)_2$

You will notice that the total number of atoms on the reactant side as well as the product side are equal in each of the above equations. Such equations are called **balanced equations**.

Now consider the reaction between hydrogen and chlorine to form hydrogen chloride.

Hydrogen + Chlorine $\longrightarrow$ Hydrogen chloride [word equation]

$H_2 + Cl_2 \longrightarrow HCl$ [skeleton equation]

Note : We cannot write the equation as $H + Cl \rightarrow HCl$, because H and Cl represent the atoms of hydrogen and chlorine, which have no independent existence. The two reactants should be written as H_2 and Cl_2, *i.e.* in their molecular form. Molecules have independent existence.

In this reaction, the numbers of hydrogen and chlorine atoms on the left hand side are not equal to their numbers on the right hand side.

*Such an equation is an unbalanced one and it is known as a **skeletal equation**.*

Some more examples of skeletal equations are :

$$H_2 + O_2 \longrightarrow H_2O$$
$$Mg + O_2 \longrightarrow MgO$$
$$N_2 + H_2 \longrightarrow NH_3$$
$$Na + Cl_2 \longrightarrow NaCl$$

Remember, all chemical equations must be balanced in order to follow the law of conservation of matter or mass.

Law of conservation of mass states that matter can neither be created nor be destroyed, it can only be transformed from one form to another.

The need for balancing a chemical equation

A chemical equation needs to be balanced so as to make the number of atoms of the reactants equal to the number of atoms of the products. *This is because a chemical reaction is just a rearrangement of atoms and the atoms themselves are neither created nor destroyed during the course of a chemical reaction.*

Therefore, the balanced chemical equation for the reaction between hydrogen and chlorine is written as : $H_2 + Cl_2 \rightarrow 2HCl$.

A balanced chemical reaction is the one in which the number of atoms of each element on the reactant side is equal to the number of atoms of that element on the product side.

The significance of a balanced chemical equation

A balanced chemical equation is a wonderful way of representing a lot of information in a concise manner.

1. It shows which substances are taking part in a chemical reaction and what products are obtained as a result of it (Qualitative).

2. It shows both the number of molecules and the number of atoms of each type involved in the reaction.
3. It enables us to calculate the actual amount (mass) of reactants and products involved and formed in the reaction if atomic mass of each of the elements involved in the reaction is known (Quantitative).
4. It makes the study of chemistry universally standardized. *Example* :

$$2Ca + O_2 \rightarrow 2CaO$$

This chemical equation indicates that :

(i) calcium reacts with oxygen to form calcium oxide.

(ii) two monoatomic molecules of calcium combine with one diatomic molecule of oxygen to produce two molecules of calcium oxide.

(iii) The atomic mass of calcium is 40 amu and that of oxygen is 16 amu.

(iv) Total mass of calcium taking part in the reaction is $2 \times 40 = 80$ units and that of oxygen is $2 \times 16 = 32$ units.

(v) Total mass of reactants = 80 + 32 = 112 units and mass of calcium oxide formed is 2(40 + 16) = 112 units.

Hence 80 units of calcium combines with 32 units of oxygen to produce 112 units of calcium oxide. This is also in accordance with the law of conservation of matter. Because, **total mass of reactants = total mass of the product.**

BALANCING A CHEMICAL EQUATION

A chemical equation is balanced by following the steps given in the examples below :

***Example 1** : Burning of magnesium in the presence of oxygen.*

Magnesium burns in oxygen to give magnesium oxide.

∴ The word equation is :

Magnesium + Oxygen → Magnesium oxide

Formulae for the reactants are Mg and O_2.

Formula for the product is MgO.

∴ The skeletal equation is :

$$Mg + O_2 \rightarrow MgO$$

Steps for balancing the equation

Step I : *Count the number of atoms of each element on either side. It is convenient to start balancing with the compound that contains the maximum number of atoms.*

Number of atoms of each element on the reactant side :

Magnesium = 1
Oxygen = 2

Number of atoms of each element on the product side :

Magnesium = 1
Oxygen = 1

Therefore, we can see that there is an extra oxygen atom present on the reactant side.

Step II : *Multiply the product side by 2, because there are two atoms of oxygen present on the reactant side.*

$$Mg + O_2 \rightarrow \boxed{2}\,MgO$$

[Now we have Mg = 1, O = 2 → Mg = 2, O = 2]

But this means that there is now one extra atom of magnesium present on the right hand side (product side).

Step III : *Multiply the magnesium atom on the left hand side by 2.*

$$\boxed{2}\,Mg + O_2 \rightarrow 2MgO$$

[Number of atoms of Mg=2, O=2→ Mg=2, O=2]

This equation is now balanced and can be written as $\boxed{2Mg + O_2 = 2MgO.}$

Note : Whatever number you are writing to balance the atoms of a molecule, you are neither supposed to insert the number in between the formula nor write it as a subscript because it will change the formula.

$$\left.\begin{array}{l} Mg + O_2 \rightarrow Mg2O \\ Mg + O_2 \rightarrow MgO_2 \end{array}\right] \begin{array}{l} \times \text{ wrong} \\ \times \text{ wrong} \end{array}$$

Example 2 : Reaction of hydrogen with oxygen to produce water.

The word equation is

Hydrogen + Oxygen → Water

Formulae for the reactants are H_2 and O_2 while for the product the formula is H_2O.

Therefore, the skeletal equation is

$H_2 + O_2 \rightarrow H_2O$

Steps for balancing the equation :

Step I : Count the number of atoms of each element on either side.

Reactant side : Hydrogen = 2
Oxygen = 2

Product side : Hydrogen = 2
Oxygen = 1

There is an extra oxygen atom on the reactant side.

Step II : Multiply the product side by 2.

∴ The equation now becomes,

$$H_2 + O_2 \rightarrow \boxed{2} H_2O$$

[Now we have H = 2, O = 2 → H = 4, O = 2].

That means there are two extra hydrogen atoms on the product side.

Step III : Multiply hydrogen on the left hand side by 2.

$$\boxed{2} H_2 + O_2 \rightarrow 2H_2O$$

[Now, H = 4, O = 2 → H = 4, O = 2]

This equation is now balanced and can be written as

$$\boxed{2H_2 + O_2 \rightarrow 2H_2O}$$

Example 3 : Reaction of iron (II) with hydrochloric acid.

This gives iron (II) chloride and hydrogen. The word equation is :

Iron + Hydrochloric acid → Iron (II) chloride + Hydrogen

Formulae for the reactants are Fe and HCl.

Formulae for the products are $FeCl_2$ and H_2.

∴ The skeletal equation is :

$$Fe + HCl \rightarrow FeCl_2 + H_2$$

Step I : *Count the number of atoms of each element on either side.*

Reactant side : Iron = 1
Hydrogen = 1
Chlorine = 1

Product side : Iron = 1
Chlorine = 2
Hydrogen = 2

Note that both hydrogen and chlorine have an extra atom present on the product side.

Step II : *Multiply* HCl *on the reactant side by 2.*

$$Fe + \boxed{2} HCl \rightarrow FeCl_2 + H_2$$

[Now we have Fe=1, H=2, Cl=2 → Fe=1, Cl=2, H=2]

Thus, the number of atoms of each element on the reactant side = the number of atoms of that element on the product side.

Further steps are not required here, since the equation is already balanced.

$\therefore$ The balanced chemical equation is :

$$\boxed{Fe + 2HCl = FeCl_2 + H_2}$$

Example 4 :

Reaction of zinc with sulphuric acid.

Zinc + Sulphuric acid $\rightarrow$ Zinc sulphate + Hydrogen

The above word equation may be represented by the following chemical equation

$$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$$

Let us examine the number of atoms of different elements on both sides of the arrow.

Element	Number of atoms of reactants	Number of atoms of products
Zn	1	1
H	2	2
S	1	1
O	4	4

As the number of atoms of each element is the same on both sides of the arrow, the above equation is a balanced chemical equation.

LIMITATIONS OF A CHEMICAL EQUATION

Although a chemical equation provides a number of information about a chemical reaction, it has limitations too. It does not inform about :

(i) the physical states of the reactants and the products, *i.e.* whether they are solids, liquids or gases.

(ii) the concentration of reactants and products.

(iii) the time taken for the completion of the reaction.

(iv) the rate at which a reaction proceeds.

(v) the heat changes during the reaction, *i.e.* whether heat is given out or absorbed.

(vi) the conditions such as temperature, pressure, catalyst, etc. which affect the reaction.

(vii) the nature of the reaction, *i.e.* whether it is reversible or irreversible.

How can a chemical equation be made more informative ?

A chemical equation can give more information in the following ways :

(i) The physical state of the reactants and products can be indicated by putting (s) for solid, (*l*) for liquid, (g) for gas and (aq.) for aqueous state beside the symbols for the reactants and products.

(ii) Evolution or absorption of heat during the reaction can be denoted by adding or subtracting a heat term on the product side.

(iii) Temperature, pressure and catalyst can be indicated above the arrow ($\rightarrow$) separating the reactants and products.

(iv) Concentration of reactants and products are indicated by adding the words (dil.) for dilute and (conc.) for concentrated before their formulae.

(v) By the sign $\rightarrow$ or $\rightleftharpoons$ information about irreversible and reversible reactions can be depicted.

Examples :

Nitrogen + Hydrogen $\rightleftharpoons$ Ammonia

Can be represented as,

$$N_2(g) + 3H_2(g) \underset{\substack{400\text{–}450°C \\ 200\text{–}900 \text{ atmosphere}}}{\overset{\substack{\text{Fe (catalyst)} \\ \text{Mo (promoter)}}}{\rightleftharpoons}} 2NH_3(g) + \text{heat}$$

From the above equation it is clear that

- the reactants nitrogen and hydrogen and the product ammonia are gases.

- the reaction is reversible ($\rightleftharpoons$) and exothermic (+ heat on the product side).
- the reaction took place at a temperature between 400–500°C at a high pressure between 200–900 atmosphere.
- The catalyst is iron and molybdenum acts as a promoter which increases the efficiency of the catalyst.

(vi) Upward arrow ($\uparrow$) indicates a gas is evolved, downward arrow ($\downarrow$) indicates that a precipitate (an insoluble solid) is formed, a sign delta (Δ) indicates that heat is evolved or absorbed.

$$C + O_2 \rightarrow CO_2\uparrow + \Delta$$

(Δ with (+) sign on the product side means heat is evolved and with (–) sign means heat is absorbed).

$$\underset{(aq.)}{NaCl} + \underset{(aq.)}{AgNO_3} \rightarrow \underset{(white\ precipitate)}{AgCl\downarrow} + \underset{(aq.)}{NaNO_3}$$

RECAPITULATION

- **Symbol** is the short form of the name of an element that stands for the atom of an element.
- **Valency** is the combining capacity of an atom or a radical with the atoms of same or other elements.
- **Radical** is an atom or a group of atoms of the same or different elements that behave as a single unit and has positive or negative charge.
- A **molecular formula** of a compound is the symbolic representation of its one molecule which shows the number of atoms of each element present in it.
- A **chemical equation** is the symbolic representation of a chemical reaction using symbols and formulae of the substances involved in the reaction.
- **Law of conservation of matter** states that matter can neither be created nor be destroyed during a chemical reaction.
- Every chemical equation needs to be balanced to follow the law of conservation of mass.
- An unbalanced equation is called a **skeletal equation**.
- A balanced chemical equation tells us about the substances taking part in the reaction, the products formed and the number of molecules of each substance involved in the reaction.
- A chemical equation has some limitations but it can be made more informative.

EXERCISE

1. Define :
   (a) Radical (b) Valency
   (c) Molecular formula
2. Give the symbols and valencies of the following radicals :
   (a) Hydroxide (b) Chloride
   (c) Carbonate (d) Ammonium
   (d) Nitrate
3. Write the molecular formulae for the oxides and sulphides of the following elements.
   (a) Sodium (b) Calcium
   (c) Hydrogen
4. Write the molecular formulae for the following compounds and name the elements present.
   (a) Baking soda (b) Common salt
   (c) Sulphuric acid (d) Nitric acid
5. The valency of aluminium is 3. Write the valency of other radicals present in the following compounds.
   (a) Aluminium chloride
   (b) Aluminium oxide
   (c) Aluminium nitride
   (d) Aluminium sulphate
6. What is variable valency ? Give two examples of elements showing variable valency.
7. (a) What is a chemical equation ?
   (b) Why is it necessary to balance a chemical equation ?
   (c) What are the limitations of a chemical equation ?
8. What are the ways by which a chemical equation can be made more informative ?
9. State the law of conservation of mass.
10. Differentiate between :
    (a) Reactants and products
    (b) A balanced and an unbalanced chemical equation
11. Balance the following equations :
    (a) $N_2 + H_2 \rightarrow NH_3$
    (b) $H_2 + O_2 \rightarrow H_2O$
    (c) $Na_2O + H_2O \rightarrow NaOH$
    (d) $CO + O_2 \rightarrow CO_2$
    (e) $Zn + HCl \rightarrow ZnCl_2 + H_2$
12. Write balanced chemical equations for the following word equations :
    (a) Iron + Chlorine → Iron (III) chloride
    (b) Magnesium + dil. sulphuric acid → Magnesium sulphate + hydrogen
    (c) Magnesium + oxygen → Magnesium oxide
    (d) Calcium oxide + water → Calcium hydroxide
    (d) Sodium + chlorine → Sodium chloride
13. What information do you get from the following chemical equation ?
    $Zn(s) + 2HCl\ (dil) \rightarrow ZnCl_2(aq) + H_2(g)$

Chemical Reactions

6

Theme : Several oxides, carbonates, hydrates on heating are converted to other compounds. Oxides of metals and non-metals have basic and acidic character respectively. There are different types of reactions such as combination, decomposition, displacement, double displacement, exothermic and endothermic reactions.

In this chapter you will learn :

- *Different types of chemical reactions : Combination, Decomposition, Displacement, Double displacement*
- *Chemical reactions involve energy changes which may be obvious or at subatomic level, exothermic and endothermic reactions.*
- *Oxides and different types of oxides and their properties.*
- *Reactivity series (metal activity series)*

LEARNING OUTCOMES

The children will be able to :

- describe different types of chemical reactions with examples.
- identify the type of chemical reaction.
- identify different oxides as acidic, basic, amphoteric and neutral.
- explain the effect of heat on oxides of some metals.

A. Chemical Reaction

CHEMICAL REACTION

You now know that there are various types of chemical changes taking place in our surroundings.

Any chemical change in matter which involves transformation into one or more substances with entirely different properties is called a *chemical reaction.*

But what happens in a chemical reaction? A chemical reaction involves breaking of chemical bonds between the atoms or groups of atoms of reacting substances and rearrangement of atoms making new bonds to form new substances with absorption or release of energy normally in the form of heat and light.

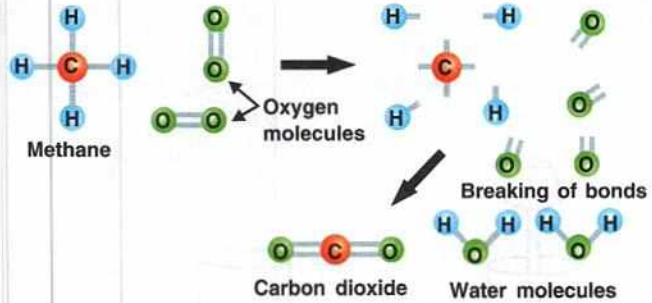


Fig. 6.1 A diagrammatic representation of breaking of bonds and formation of new bonds.

A chemical bond is the attractive force that holds the atoms of a molecule together in a compound.

Chemical reactions are represented by chemical equations. The substances taking part in a reaction are called *reactants* and the new substances formed are called *products*.

$$\text{Reactant (s)} \xrightarrow{\text{chemical reaction}} \text{Product (s)}$$

Example : Sodium hydroxide reacts with dilute hydrochloric acid to form salt and water. This chemical reaction can be represented as :

$$NaOH + HCl \rightarrow NaCl + H_2O$$

[sodium hydroxide] [hydrochloric acid] (Reactants) → [sodium chloride] [water] (Products)

Sodium hydroxide and dilute hydrochloric acid are the reactants and sodium chloride and water are the products.

Characteristics of chemical reactions

Chemical reactions are characterised by changes that are quite easily observed. Some of these typical changes are :

(i) Evolution of gas : In many chemical reactions, one of the products is a gas.

Examples

(a) When zinc reacts with dilute sulphuric acid, hydrogen gas is evolved with an effervescence.

$$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$$

[zinc] [dil. sulphuric acid] [zinc sulphate] [hydrogen]

(b) When potassium chlorate is heated strongly, it breaks up to produce oxygen gas along with potassium chloride.

$$2KClO_3 \xrightarrow[\text{strongly}]{\text{heated}} 2KCl + 3O_2$$

[potassium chlorate] [potassium chloride] [oxygen]

(c) When sodium sulphite is treated with dilute hydrochloric acid, a gas with suffocating odour (like burning sulphur) sulphur dioxide is liberated.

$$Na_2SO_3 + 2HCl \rightarrow 2NaCl + H_2O + SO_2$$

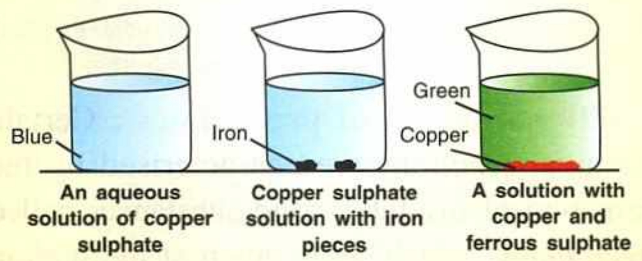
[sodium sulphite] [dil. hydrochloric acid] [sodium chloride] [water] [sulphur dioxide]

Effervescence : The formation of gas bubbles in a liquid during a chemical reaction is called *effervescence*. In reactions (a) and (c), one of the reactants is a liquid. In cases, where one of the reactants is a liquid, the gas produced forms bubbles in the liquid.

(ii) Change of colour : Certain chemical reactions are characterised by a change in the colour of the reactants.

Activity 1

Take some copper sulphate in a beaker and add some water to it.



What is the colour of the solution so formed ?

....

Now add a piece of iron to it and allow it to stand for a few hours. What do you observe ?

....

Why has the colour changed ?

....

What conclusion do you draw ?

....

Give a chemical equation for the reaction.

....

Examples

(a) When a few pieces of iron are dropped into a blue coloured copper sulphate solution, the blue colour of the solution fades and eventually turns into light green.

Fe	+	$CuSO_4\ (aq)$	$\rightarrow$	$FeSO_4$	+	Cu
[iron]		(blue solution)		(green solution)		[copper] (red deposit)

(b) When blue coloured copper sulphate reacts with hydrogen sulphide gas, a black coloured substance, copper sulphide, is formed.

$CuSO_4$	+	H_2S	$\rightarrow$	CuS	+	H_2SO_4
[copper sulphate solution] (blue)		[hydrogen sulphide]		[copper sulphide] (black solid)		[sulphuric acid]

(c) Lead nitrate is a white, crystalline solid. When heated strongly, it decomposes to produce light yellow solid lead monoxide, reddish brown nitrogen dioxide gas and colourless oxygen gas.

$2Pb(NO_3)_2$	$\rightarrow$	$2PbO$	+	$4NO_2\uparrow$	+	$O_2\uparrow$
[lead nitrate] (white)		[lead monoxide] (pale yellow)		[nitrogen dioxide] (reddish brown)		[oxygen]

(iii) Formation of precipitates : Certain chemical reactions are characterised by the formation of insoluble solid substances called ***precipitates***, which settle down at the bottom of the reaction tube.

Examples

(a) When a solution of silver nitrate is added to a solution of sodium chloride, a white insoluble substance (precipitate), silver chloride, is formed.

$AgNO_3\ (aq)$	+	$NaCl\ (aq)$	$\rightarrow$	$AgCl\downarrow$	+	$NaNO_3\ (aq)$
[silver nitrate solution]		[sodium chloride solution]		[silver chloride] (white ppt)		[sodium nitrate solution]

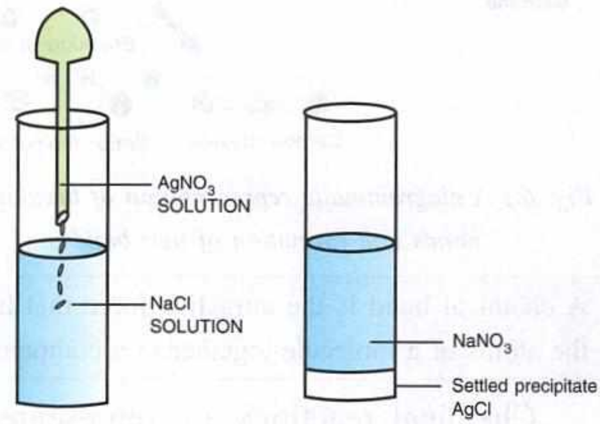


Fig. 6.2 Precipitation of silver chloride.

(b) When ferrous sulphate solution is added to sodium hydroxide solution, a dirty green precipitate of ferrous hydroxide is formed.

$FeSO_4\ (aq)$	+	$2NaOH\ (aq)$	$\rightarrow$	$Fe(OH)_2\downarrow$	+	$Na_2SO_4\ (aq)$
[ferrous sulphate]		[sodium hydroxide]		[ferrous hydroxide] (dirty green ppt)		[sodium sulphate]

(c) When few drops of dilute sulphuric acid is added to barium chloride solution, a white precipitate of barium sulphate is formed.

$BaCl_2$	+	H_2SO_4	$\rightarrow$	$BaSO_4\downarrow$	+	$2HCl$
[barium chloride solution]		[Sulphuric acid]		[Barium sulphate] (White precipitate)		[Hydrochloric acid]

(iv) Change of state : In many chemical reactions, a change of state is observed. *For example*, the reaction might start with solid or liquid reactants and end up with gaseous products, and vice-versa.

The physical states in the reaction can be represented by (s) for solid state, (l) for liquid state and (g) for gaseous state and (aq.) for aqueous solution.

(a) The reaction between hydrogen sulphide and chlorine (both gases) produces sulphur (solid) and hydrogen chloride (gas).

$$\underset{\text{[hydrogen sulphide]}}{H_2S\,(g)} + \underset{\text{[chlorine]}}{Cl_2\,(g)} \rightarrow \underset{\text{[hydrogen chloride]}}{2HCl\,(g)} + \underset{\text{[sulphur] (yellow solid)}}{S\,(s)}$$

(b) The reaction between ammonia and hydrogen chloride gases produces ammonium chloride which is a white solid.

$$\underset{\text{[ammonia]}}{NH_3\,(g)} + \underset{\text{[hydrogen chloride]}}{HCl\,(g)} \rightarrow \underset{\text{[ammonium chloride]}}{NH_4Cl\,(s)}$$

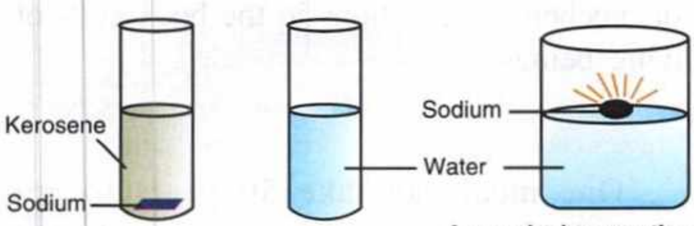
(v) Change in energy : During a chemical reaction, change in energy takes place. e.g. (a) When burning a fuel, a large amount of energy is released in the form of heat and light. (b) During decomposition of calcium carbonate, heat energy is absorbed.

Conditions necessary for chemical reactions

(i) Close contact : A chemical reaction takes place when two or more substances are mixed together, *i.e.* when they directly come into contact with each other.

Example : When sodium (metal) comes in contact with cold water, an explosive reaction occurs.

$$\underset{\text{[sodium]}}{2Na} + \underset{\text{[cold water]}}{2H_2O} \xrightarrow[\text{reaction}]{\text{vigorous}} \underset{\text{[sodium hydroxide]}}{2NaOH} + \underset{\text{[hydrogen]}}{H_2}$$



No reaction takes place because sodium is not in contact with water

An explosive reaction takes place when sodium and water come in contact

(ii) Attraction in the physical state of the reactants (through solution) : In some cases, a chemical reaction occurs when substances are mixed in solution form. This brings the reacting molecules in contact with each other at a faster rate.

Example : Sodium chloride and silver nitrate react with each other in aqueous solution form to produce a precipitate, silver chloride.

$$\underset{\text{[sodium chloride]}}{NaCl\,(aq)} + \underset{\text{[silver nitrate]}}{AgNO_3\,(aq)} \rightarrow \underset{\text{[silver chloride] (white precipitate)}}{AgCl\downarrow} + \underset{\text{[sodium nitrate]}}{NaNO_3\,(aq)}$$

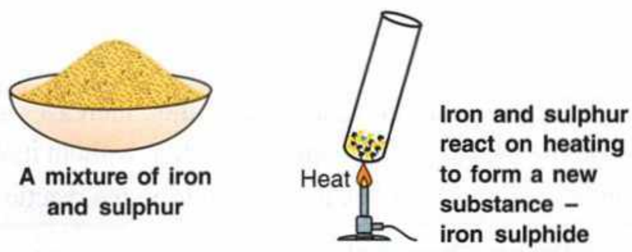
(iii) Heat energy : Some chemical reactions take place only in the presence of heat.

Example : (a) When lead nitrate is heated, it breaks down into lead monoxide, nitrogen dioxide and oxygen.

$$\underset{\text{[lead nitrate]}}{2Pb\,(NO_3)_2} \xrightarrow{\text{heat}} \underset{\text{[lead monoxide]}}{2PbO} + \underset{\text{[nitrogen dioxide]}}{4NO_2} + \underset{\text{[oxygen]}}{O_2}$$

(b) If iron powder and sulphur powder are mixed, they do not react. But when they are heated, they react to form a new substance, iron sulphide.

$$\underset{\text{[iron]}}{Fe} + \underset{\text{[sulphur]}}{S} \xrightarrow{\text{heat}} \underset{\text{[iron sulphide]}}{FeS}$$



(iv) Light energy : Some chemical reactions can take place only in the presence of light. They are called *photochemical reactions.*

Example : Photosynthesis is a chemical reaction in which food (glucose) is prepared by the green leaves of a plant but light is necessary for the reaction to take place.

$$\underset{\text{[carbon dioixde]}}{6CO_2} + \underset{\text{[water]}}{6H_2O} \xrightarrow[\text{chlorophyll}]{\text{sunlight}} \underset{\text{[glucose]}}{C_6H_{12}O_6} + \underset{\text{[oxygen]}}{6O_2}$$

Photosynthesis is a photochemical reaction in which the leaves of green plants prepare glucose and oxygen from carbon dioxide and water in the presence of chlorophyll and sunlight.

(v) Electricity : Some chemical reactions occur only when electricity is passed through the reactants. Such reactions are called electrochemical reactions.

Example : Water decomposes into hydrogen and oxygen (gases) when electric current is passed through it.

$$\underset{\text{[water]}}{2H_2O} \xrightarrow[\text{current}]{\text{electric}} \underset{\text{[hydrogen]}}{2H_2} + \underset{\text{[oxygen]}}{O_2}$$

Note : Pure water is a bad conductor of electricity. When a small amount of acid, alkali or salt is added to it, it conducts electricity.

(vi) Pressure : Some chemical reactions take place when the reactants are subjected to high pressure.

Example : Nitrogen and hydrogen when subjected to high pressure, produce ammonia gas*.

$$\underset{\text{[nitrogen]}}{N_2} + \underset{\text{[hydrogen]}}{3H_2} \underset{200\text{ atm}}{\overset{450°C}{\rightleftharpoons}} \underset{\text{[ammonia]}}{2NH_3}$$

(vii) Catalysts :

A catalyst is a substance that either increases or decreases the rate of a chemical reaction without itself undergoing any chemical change during the reaction.

Some chemical reactions need a catalyst to change the rate of the reaction, in case it is too slow or too fast.

(a) *Positive catalyst* : When a catalyst increases the rate of a chemical reaction, it is known as a **positive catalyst**.

Examples

(a) On being heated to 700°C, potassium chlorate decomposes to evolve oxygen gas. But when manganese dioxide is mixed with it (in the ratio 1 : 4), the decomposition takes place at a much lower temperature, *i.e.* at about 300°C. In this reaction, manganese dioxide acts as a positive catalyst and remains unaffected.

$$\underset{\text{[potassium chlorate]}}{2KClO_3} \xrightarrow[300°C]{MnO_2} \underset{\text{[potassium chloride]}}{2KCl} + \underset{\text{[oxygen]}}{3O_2}$$

(b) Similarly, finely divided iron is used as a positive catalyst in the manufacture of ammonia from hydrogen and nitrogen.

$$\underset{\text{[nitrogen]}}{N_2} + \underset{\text{[hydrogen]}}{3H_2} \underset{\substack{450°C \\ \text{Mo (promoter)}}}{\overset{\substack{\text{Iron (catalyst)} \\ 200-900\text{ atm}}}{\rightleftharpoons}} \underset{\text{[ammonia]}}{2NH_3} + \text{Heat}$$

Promoters : Substances that improve the efficiency of a catalyst are called promoters. *For example*, molybdenum acts as a promoter to increase the efficiency of the catalyst, iron, in the formation of ammonia gas from hydrogen and nitrogen. Promoters cannot work without a catalyst.

(b) *Negative catalyst* : When a substance decreases the rate of a chemical reaction, it is known as a **negative catalyst**.

Examples : Phosphoric acid acts as a negative catalyst to decrease the rate of decomposition of hydrogen peroxide. Alcohol too acts as a negative catalyst in certain chemical reactions.

Enzymes as catalysts in biochemical reactions

Enzymes are complex organic compounds, made up of protein units, that act as catalysts for biochemical reactions in the body cells of living beings.

For example : The process of digesting food is a series of complex chemical reactions which is a very slow process.

One meal can take 50 years to get digested. The presence of enzymes accelerates the digestion process and reduces its time to 2-3 hours.

* *This reaction takes place in the presence of iron as catalyst.*

Amylase, present in saliva, breaks starch into simple sugars.

Pepsin, in the stomach, breaks proteins into amino acids.

Amylase, trypsin and *lipase,* in the small intestine, help to break down starch, fats and proteins into glucose, fatty acids and glycerol and amino acids respectively.

Maltase, an enzyme present in the living cells of yeast, helps in the conversion of maltose (a carbohydrate) into glucose.

$$\text{Maltose + Water} \xrightarrow{\text{Maltase}} \text{Glucose}$$

The enzyme *zymase,* which is present in yeast cells, helps in the fermentation of glucose into alcohol.

$$\underset{\text{[Glucose]}}{C_6H_{12}O_6} \xrightarrow{\text{Zymase}} \underset{\text{[Ethyl alcohol]}}{2C_2H_5OH} + \underset{\text{[Carbon dioxide]}}{2CO_2}$$

EXERCISE– I

1. *(a)* Define a chemical reaction.
   (b) What happens during a chemical reaction?
   (c) What do you understand by a chemical bond ?
2. Give *one* example each that illustrates the following characteristics of a chemical reaction:
   (a) evolution of a gas
   (b) change of colour
   (c) change in state
3. How do the following help in bringing about a chemical change ? Explain each with a suitable example.
   (a) pressure *(b)* light
   (c) catalyst *(d)* heat.
4. *(a)* Define catalyst.
   (b) What are *(i)* positive catalysts and *(ii)* negative catalysts ? Support your answer with one example for each of them.
   (c) Name three biochemical catalysts found in the human body.
5. What do you observe when
   (a) dilute sulphuric acid is added to granulated zinc ?
   (b) a few pieces of iron are dropped in a blue solution of copper sulphate ?
   (c) silver nitrate is added to a solution of sodium chloride ?
   (d) ferrous sulphate solution is added to an aqueous solution of sodium hydroxide ?
   (e) solid lead nitrate is heated ?
   (f) when dilute sulphuric acid is added to barium chloride solution ?
6. Complete and balance the following chemical equations :
   (a) $N_2 + O_2 \rightarrow$
   (b) $H_2S + Cl_2 \rightarrow$
   (c) $Na + H_2O \rightarrow$
   (d) $NaCl + AgNO_3 \rightarrow$
   (e) $Zn + \underset{\text{(dil)}}{H_2SO_4} \rightarrow$
   (f) $\underset{\text{(aq.)}}{FeSO_4} + \underset{\text{(aq.)}}{NaOH} \rightarrow$
   (g) $Pb(NO_3)_2 \xrightarrow{\text{heat}}$
   (h) $\underset{\text{(aq.)}}{BaCl_2} + \underset{\text{(aq.)}}{H_2SO_4} \rightarrow$

(B.) Types of Chemical Reactions

TYPES OF CHEMICAL REACTIONS

There are different types of chemical reactions. The most common ones are :

1. Combination reaction
2. Decomposition reaction
3. Displacement reaction
4. Double displacement reaction

Combination reaction

A reaction in which two or more substances combine to form a single substance is called a *combination reaction*. This type of a reaction is also called a *synthesis reaction*.

$$A + B \rightarrow AB$$

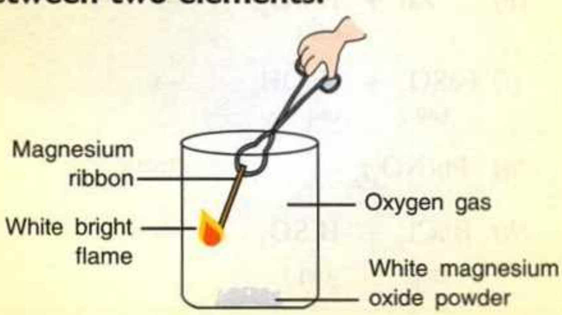
In the above reaction, the combination of substances A and B takes place to give a molecule of a new substance AB.

In combination reactions :

(i) two elements combine to form a compound.

Activity 2

To demonstrate a combination reaction between two elements.



Take a magnesium ribbon and hold it with a pair of tongs and heat it in the air. Now plunge it into a gas jar containing oxygen gas.

You will observe that magnesium ribbon burns with a dazzling white light and produces a white powder which is magnesium oxide.

The reaction can be represented as

$$2Mg + O_2 \rightarrow 2MgO \text{ (white powder)}$$

Examples

(a) When iron and sulphur (both elements) are heated together, they combine to form a compound, iron sulphide.

$$\underset{\text{[iron]}}{Fe\,(s)} + \underset{\text{[sulphur]}}{S\,(s)} \xrightarrow{\text{heat}} \underset{\text{[iron sulphide]}}{FeS\,(s)}$$

(b) Similarly, carbon burns in oxygen to form a gaseous compound, carbon dioxide.

$$\underset{\text{[carbon]}}{C\,(s)} + \underset{\text{[oxygen]}}{O_2\,(g)} \xrightarrow{\text{heat}} \underset{\text{[carbon dioxide]}}{CO_2\,(g)} + \text{Heat}$$

(ii) an element and a compound can combine to give one product.

Example : Carbon monoxide, a compound, burns in the presence of oxygen, an element, to form a single product, carbon dioxide.

$$\underset{\text{[carbon monoxide]}}{2CO\,(g)} + \underset{\text{[oxygen]}}{O_2\,(g)} \xrightarrow{\text{heat}} \underset{\text{[carbon dioxide]}}{2CO_2\,(g)}$$

(iii) two or more compounds can combine to form a single product.

Example : Ammonia and hydrogen chloride, both compounds, combine to form a compound, ammonium chloride.

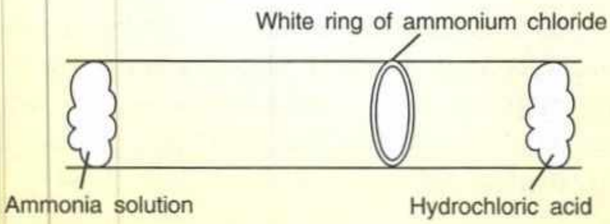
$$\underset{\text{[ammonia]}}{NH_3\,(g)} + \underset{\text{[hydrogen chloride]}}{HCl\,(g)} \rightarrow \underset{\text{[ammonium chloride]}}{NH_4Cl\,(s)}$$

Activity 3

To demonstrate a combination reaction between two compounds.

Take a wide glass tube. Fix a cotton soaked in hydrochloric acid at one end of the tube and another cotton soaked in ammonia solution at the other end of the tube.

After few minutes you will observe a white ring in the test tube which is of ammonium chloride formed due to the combination of ammonia and hydrogen chloride gases.



The ring is nearer to the cotton soaked in acid as ammonia moves faster than hydrogen chloride.

Some more examples of a combination reaction are as follows :

(i) PCl_3 [phosphorus trichloride] + Cl_2 [chlorine] $\xrightarrow{heat}$ PCl_5 [phosphorus pentachloride]

(ii) SO_2 [sulphur dioxide] + H_2O [water] $\rightarrow$ H_2SO_3 [sulphurous acid]

(iii) N_2O_5 [nitrogen pentoxide] + H_2O [water] $\rightarrow$ $2HNO_3$ [nitric acid]

(iv) CaO [calcium oxide] + H_2O [water] $\rightarrow$ $Ca(OH)_2$ [calcium hydroxide]

Decomposition reaction

A reaction in which a compound breaks up on heating into two or more simpler substances is called *decomposition reaction.*

$$AB \xrightarrow{heat} A + B$$

Here the decomposition of the molecule AB takes place to give two new substances, A and B.

Since most of the decomposition reactions are carried out by heating, they are called 'thermal decomposition reactions'.

Due to the decomposition :

(i) a compound can break up to form two or more elements.

Examples

(a) The compound mercuric oxide, when heated, decomposes to form two elements, mercury and oxygen.

$$\underset{\text{[mercuric oxide]}}{2HgO\,(s)} \xrightarrow{heat} \underset{\text{[mercury]}}{2Hg\,(s)} + \underset{\text{[oxygen]}}{O_2\,(g)}$$

(b) When electric current is passed through acidulated water, the latter decomposes into hydrogen and oxygen.

$$2H_2O\,(l) \xrightarrow[current]{electric} 2H_2\,(g) + O_2\,(g)$$

(ii) a compound can break up to form both elements and compounds.

Example : The compound potassium nitrate decomposes on heating to produce a compound, potassium nitrite and an element, oxygen.

$$\underset{\text{[potassium nitrate]}}{2KNO_3\,(s)} \xrightarrow{heat} \underset{\text{[potassium nitrite]}}{2KNO_2\,(s)} + \underset{\text{[oxygen]}}{O_2\,(g)}$$

(iii) a compound can break up to form two or more new simpler compounds.

Example : The compound calcium carbonate decomposes on strong heating to form two compounds, calcium oxide and carbon dioxide.

$$\underset{\text{[calcium carbonate]}}{CaCO_3\,(s)} \xrightarrow[100^\circ C]{heat} \underset{\text{[calcium oxide]}}{CaO\,(s)} + \underset{\text{[carbon dioxide]}}{CO_2\,(g)}$$

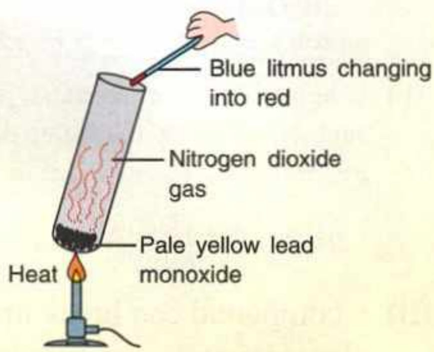
Activity 4

To demonstrate thermal decomposition of a compound.

Take some lead nitrate (a white solid) in a hard glass test tube and heat it for sometime.

You will observe that a reddish brown gas is evolved which turns moist blue litmus paper into red. Also, a gas is evolved in which a glowing splinter bursts into flame. When heating is stopped, a pale yellow solid is left behind.

The pale yellow solid is lead monoxide, the reddish brown gas is nitrogen dioxide which is produced along with a colourless oxygen gas.



$$2Pb(NO_3)_2 \xrightarrow{heat} 2PbO + 4NO_2 + O_2$$

More examples of a decomposition reaction are :

(i) $2KClO_3\,(s) \xrightarrow{heat} 2KCl\,(s) + 3O_2\,(g)$
[potassium chlorate] [potassium chloride] [oxygen]

(ii) $Cu(OH)_2\,(s) \xrightarrow{heat} CuO\,(s) + H_2O\,(l)$
[copper hydroxide] [copper oxide] [water]

(iii) $CuCO_3\,(s) \xrightarrow{heat} CuO\,(s) + CO_2\,(g)$
[copper carbonate] [copper oxide] [carbon dioxide]

(iv) $2H_2O_2\,(s) \xrightarrow{heat} 2H_2O\,(s) + O_2\,(g)$
[hydrogen peroxide] [water] [oxygen]

Displacement reaction

A reaction in which a more active element displaces a less active element from a compound is called *displacement reaction*. In such reactions, one constituent of the reactant molecule is replaced by another. *For example,*

$$AB + C \rightarrow CB + A$$

Here, C displaces A from the molecule AB, since C is chemically more reactive as compared to A.

(i) A more reactive metal displaces a less reactive metal from its salt solution.

Examples : (a) Zinc being more reactive than copper, displaces copper from copper sulphate solution. The blue copper sulphate solution turns colourless due to the formation of zinc sulphate and a reddish brown copper is deposited on zinc.

$Zn + CuSO_4\,(aq) \rightarrow ZnSO_4\,(aq) + Cu$
(grey) (blue) (colourless) (reddish brown)

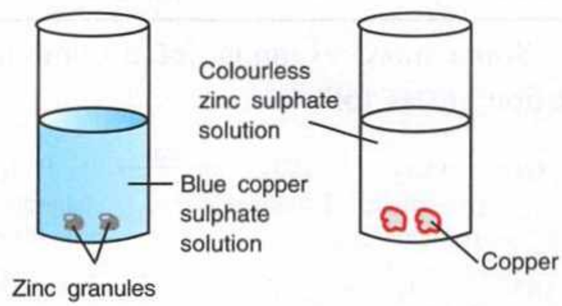


Fig. 6.3 Showing displacement of copper by zinc from copper sulphate solution.

(b) Iron displaces copper from blue copper sulphate solution to form a green iron sulphate solution, since iron is more reactive than copper.

$Fe + CuSO_4 \longrightarrow FeSO_4\,(aq) + Cu$
[grey] (blue) (green) (reddish brown)

(ii) A metal more reactive than hydrogen displaces hydrogen gas from an acid.

Example : Zinc, being an active metal, displaces hydrogen from dilute hydrochloric (or dilute sulphuric) acid.

$$Zn + 2HCl\,(dil) \rightarrow ZnCl_2 + H_2$$
[zinc] [hydrochloric acid] [zinc chloride] [hydrogen]

$$Zn + H_2SO_4\,(dil) \rightarrow ZnSO_4 + H_2$$
[zinc] [sulphuric acid] [zinc sulphate] [hydrogen]

Highly reactive metals like sodium and potassium react with water to displace hydrogen from it and form sodium hydroxide and potassium hydroxide respectively. The reactions are violent due to the very reactive nature of sodium and potassium.

$$2Na + 2H_2O \xrightarrow[\text{reaction}]{\text{vigorous}} 2NaOH + H_2$$
[sodium] [cold water] [sodium hydroxide] [hydrogen]

$$2K + 2H_2O \xrightarrow[\text{reaction}]{\text{vigorous}} 2KOH + H_2$$
[potassium] [cold water] [potassium hydroxide] [hydrogen]

(iii) A more reactive non-metal displaces a less reactive non-metal from the solution of its compound.

Example : Chlorine being more reactive than bromine, displaces it from potassium bromide solution.

$$2KBr\,(aq.) + Cl_2 \longrightarrow 2KCl + Br_2$$
[potassium bromide] [chlorine] [potassium chloride] [bromine]

Displacement reactions help us to distinguish between more reactive and less reactive elements. The one that displaces is more reactive than the other which gets displaced by it. This leads to the formation of metal activity series.

Double displacement reaction

A chemical reaction in which two compounds in their aqueous state exchange their ions or radicals to form new compounds is called a *double decomposition reaction* or a *double displacement reaction. For example,*

$$AB + CD \rightarrow CB + AD$$

Here, AB and CD are the two reactant molecules. They exchange their ions (or radicals) to form two new molecules, CB and AD. Double displacement reactions are of two types :

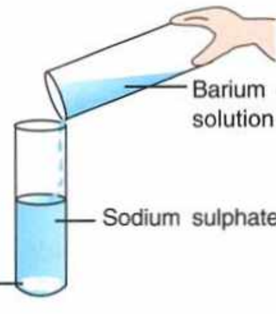
(a) Precipitation reaction

(b) Neutralization reaction

(a) *Precipitation reaction* : A chemical reaction in which two compounds in their aqueous state react to form an insoluble solid (a precipitate) as one of the products is known as a *precipitation reaction.*

Example : When barium chloride solution reacts with sodium sulphate solution, a white precipitate of barium sulphate and a soluble salt, sodium chloride, are obtained.

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4\downarrow^* + 2NaCl(aq)$$
[barium chloride] [sodium sulphate] [barium sulphate] (white precipitate) [sodium chloride]



Some more examples of precipitation reaction are :

(i) $AgNO_3(aq) + HCl \rightarrow AgCl\downarrow + HNO_3(aq)$
[silver nitrate] [hydrochloric acid] [silver chloride] (white ppt.) [nitric acid]

* *A precipitate is an insoluble solid formed when two compounds in their aqueous solution state react chemically with each other. It is indicated by an arrow pointing downwards in a chemical equation.*

(ii)	$CuSO_4(aq)$	$+2NaOH(aq)$	$\rightarrow$	$Cu(OH)_2\downarrow$	+	$Na_2SO_4(aq)$
	[copper sulphate]	[sodium hydroxide]		[copper hydroxide] (blue ppt.)		[sodium sulphate]
(iii)	$Pb(NO_3)_2$	$+2NH_4OH$	$\rightarrow$	$Pb(OH)_2\downarrow$	+	$2NH_4NO_3$
	[lead nitrate]	[ammonium hydroxide]		[lead hydroxide] (white ppt.)		[ammonium nitrate] *(aq)*
(iv)	$CaCl_2(aq)$	$+2NaOH(aq)$	$\rightarrow$	$Ca(OH)_2\downarrow$	+	$2NaCl(aq)$
	[calcium chloride]	[sodium hydroxide]		[calcium hydroxide] (white ppt.)		[sodium chloride]

(b) *Neutralization reaction* : A chemical reaction in which a base or an alkali reacts with an acid to produce a salt and water only is known as a *neutralization reaction.*

Acid + Base/Alkali → Salt + Water

Examples

(a) When an alkali, sodium hydroxide, reacts with hydrochloric acid, it forms a salt, sodium chloride, and water.

$NaOH(aq)$	+	$HCl(aq)$	$\rightarrow$	$NaCl(aq)$ +	H_2O
[sodium hydroxide]		[hydrochloric acid]		[sodium chloride]	[water]

(b) Zinc oxide, a base, reacts with nitric acid to form a salt, zinc nitrate, and water.

$ZnO(s)$	+	$2HNO_3(aq)$	$\rightarrow$	$Zn(NO_3)_2(aq)$ +	H_2O
[zinc oxide]		[nitric acid]		[zinc nitrate]	[water]

Some more examples of a neutralization reaction are :

(i)	$2KOH(aq)$ +	$H_2SO_4(aq)$	$\rightarrow$	$K_2SO_4(aq)$ +	$2H_2O$
	[potassium hydroxide]	[sulphuric acid]		[potassium sulphate]	[water]
(ii)	$CuO(s)$ +	$2HNO_3(aq)$	$\rightarrow$	$Cu(NO_3)_2(aq)$ +	H_2O
	[copper oxide]	[nitric acid]		[copper nitrate]	[water]

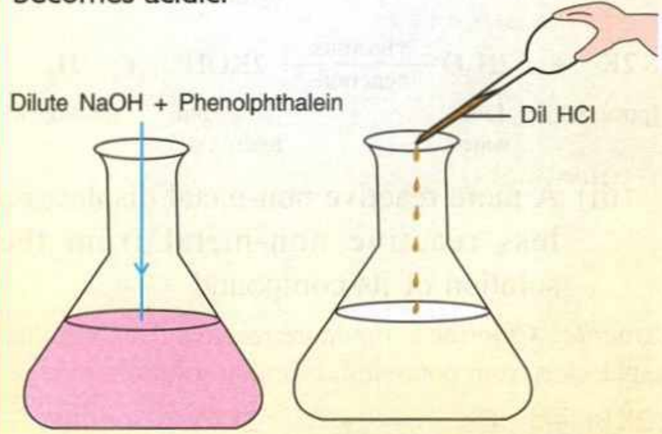
All metallic oxides and metallic hydroxides are called *bases,* whereas those bases that dissolve in water are known as alkalis.

Acids are the substances which when dissolved in water produce hydrogen ions (H^+) as the only positive ions in their solutions.

Activity 5

To show a neutralization reaction.

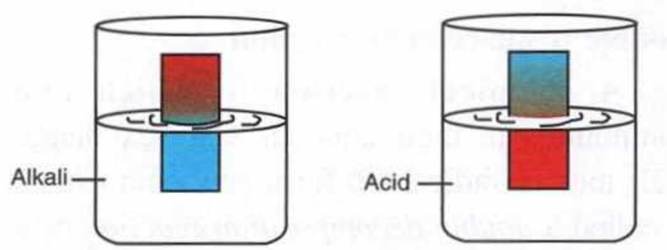
Pour about 50 ml of sodium hydroxide in a conical flask. Add few drops of phenolphthalein to it. The solution turns pink. Now add dilute hydrochloric acid dropwise to it with the help of a dropper. Shake well after adding each drop of acid. A stage comes when the solution becomes colourless. This is the neutralization point where the alkaline solution is completely neutralized by the acid. If more acid is added, the solution becomes acidic.



Solution is pink **Solution turns colourless**

Indicators : These are the organic compounds which show characteristic colours in acidic and basic solutions.

Example : Phenolphthalein, methyl orange, blue and red litmus, etc.



Red litmus changed into blue in an alkaline solution **Blue litmus changed into red in an acidic solution**

INDICATOR	COLOUR IN	
	Acidic solution	**Basic solution**
1. Litmus	Red	Blue
2. Methyl orange	Red or pink	Yellow
3. Phenolphthalein	Colourless	Pink

THE IMPORTANCE OF NEUTRALIZATION REACTIONS IN OUR DAILY LIFE

1. **Indigestion :** The acidity and indigestion can be overcome by taking antacids like milk of magnesia [$Mg(OH)_2$].
2. **Oral hygiene :** Many toothpastes contain bases to neutralize the acids formed in the mouth.
3. **Insect sting :** When a bee stings, it injects an acidic liquid (formic acid) through the skin which can be neutralized by applying a basic calamine solution or baking soda solution or slaked lime.

   But wasp stings are alkaline. They can be neutralized by vinegar which is a weak acid. Lemon juice can also be used.
4. **Soil treatment :** If the soil is acidic, it can be treated with bases like quick lime, slaked lime or chalk to make it neutral. Similarly, basic soils are neutralized by adding sulphate salt.

METAL ACTIVITY SERIES

On the basis of the rate of the reaction of metals with oxygen (air), water and dilute acids, the metals have been arranged in a decreasing order of their chemical reactivity.

A list in which the metals are arranged in a decreasing order of their chemical reactivity is called the metal reactivity series.

Among the commonly known metals, the most active metal (potassium) is kept at the top of the list and the least active metal (platinum) is at the bottom of the list.

Special features of the activity series :

1. The ease with which a metal in solution loses electron(s) and forms a positive ion decreases down the series, *i.e.* from potassium to platinum.
2. Hydrogen is included in the activity series because, like metals, it too loses an electron and becomes positively charged (H^+) in most of the chemical reactions.

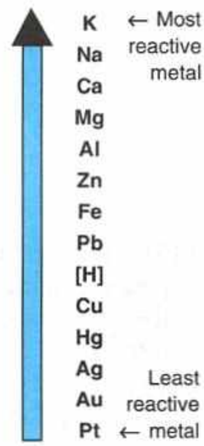


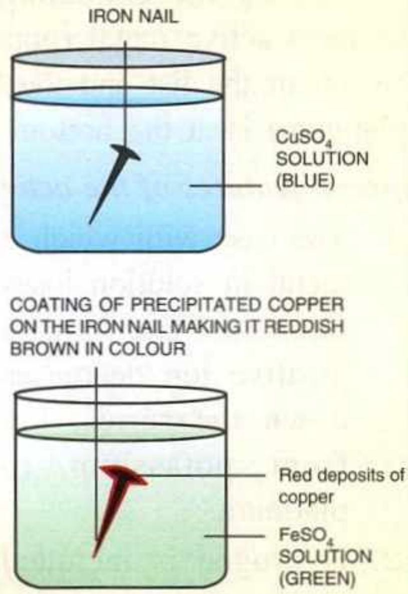
Fig 6.4 Reactivity series of metals

3. The series facilitates the comparative study of metals in terms of the degree of their reactivity.
4. The compounds of the metals (oxides, carbonates, nitrates and hydroxides) too can be easily compared.

Activity 6

To show that iron is more reactive than copper.

An iron nail is placed in blue coloured copper sulphate solution. After some time, a reddish brown coating is seen on the iron nail and the colour of the solution changes gradually from blue to light green. This is because iron, which is more reactive compared to copper, displaces copper from the solution and copper is deposited on the iron nail.



(a) Iron nail dipped in $CuSO_4$ solution

(b) The solution turns pale green due to the formation of $FeSO_4$

$$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$$

(iron) (blue copper sulphate solution) (green ferrous sulphate solution) (copper)

Activity 7

To show that copper is more reactive than silver.

A test tube is half filled with silver nitrate solution and a copper coil is dipped into it. After some time, you will notice that the copper coil gets coated with silver while the solution turns blue. This shows that copper is chemically more reactive than silver.

$$2AgNO_3 + Cu \longrightarrow 2Ag + Cu(NO_3)_2$$

(silver nitrate, colourless) (copper) (silver) (copper nitrate, blue)

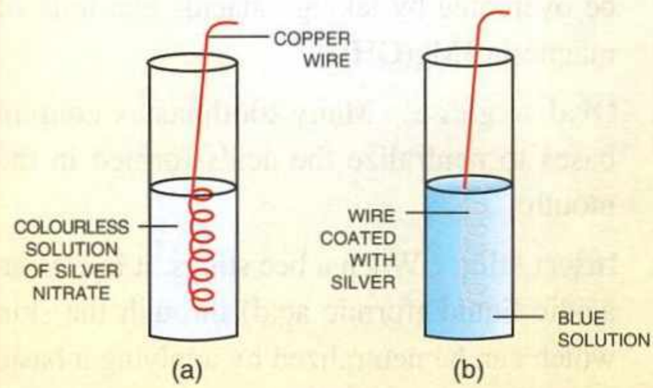


Table showing reactivity of metals with oxygen, water and acids

Elements	*Reaction with oxygen*	*Reaction with water*	*Reaction with acids*
1. K 2. Na 3. Ca	React with oxygen at ordinary temperature to form oxides.	K and Na react vigorously with cold water Ca reacts moderately with cold water.	K and Na react explosively with dilute acids to give hydrogen but Ca reacts less vigorously.
4. Mg 5. Al 6. Zn 7. Fe	Form oxides on heating, but aluminium reacts with it at ordinary temperature.	Mg reacts with hot water or steam; others react with steam only to form oxide and hydrogen.	Mg, Al, Zn and Fe react moderately with acids to produce hydrogen.
8. Pb 9. Cu 10. Hg	Form oxides on very strong heating.	No reaction with hot water or steam.	Pb reacts with conc. HCl to give H_2; Cu and Hg do not react with dilute acids.
11. Ag 12. Au 13. Pt	Do not react with oxygen even on strong heating.	No reaction with hot water or steam.	Do not react with dilute acids.

Some more displacement reactions are as follows :

$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

[Zn is more reactive than Cu]

$Mg + FeSO_4 \rightarrow MgSO_4 + Fe$

[Mg is more reactive than Fe].

$Zn + \underset{\text{(dil.)}}{H_2SO_4} \rightarrow ZnSO_4 + H_2$

(Zn is more reactive than hydrogen)

Non-metals also show displacement reaction. Chlorine is more reactive than iodine.

$$\underset{\text{(Potassium iodide)}}{2KI} + \underset{\text{(Chlorine)}}{Cl_2} \rightarrow \underset{\text{(Potassium chloride)}}{2KCl} + \underset{\text{(Iodine)}}{I_2}$$

TYPES OF CHEMICAL REACTIONS ON THE BASIS OF ENERGY CHANGE

Chemical reactions involve energy changes

A chemical reaction involves the breaking up of chemical bonds between atoms resulting in the absorption of energy in the form of heat and simultaneously the formation of bonds takes place with the release of energy. These two types of energy are different from each other *i.e.* there is either a surplus or a deficit of energy during the reaction. Therefore, in a chemical reaction energy is either absorbed or released.

Depending upon the energy absorbed or evolved, chemical reactions are of two types :

1. Exothermic reaction
2. Endothermic reaction

1. *Exothermic reaction* : A chemical reaction in which heat (a form of energy) is given out is called an *exothermic reaction.* It causes a rise in temperature.

Examples

(a) When carbon burns in oxygen to form carbon dioxide, a lot of heat is produced.

$$\underset{\text{[carbon]}}{C} + \underset{\text{[oxygen]}}{O_2} \rightarrow \underset{\text{[carbon dioxide]}}{CO_2} + \text{Heat}$$

(b) When water is added to quicklime, a lot of heat energy is produced [along with alkaline calcium hydroxide (slaked lime)], which boils the water.

$$\underset{\text{[quicklime]}}{CaO} + H_2O \rightarrow \underset{\text{[slaked lime]}}{Ca(OH)_2} + \text{Heat}$$

Respiration, rusting, burning of fuels like coal, petrol, kerosene, etc. are some common exothermic reactions. **All neutralization reactions are exothermic in nature.**

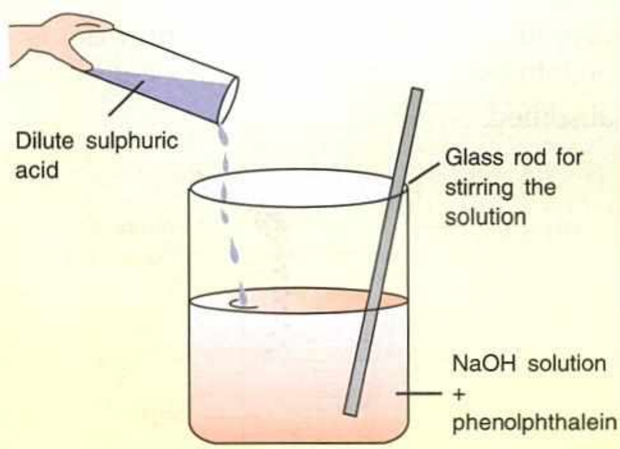
Activity 8

To show that neutralization of an acid with a base is an exothermic reaction.

Take 100 ml of a base sodium hydroxide in a beaker. Add few drops of phenolphthalein to it. The solution turns pink. Now add dilute sulphuric acid to the solution slowly and constantly stir it with a glass rod so that the acid and the base react with each other and the solution again becomes colourless. This indicates that neutralization is complete.

Now touch the beaker with your hand. You will notice that it has become hot.

This is because the neutralization of an acid with a base is an exothermic reaction in which heat energy is released.



2. ***Endothermic reaction*** : A chemical reaction in which heat is absorbed is called an *endothermic reaction*. It causes a fall in temperature.

Examples

(a) When nitrogen and oxygen together are heated to a temperature of about 3000°C, nitric oxide gas is formed. This is an endothermic reaction.

$$\underset{\text{[nitrogen]}}{N_2} + \underset{\text{[oxygen]}}{O_2} \xrightarrow[\text{(3000°C)}]{\text{heat}} \underset{\text{[nitric oxide]}}{2NO\,(g)}$$

(b) Similarly, decomposition of calcium carbonate into carbon dioxide and calcium oxide occurs when it is heated to a temperature of about 1000°C. It is an endothermic reaction.

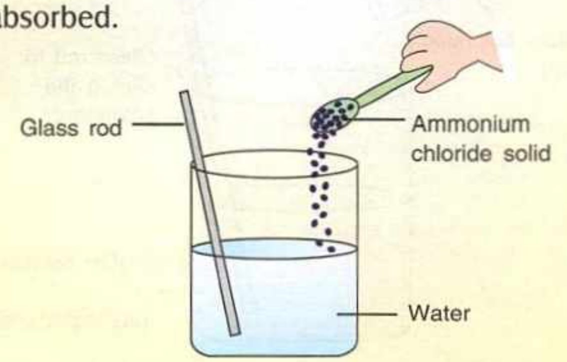
$$\underset{\text{[calcium carbonate]}}{CaCO_3\,(s)} \xrightarrow[\text{(1000°C)}]{\text{heat}} \underset{\text{[calcium oxide]}}{CaO\,(s)} + \underset{\text{[carbon dioxide]}}{CO_2\,(g)}$$

Activity 9

To show that dissolution of ammonium chloride in water is an endothermic reaction.

Take 100 ml of water in a beaker. Add 2-3 spatula of ammonium chloride solid to it and constantly stir with a glass rod to dissolve it in water. Now touch the beaker.

You will notice that the beaker has become colder than it was earlier. This is because dissolution of ammonium chloride is an endothermic reaction in which heat energy is absorbed.



Metallic and non-metallic oxides

An oxide is a compound which essentially contains oxygen in its molecule, chemically combined with a metal or a non-metal.

Oxides are mainly of *two* types :

- Metallic oxides
- Non-metallic oxides

Oxides

- Metallic oxides
  - Basic oxides [*eg.* Na_2O, CaO, MgO, etc.]
  - Amphoteric oxides [*eq.* ZnO, Al_2O_3, PbO, etc.]
- Non-metallic oxides
  - Acidic oxides [*eq.* CO_2, SO_2, P_2O_5, etc.]
  - Neutral oxides [*eq.* H_2O, CO, NO, etc.]

Metallic oxides : Metallic oxides are formed when a metal reacts with oxygen on heating or without heating.

Metals like sodium, potassium, calcium, magnesium, zinc, aluminium, iron, lead, etc. react with oxygen to produce their oxides.

(a) $\underset{\text{(Sodium)}}{4Na} + O_2 \longrightarrow \underset{\text{(Sodium oxide)}}{2Na_2O}$

(b) $\underset{\text{(Calcium)}}{2Ca} + O_2 \xrightarrow{\text{heat}} \underset{\text{(Calcium oxide)}}{2CaO}$

(c) $\underset{\text{(Copper)}}{2Cu} + O_2 \xrightarrow{\text{heat}} \underset{\text{(Copper (II) oxide)}}{2CuO}$

(d) $\underset{\text{(Lead)}}{2Pb} + O_2 \xrightarrow{\text{heat}} \underset{\text{(Lead monoxide)}}{2PbO}$

- Some metallic oxides are also formed on heating metallic carbonates, nitrates, etc.

Examples

(a) Lead nitrate on heating produces lead monoxide.

$$\underset{\substack{\text{[lead nitrate]}\\\text{(white solid)}}}{2Pb(NO_3)_2} \xrightarrow{\text{heat}} \underset{\substack{\text{[lead monoxide]}\\\text{(yellow solid)}}}{2PbO} + \underset{\substack{\text{[nitrogen}\\\text{dioxide]}}}{4NO_2} + \underset{\text{[oxygen]}}{O_2}$$

(b) Lead carbonate on strong heating produces lead monoxide.

$$\underset{\substack{\text{[lead carbonate]}\\\text{(white solid)}}}{PbCO_3} \xrightarrow{\text{heat}} \underset{\substack{\text{[lead oxide]}\\\text{(yellow solid)}}}{PbO} + \underset{\text{[carbon dioxide]}}{CO_2\uparrow}$$

(c) Limestone (calcium carbonate) on heating also decomposes to produce a white solid calcium oxide.

$$\underset{\text{[calcium carbonate]}}{CaCO_3} \xrightarrow{\text{heat}} \underset{\text{[calcium oxide]}}{CaO} + \underset{\text{[carbon dioxide]}}{CO_2}$$

- Most of the metallic oxides are **basic in nature** hence called as **basic oxides**.
- Some of these basic oxides dissolve in water to produce **soluble bases known as alkalis**.

(a) $\underset{\text{[sodium oxide]}}{Na_2O} + \underset{\text{[water]}}{H_2O} \longrightarrow \underset{\text{[sodium hydroxide]}}{2NaOH}$

(b) $\underset{\text{[potassium oxide]}}{K_2O} + \underset{\text{[water]}}{H_2O} \longrightarrow \underset{\text{[potassium hydroxide]}}{2KOH}$

Sodium hydroxide and potassium hydroxide are strong alkalis. They change the colour of indicators. Red litmus turns into blue in an alkaline solution.

All alkalis are bases but all bases are not alkalis because all bases are not soluble in water.

> ***Note :*** Ammonium hydroxide [NH_4OH] is a weak alkali although it does not contain any metal atom in its molecule but it has a group of non-metals with a positive charge which is called as ammonium (NH_4^+) ion.

- Basic oxides react with acids to produce salt and water.

(a) $\underset{\text{[calcium oxide]}}{CaO} + \underset{\text{[dil. hydrochloric acid]}}{2HCl} \rightarrow \underset{\text{[calcium chloride]}}{CaCl_2} + \underset{\text{[water]}}{H_2O}$

(b) $\underset{\text{[copper (II) oxide]}}{CuO} + \underset{\text{[dil. sulphuric acid]}}{H_2SO_4} \rightarrow \underset{\text{[copper sulphate]}}{CuSO_4} + \underset{\text{[water]}}{H_2O}$

- Metallic oxides like zinc oxide, lead monoxide and aluminium oxide are called **amphoteric oxides** because they react both with acids as well as bases to produce salt and water.

(a) $\underset{\text{[zinc oxide]}}{ZnO} + \underset{\text{[hydrochloric acid]}}{2HCl} \rightarrow \underset{\text{[zinc chloride]}}{ZnCl_2} + \underset{\text{[Water]}}{H_2O}$

(b) $ZnO + 2NaOH \rightarrow \underset{\text{[sodium zincate]}}{Na_2ZnO_2} + \underset{\text{[water]}}{H_2O}$

Zinc oxide reacts with dilute hydrochloric acid to produce a salt zinc chloride, while with sodium hydroxide, a base, it produces a salt sodium zincate and water.

(a) $\underset{\text{[lead monoxide]}}{PbO} + \underset{\text{[hydrochloric acid]}}{2HCl} \rightarrow \underset{\text{[lead chloride]}}{PbCl_2} + \underset{\text{[water]}}{H_2O}$

$\underset{\text{[lead monoxide]}}{PbO} + \underset{\text{[sodium hydroxide]}}{2NaOH} \rightarrow \underset{\text{[sodium plumbite]}}{Na_2PbO_2} + \underset{\text{[water]}}{H_2O}$

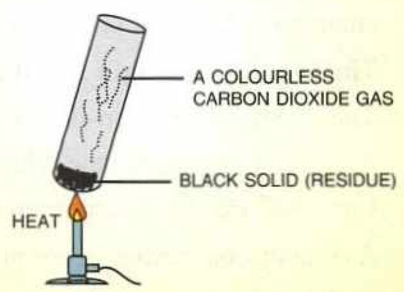
(b) $\underset{\text{[aluminium oxide]}}{Al_2O_3} + \underset{\text{[hydrochloric acid]}}{6HCl} \rightarrow \underset{\text{[aluminium chloride]}}{2AlCl_3} + \underset{\text{[water]}}{3H_2O}$

$\underset{\text{[aluminium oxide]}}{Al_2O_3} + \underset{\text{[sodium hydroxide]}}{2NaOH} \rightarrow \underset{\text{[sodium aluminate]}}{2NaAlO_2} + \underset{\text{[water]}}{H_2O}$

Activity 10

To show formation of an oxide on heating a metal carbonate.

Take some copper carbonate, a green solid, in a hard glass tube and heat it for sometime.



You will observe that the green solid changes into a black solid. Also a colourless, odourless gas is evolved which turns lime water milky when passed through it.

The gas evolved is carbon dioxide and the black residue is copper oxide, a metallic oxide.

Copper carbonate $\xrightarrow{\text{heat}}$ copper oxide + carbon dioxide

$$\underset{\text{(green solid)}}{CuCO_3(s)} \xrightarrow{\text{heat}} \underset{\text{(black solid)}}{CuO(s)} + CO_2(g)$$

Non-metallic oxides : Non-metallic oxides are formed when a non-metal is heated with oxygen. Non-metals like carbon, sulphur, phosphorus, nitrogen, hydrogen combine with oxygen to produce their respective oxides.

(a) $\underset{\text{[carbon]}}{C} + \underset{\text{[oxygen]}}{O_2} \xrightarrow{\text{heat}} \underset{\text{[carbon dioxide]}}{CO_2}$

(b) $\underset{\text{[sulphur]}}{S} + \underset{\text{[oxygen]}}{O_2} \xrightarrow{\text{heat}} \underset{\text{[sulphur dioxide]}}{SO_2}$

(c) $\underset{\text{[nitrogen]}}{N_2} + \underset{\text{[oxygen]}}{O_2} \xrightarrow{\text{high temperature}} \underset{\text{[nitric oxide]}}{2NO}$

(d) $\underset{\text{[Hydrogen]}}{2H_2} + \underset{\text{[oxygen]}}{O_2} \xrightarrow{\text{high temperature}} \underset{\text{[Water]}}{2H_2O}$

- Most of the non-metallic oxides like carbon dioxide, sulphur dioxide, sulphur trioxide, phosphorus pentoxide and nitrogen dioxide are all acidic in nature. They turn moist blue litmus into red. They dissolve in water to produce their respective acids.

(a) $\underset{\text{[carbon dioxide]}}{CO_2} + \underset{\text{[water]}}{H_2O} \rightarrow \underset{\text{[carbonic acid]}}{H_2CO_3}$

(b) $\underset{\text{[sulphur dioxide]}}{SO_2} + \underset{\text{[water]}}{H_2O} \rightarrow \underset{\text{[sulphurous acid]}}{H_2SO_3}$

(c) $\underset{\text{[sulphur trioxide]}}{SO_3} + \underset{\text{[water]}}{H_2O} \rightarrow \underset{\text{[sulphuric acid]}}{H_2SO_4}$

(d) $\underset{\text{[nitrogen dioxide]}}{2NO_2} + \underset{\text{[water]}}{H_2O} \rightarrow \underset{\text{[nitrous acid]}}{HNO_2} + \underset{\text{[nitric acid]}}{HNO_3}$

- Acidic oxides react with bases and alkalis to produce salt and water.

(a) $\underset{\text{[sulphur dioxide]}}{CO_2} + \underset{\text{[sodium hydroxide]}}{2NaOH} \rightarrow \underset{\text{[potassium sulphite]}}{Na_2CO_3} + \underset{\text{[water]}}{H_2O}$

(b) $\underset{\text{[sulphur dioxide]}}{SO_2} + \underset{\text{[potassium hydroxide]}}{2KOH} \rightarrow \underset{\text{[potassium sulphite]}}{K_2SO_3} + \underset{\text{[water]}}{H_2O}$

Non-metallic oxides like water [H_2O], carbon monoxide [CO], nitric oxide [NO], nitrous oxide [N_2O] are neutral in nature. They do not change the colour of indicators.

RECAPITULATION

- A chemical reaction is the process of breaking and making of chemical bonds of the substances to form new substances.
- Chemical reactions are characterized by evolution of gas, change of colour, formation of precipitate, change of state and change of energy.
- The chemical reaction in which heat is given out is called an exothermic reaction.
- The chemical reaction in which heat is absorbed is called an endothermic reaction.
- A catalyst is a substance which alters the rate of a reaction but does not undergo any change itself.
- Enzymes are the complex protein molecules which act as catalysts in biochemical reactions.
- A combination reaction is one in which two or more substances combine together to form a single new compound.
- A decomposition reaction is one in which a substance (compound) is broken down into two or more

simpler substances which can be elements or compounds.

- A displacement reaction is one in which one part of the molecule is replaced by a more reactive element.
- A chemical reaction in which two compounds react in their aqueous solutions by exchanging their radicals is called a double displacement reaction.
- A precipitation reaction is a double displacement reaction in which one of the products is an insoluble substance called precipitate.
- The reaction of an alkali with an acid or vice-versa to form a salt and water is called a neutralisation reaction.
- The list in which metals are arranged in the decreasing order of their chemical reactivity is called the Metal Activity Series.
- Metals react with oxygen to form basic oxides. A basic oxide is a compound that reacts with water to produce a metallic hydroxide. A basic oxide reacts with acids to produce salt and water.
- Soluble bases are called *alkalis,* which turn red litmus paper blue. All alkalis are bases but all bases are not alkalis, since all bases are not soluble in water.
- Some metallic oxides (ZnO, Al_2O_3, PbO) are amphoteric in nature, *i.e.* they react with an acid or a base to produce salt and water.
- Non-metals react with oxygen to form acidic oxides or neutral oxides (water, nitric oxide, etc.). Acidic oxides dissolve in water to produce acids. These oxides react with bases to produce salt and water.

EXERCISE– II

1. Fill in the blanks :

(a) A reaction in which two or more substances combine to form a single substance is called a .... reaction.

(b) A .... is a substance which changes the rate of a chemical reaction without undergoing a chemical change itself.

(c) The formation of gas bubbles in a liquid during a reaction is called ....

(d) The reaction between an acid and a base is called ....

(e) Soluble bases are called ....

(f) The chemical change involving iron and hydrochloric acid illustrates a .... reaction.

(g) In the type of reaction called .... two compounds exchange their positive and negative radicals respectively.

(h) A catalyst either .... or .... the rate of a chemical change but itself remains .... at the end of the reaction.

(i) The chemical reaction between hydrogen and chlorine is a .... reaction.

(j) When a piece of copper is added to silver nitrate solution, it turns .... in colour.

2. Classify the following reactions as combination, decomposition, displacement, precipitation and neutralization. Also balance the equations.

(a) $CaCO_3(s) \xrightarrow{\text{heat}} CaO(s) + CO_2(g)$

(b) $Zn(s) + H_2SO_4 \rightarrow ZnSO_4(s) + H_2(g)$

(c) $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3$

(d) $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

(e) $CuSO_4(aq) + H_2S(g) \rightarrow CuS(s) + H_2SO_4(l)$

(f) $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

(g) $Ca(s) + O_2(g) \xrightarrow{\text{heat}} CaO(s)$

(h) $NaOH + HCl \rightarrow NaCl + H_2O$

(i) $KOH + H_2SO_4 \rightarrow K_2SO_4 + H_2O$

3. Define :

(a) precipitation (b) neutralization

(c) catalyst

4. Explain the following types of chemical reactions giving two examples for each of them:

(a) combination reaction

(b) decomposition reaction

(c) displacement reaction

(d) double displacement reaction

5. Write the missing reactants and products and balance the equations :

(a) $NaOH + \ldots\ldots\ldots\ldots \rightarrow NaCl + \ldots\ldots\ldots\ldots$

(b) $KClO_3 \xrightarrow{\text{heat}} \ldots\ldots\ldots\ldots + \ldots\ldots\ldots\ldots$

(c) $\ldots\ldots\ldots\ldots + HCl \rightarrow NaCl + H_2O + \ldots\ldots\ldots\ldots$

6. How will you obtain ?

(a) Magnesium oxide from magnesium.

(b) Silver chloride from silver nitrate.

(c) Nitrogen dioxide from lead nitrate.

(d) Zinc chloride from zinc.

(e) Ammonia from nitrogen.

Also give balanced equations for the reactions.

7. What do you observe when

(a) Iron nail is kept in copper sulphate solution for sometime ?

(b) Phenolphthalein is added to sodium hydroxide solution ?

(c) Blue litmus paper is dipped in dilute hydrochloric acid ?

(d) Lead nitrate is heated ?

(e) Magnesium ribbon is burnt in oxygen ?

(f) Ammonia is brought in contact with hydrogen chloride gas ?

8. Give reason :

(a) A person suffering from acidity is advised to take an antacid.

(b) Acidic soil is treated with quicklime.

(c) Wasp sting is treated with vinegar.

9. What is meant by metal reactivity series ? State its importance (any two points).

10. What are oxides ? Give two examples of each of the following oxides.

(a) Basic oxide (b) Acidic oxide

(c) Amphoteric oxide (d) Neutral oxide

11. Define exothermic and endothermic reactions. Give two examples of each.

12. State the effect of the following on the surroundings :

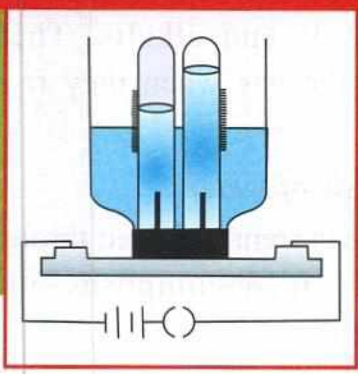
(a) an endothermic reaction

(b) an exothermic reaction.

13. What do you observe when

(a) an acid is added to a basic solution ?

(b) ammonium chloride is dissolved in water ?



Hydrogen

7

Theme : Hydrogen is an important constituent of several compounds. It is found in acids and organic compounds. It acts as a fuel which makes its study useful.

In this chapter you will learn :

- *Hydrogen is an important constituent of water, acids, alkalies, organic compounds, etc.*
- *Preparation of hydrogen by electrolysis of water and by the action of dilute acids on active metals.*
- *Physical and chemical properties of hydrogen, uses and its chemical tests.*
- *Oxidation and reduction reactions.*

LEARNING OUTCOMES

The children will be able to :

- describe the preparation of hydrogen from electrolysis of water.
- prepare hydrogen in the lab, using zinc and acid.
- describe properties and uses of hydrogen.
- correlate concepts of oxidation and reduction with addition and removal of oxygen and removal and addition of hydrogen respectively.

Symbol of hydrogen : H, ***Formula of hydrogen :*** H_2, ***Valency :*** 1, ***Atomic number :*** 1, ***Atomic mass :*** 1·00794 amu.

INTRODUCTION

Hydrogen is the *lightest* of all elements known to mankind. This gas was first prepared by **Robert Boyle** in 1672. **Henry Cavendish** studied the properties of hydrogen in 1776 and called it the *inflammable gas* due to its *combustible* nature. In 1783, **Lavoisier** named the gas "*hydrogen*" meaning *water producer* [hydra-water, gen-producing].

OCCURRENCE

In free state : Hydrogen is the most abundant element in the universe. The outer atmospheres of the sun and the stars consist largely of hydrogen. The outermost layer of the earth's atmosphere contains hydrogen, but in traces. On the earth, free hydrogen occurs in very small quantities, mainly in volcanic gases.

In combined state : Hydrogen is an important constituent of several compounds like water, acids, alkalis, organic compounds, etc. Thus, it is much more common in the form of compounds.

- Water is the most important compound of hydrogen, which covers more than 70% of the earth's surface.
- Acids, alkalis, petroleum, natural gas, etc. are all compounds containing hydrogen.
- Hydrogen is one of the main components of the organic compounds which form the body of animals and vegetable matter.

Note : Organic compounds are carbon compounds that also contain hydrogen. Earlier most of them were obtained from living organisms, hence they are still known as organic compounds.

Today hydrogen is widely seen as the mass fuel of the future, when there will be no petroleum or coal or gases acting as fuel.

Do You Know ?

There would have been no sunlight or heat if there was no hydrogen in the universe.

PREPARATION OF HYDROGEN

The principal sources of hydrogen on the earth are water, acids and alkalis. These compounds liberate the gas when they react chemically.

1. By the electrolysis of water

When an electric current is passed through acidulated water*, it decomposes into hydrogen and oxygen.

$$\underset{\text{(Acidulated water)}}{2H_2O} \xrightarrow[\text{Current}]{\text{Electric}} \underset{\text{(Hydrogen)}}{2H_2} + \underset{\text{(Oxygen)}}{O_2}$$

Hydrogen is collected at the cathode (negative electrode) and oxygen is collected at the anode (positive electrode). The ratio of hydrogen and oxygen thus collected is 2 : 1 by volume.

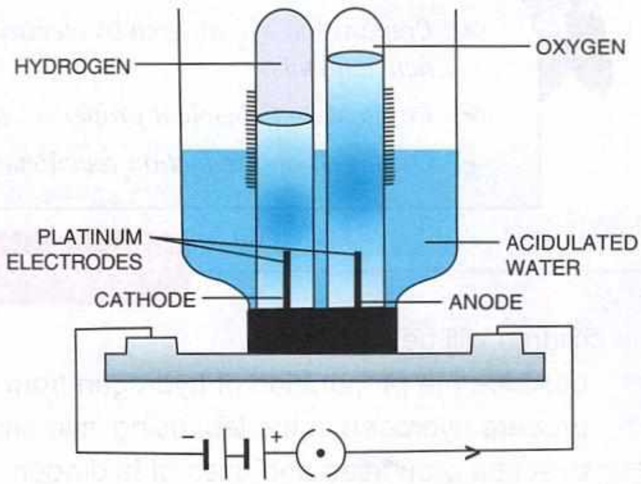


Fig. 7.1 Electrolysis of water.

By this method, hydrogen is prepared on a large scale.

Note : Addition of acid to water makes it a good conductor of electricity.

ELECTROLYSIS AND RELATED TERMS

Electrolysis is a process in which an electric current is passed through an aqueous

* *Water containing a little sulphuric acid is called acidulated water.*

solution or a molten state of a compound to bring about a chemical change.

The passage of electric current takes place because the compound in liquid state or solution state produces free ions. These ions carry charge.

The process involves *two* steps :

- In the first step, the ions present in the solution or molten compound (also called as an electrolyte) migrate towards the solid conductors called electrodes.
- In the second step, these ions release their charges at the electrodes and become neutral.

The terms associated with electrolysis are :

1. Electrolytes : Electrolytes are the compounds in aqueous or molten state that conduct electric current through them because they contain free mobile ions. Electrolytes are of two types.

(a) Strong electrolytes : They dissociate almost completely into ions in their aqueous or molten state. They contain mostly ions in their solutions. *Example :* Sodium chloride, lead bromide, sodium hydroxide, nitric acid, sulphuric acid, hydrochloric acid, etc.

(b) Weak electrolytes : These substances partially dissociate into ions in their liquid states. Hence their solutions contain both ions and as well as molecules. *Example :* Ammonium hydroxide, acetic acid, etc.

2 Non-electrolytes : The compounds that do not conduct electric current in their aqueous solution or molten state and do not decompose into ions are called non-electrolytes. *Example :* Sugar, pure water, alcohol, etc.

3. Electrodes : The two conducting solid poles through which an electric current enters or leaves an electrolyte are called electrodes. They are made up of metals or graphite. The two electrodes are called cathode and anode.

Anode	*Cathode*
1. The electrode connected to the positive terminal of the battery is called *anode.*	1. The electrode connected to the negative terminal of the battery is called *cathode.*
2. Current enters the electrolyte through the *anode.*	2. Current leaves the electrolyte through the cathode.
3. *Anions* migrate and discharge at the *anode.*	3. *Cations* migrate and discharge at the *cathode.*

4. Ions : *Ions* are the electrically charged atoms or groups of atoms formed due to the decomposition of an electrolyte during electrolysis. There are two types of ions, cation and anion.

Anion	*Cation*
1. Anions are negatively charged atoms or group of atoms.	1. Cations are positively charged atoms or group of atoms.
2. During electrolysis, anions migrate towards the anode.	2. During electrolysis, cations migrate towards the cathode.
3. Examples : Cl^-, OH^-, Br^-, SO_4^{2-}, HCO_3^-, CO_3^{2-}, etc.	3. Examples : Na^+, Ba^{2+}, Ca^{2+}, Al^{3+}, Pb^{2+}, NH_4^+, etc.

5. Electrolytic cell : It is the container in which electrolysis takes place.

2. *By the action of dilute acids on active metals*

Metals like magnesium, zinc, iron, etc.

react with dilute hydrochloric (or dilute sulphuric) acid to liberate hydrogen gas along with the formation of their respective salts.

Metal	+ *Dilute acid*	→	*Salt*	+ *Hydrogen*
Mg	+ H_2SO_4	→	$MgSO_4$ Magnesium sulphate	+ H_2
Zn	+ H_2SO_4	→	$ZnSO_4$ Zinc sulphate	+ H_2
Zn	+ 2HCl	→	$ZnCl_2$ Zinc chloride	+ H_2
Fe	+ 2HCl	→	$FeCl_2$ Ferrous chloride	+ H_2

LABORATORY PREPARATION OF HYDROGEN

Hydrogen is prepared in the laboratory by the action of dilute hydrochloric acid or dilute sulphuric acid on granulated zinc.

Balanced equation for the reaction :

Metal	+ *Dilute acid*	→	*Salt*	+ *Hydrogen*
Zn	+ 2HCl	→	$ZnCl_2$	+ $H_2(g)$
Zn	+ H_2SO_4	→	$ZnSO_4$	+ $H_2(g)$

Why is granulated zinc preferred ?

Granulated zinc contains an impurity, copper, which acts as a positive catalyst* for the reaction. That is why granulated zinc is preferred over pure zinc for laboratory preparation of hydrogen.

The set-up for the experiment is as shown in Fig. 7.2.

* *Catalyst is a substance (an element or a compound) that alters the rate of a chemical reaction, but does not itself undergo any change in its composition during the reaction. A positive catalyst increases the rate of a reaction.*

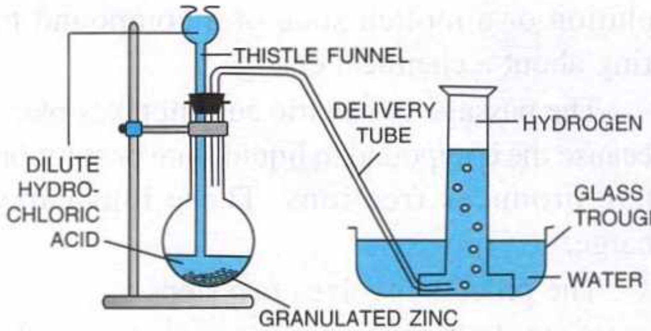


Fig. 7.2 Preparation of hydrogen.

Procedure : Take a few pieces of granulated zinc in a round bottom flask. Fix a two-bored air tight stopper in the mouth of the flask. Pass a thistle funnel through a bore and pass one end of a delivery tube through the other bore. Place the other end of the delivery tube inside a beehive shelf kept in a trough containing water.

Now pour some dilute hydrochloric acid through the thistle funnel. A brisk effervescence occurs. This indicates the liberation of hydrogen gas. Allow the first few bubbles of gas to escape. Then invert over the beehive shelf a gas jar completely filled with water. Hydrogen displaces water and gets collected in the gas jar. This method of collecting gas is known as *"downward displacement of water"*.

Precautions :

(i) Do not collect the first few bubbles of hydrogen gas, since they may contain air too, which is present in the flask and delivery tube.

(ii) Do not bring any flame near the apparatus, since hydrogen is inflammable.

Note : (i) Hydrogen is sparingly soluble in water. Therefore, as it is collected over water, it picks up moisture. Dry hydrogen gas is obtained when it is collected over mercury.

(ii) Since hydrogen is lighter than air, it is possible to collect the gas by downward displacement of air. But it is not safe to do so since a mixture of hydrogen and air can lead to an explosion.

Do You Know ?

Why is dilute nitric acid not used to prepare hydrogen? This is because nitric acid is a strong oxidising agent (even when it is dilute), and oxidises hydrogen gas to convert it into water and defeats the purpose. Dilute hydrochloric acid and dilute sulphuric acid do not oxidise hydrogen into water, hence they are preferred for the preparation of hydrogen gas.

However, a very dilute nitric acid reacts with magnesium and manganese to produce hydrogen.

(i) $Mg + \underset{\text{(very dilute)}}{2HNO_3} \rightarrow Mg(NO_3)_2 + H_2$

(ii) $Mn + \underset{\text{(very dilute)}}{2HNO_3} \rightarrow Mg(NO_3)_2 + H_2$

The metals are arranged in a series in the descending order on the basis of their reactivity which is known as **metal reactivity series**. The most reactive metal is placed at the top and the least reactive one is placed at the bottom of the series. Though hydrogen is not a metal, it is placed in the series as many of its chemical reactions are similar to those of metals.

Metals such as copper, silver and gold, which are below hydrogen in the reactivity series, do not displace hydrogen from water or dilute acids.

Reactivity series

K
Na
Ca
Mg
Al
Zn
Fe
Pb
[H]
Cu
Hg
Ag
Au
Pt

Metals arranged in the descending order of their reactivity.

3. *By the action of water or steam on metals*

(a) From cold water : *The very active metals like sodium, potassium and calcium liberate hydrogen gas from cold water, the other product being a metallic hydroxide.* Sodium and potassium react violently with water, while calcium reacts comparatively slowly.

To prevent direct contact of sodium and potassium with air or water vapour, they are stored in kerosene oil.

$$\textit{Metal} + \underset{\textit{water}}{\textit{Cold}} \rightarrow \underset{\textit{hydroxide}}{\textit{Metallic}} + \textit{Hydrogen}$$

$$2Na + 2H_2O \xrightarrow{\text{Fast}} \underset{\substack{\text{Sodium}\\\text{hydroxide}}}{2NaOH} + H_2$$

$$2K + 2H_2O \xrightarrow{\text{Fast}} \underset{\substack{\text{Potassium}\\\text{hydroxide}}}{2KOH} + H_2$$

$$Ca + 2H_2O \xrightarrow{\text{Slow}} \underset{\substack{\text{Calcium}\\\text{hydroxide}}}{Ca(OH)_2} + H_2$$

(b) From hot water : Magnesium liberates hydrogen from boiling water. However, the reaction is a slow one.

$$Mg + \underset{\substack{\text{Hot}\\\text{water}}}{H_2O} \xrightarrow{\text{Slow}} \underset{\substack{\text{Magnesium}\\\text{hydroxide}}}{Mg(OH)_2} + H_2$$

(c) From steam : When steam is passed over heated metals like magnesium, zinc, aluminium and iron, hydrogen is liberated. The other product is a metallic oxide.

$$\underset{\textit{metal}}{\textit{Hot}} + \textit{Steam} \rightarrow \underset{\textit{oxide}}{\textit{Metallic}} + \textit{Hydrogen}$$

$$Mg + \underset{\text{(steam)}}{H_2O} \rightarrow MgO + H_2$$

$$Zn + \underset{\text{(steam)}}{H_2O} \rightarrow ZnO + H_2$$

$$2Al + \underset{\text{(steam)}}{3H_2O} \rightarrow Al_2O_3 + 3H_2$$

$$\underset{\text{(red hot)}}{3Fe} + \underset{\text{(steam)}}{4H_2O} \rightleftharpoons^{*} \underset{\text{(triferric tetraoxide)}}{Fe_3O_4} + 4H_2$$

Metals other than zinc are not used in the laboratory preparation of hydrogen gas because of the following reasons :

(i) Sodium and potassium react violently with acids.

(ii) Calcium and magnesium are expensive.

(iii) Aluminium has a great affinity for oxygen and forms a protective coating of aluminium oxide, hence it does not react effectively with acids to produce hydrogen, after reacting for sometime.

(iv) Iron needs to be heated and the reaction is reversible.

(v) Lead forms insoluble chloride and sulphate salts with dilute hydrochloric acid and dilute sulphuric acid respectively which stop further reaction.

(vi) Copper, mercury and silver are below hydrogen in metal activity series, hence cannot react with dilute acids to produce hydrogen.

MANUFACTURE OF HYDROGEN

Bosch process : Hydrogen is prepared on large scale by this process. Following are the steps :

(i) Steam is passed over hot coke at 1000°C in a furnace called a converter. As a result, water gas is produced which is a mixture of carbon monoxide and hydrogen gases.

$$\underset{\text{hot coke}}{C} + \underset{\text{steam}}{H_2O} \xrightarrow{1000^\circ C} \underset{\text{water gas}}{(CO + H_2)} - \text{Heat}$$

This reaction is endothermic in nature.

(ii) Water gas is mixed with excess of steam and passed over a catalyst ferric oxide (Fe_2O_3) and a promoter* chromium trioxide (Cr_2O_3).

$$\underset{\text{water gas}}{(CO + H_2)} + \underset{\text{steam}}{H_2O} \xrightarrow[450^\circ C]{Fe_2O_3/Cr_2O_3} CO_2 + 2H_2 + \text{Heat}$$

This reaction is exothermic in nature.

(iii) The products are hydrogen, carbon dioxide and some unreacted carbon monoxide. Hydrogen is separated from carbon dioxide by passing the mixture through water under pressure, in which carbon dioxide gets dissolved leaving behind hydrogen. Carbon dioxide can also be separated by passing it through caustic potash (KOH) solution

$$2KOH + CO_2 \longrightarrow K_2CO_3 + H_2O$$

(iv) To separate carbon monoxide, the gaseous mixture is passed through ammoniacal cuprous chloride in which carbon monoxide dissolves leaving behind hydrogen.

$$CuCl + CO + 2H_2O \longrightarrow \underset{\text{(addition product)}}{CuCl \cdot CO \cdot 2H_2O}$$

Thus, hydrogen gas is obtained.

PROPERTIES OF HYDROGEN GAS

(a) Physical properties :

1. **Nature :** Hydrogen is a colourless, odourless and tasteless gas. It is non-poisonous in nature.

* *Reaction between iron and steam is reversible.*

* ***A promoter activates the catalyst.***

2. **Solubility :** Hydrogen is sparingly soluble in water. One litre of water dissolves about 20 ml of hydrogen gas at ordinary temperature and pressure.

3. **Liquefaction :** Hydrogen gas cannot be easily liquefied. Only at –240°C and at 20 atmospheric pressure, it liquefies into a colourless liquid.

4. **Density :** Hydrogen is the lightest of all the gases. Air is 14·4 times heavier than hydrogen.

Activity 1

To prove that hydrogen gas is lighter than air.

Take a delivery tube and place one of its ends in a liquid soap solution kept in a trough and the other one in a flat bottom jar containing hydrogen gas freshly prepared by the action of dilute hydrochloric acid on zinc granules as shown in figure 7.3. The hydrogen gas enters the soap solution. The soap bubbles containing hydrogen rise upward in the air. The rising soap bubbles prove that hydrogen is lighter than air.

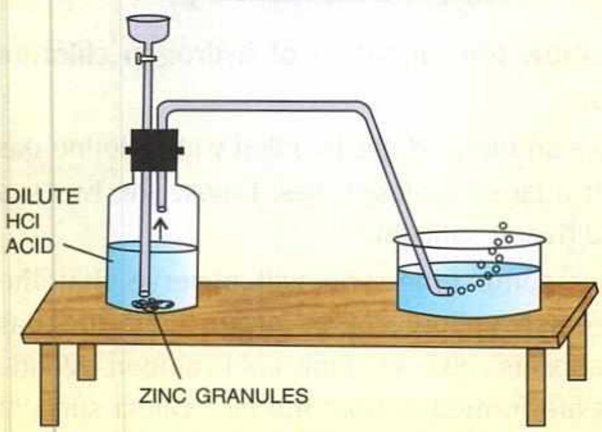


Fig. 7.3 Hydrogen-filled soap bubbles rising upward in the soap solution and into the air shows that hydrogen is lighter than air.

Activity 2

Another example of hydrogen gas being lighter than air.

Take two jars of equal size. One jar contains air while the other one has an equal volume of hydrogen. Place these two jars, with their lids, on the two pans of a balance.

What do you observe ? You will find that the jar containing hydrogen is lighter than the jar with air.

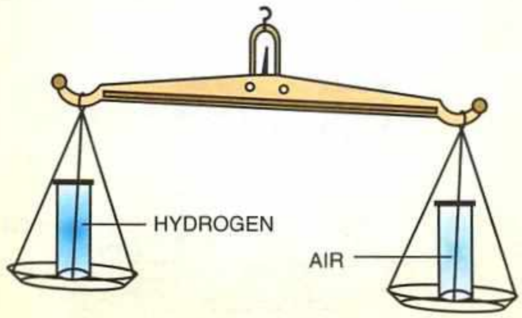


Fig. 7.4 Hydrogen is lighter than air.

Hydrogen gas is lighter than air is best seen in hydrogen-filled balloons that immediately rise up in the air.

(b) Chemical properties :

1. Action with litmus

Hydrogen gas is neutral to litmus, *i.e.* no change is observed in the colour of either moist blue or red litmus paper when it is introduced into a jar containing hydrogen.

2. Combustibility

Hydrogen is **combustible** by nature. In air, pure hydrogen burns silently, with a pale blue flame. But ordinary, *i.e.* impure hydrogen gas burns in air with a pop sound. This is because of the presence of impurities in it. This method is widely used as a test to identify hydrogen. However, hydrogen itself is a **non-supporter of combustion**.

Activity 3

To show that hydrogen is combustible but is a non-supporter of combustion.

Hold a hydrogen gas-filled jar with its mouth downwards. Place a lighted candle and cover it with the jar. The candle gets extinguished but the gas burns with a pop sound. This shows that hydrogen is combustible but is a non-supporter of combustion.

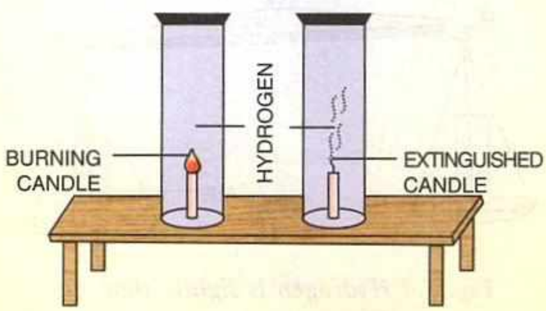


Fig. 7.5 Hydrogen gas does not support combustion.

3. Reactions of hydrogen with some non-metals

(i) *Action with oxygen* : Hydrogen burns in oxygen with a blue flame to form steam, which on condensation forms water. The process is as shown in Fig. 7.6.

$2H_2 + O_2 \rightarrow 2H_2O + \text{Heat}$

The reaction is highly exothermic. In fact, it is this released energy that is harnessed to propel some space rockets. Inside these rockets, liquid hydrogen and liquid oxygen (LOX) are stored together in tanks. Their leakage is prevented at all costs.

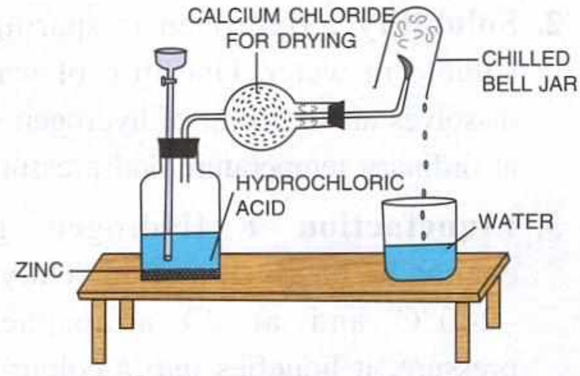


Fig. 7.6 Hydrogen when burns in air forms water.

Note : Hydrogen forms an explosive mixture with air (due to the presence of oxygen in air). If the amount of air is less in the mixture, the explosion is not dangerous *i.e.* the gas burns with only a pop sound. But a large volume of mixture causes a dangerous explosion leading to serious damage and injuries. (*Hydrogen-oxygen mixture is called* **detonating mixture**).

(ii) *Action of hydrogen with chlorine gas* : In diffused sunlight, hydrogen combines with an equal volume of greenish yellow chlorine gas to form colourless hydrogen chloride gas.

$$\underset{\text{(hydrogen)}}{H_2} + \underset{\text{(chlorine)}}{Cl_2} \rightarrow \underset{\text{(hydrogen chloride)}}{2HCl}$$

Activity 4

To show the formation of hydrogen chloride gas.

Place an inverted gas jar filled with chlorine gas over a jar of hydrogen gas. Expose the two jars to diffused sunlight.

After some time, you will observe that the greenish yellow colour of the chlorine gas disappears and white fumes of hydrogen chloride gas are formed in both the jars. Direct sunlight should not fall on the jars because reaction is explosive in direct sunlight.

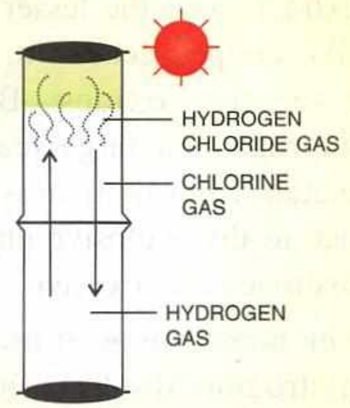


Fig. 7.7 Reaction between hydrogen and chlorine to give hydrogen chloride (g).

(iii) ***Action of hydrogen with nitrogen gas :*** In the presence of iron as a catalyst and at 450°C temperature and 200 atmospheric pressure, hydrogen reacts with nitrogen in the ratio 3 : 1 by volume to produce ammonia gas.

$$3H_2 + N_2 \underset{\substack{450°C \\ 200\text{-}900\ atm}}{\overset{Fe\ (cat)/Mo(pro)}{\rightleftharpoons}} 2NH_3 + Heat$$

In this reaction, molybdenum is used as a promoter.

(iv) ***Action of hydrogen with sulphur :*** When hydrogen is passed through boiling sulphur, hydrogen sulphide gas (H_2S) is formed which smells like rotten egg.

$$\underset{}{H_2} + \underset{\text{(boiling sulphur)}}{S} \rightarrow \underset{\text{(hydrogen sulphide)}}{H_2S}$$

4. Reactions of hydrogen with metals

When metals are heated, hydrogen reacts with certain metals to form metal hydrides.

$$2Na + H_2 \rightarrow \underset{\text{(sodium hydride)}}{2NaH}$$

$$Ca + H_2 \rightarrow \underset{\text{(calcium hydride)}}{CaH_2}$$

Most of the metallic hydrides are unstable in nature.

5. Action of hydrogen as a reducing agent

Hydrogen acts as a good reducing agent. When hydrogen gas is passed over hot metallic oxides of copper, lead, iron, etc., it removes oxygen from them and thus reduces them to their corresponding metals.

$$CuO + H_2 \rightarrow Cu + H_2O$$

$$PbO + H_2 \rightarrow Pb + H_2O$$

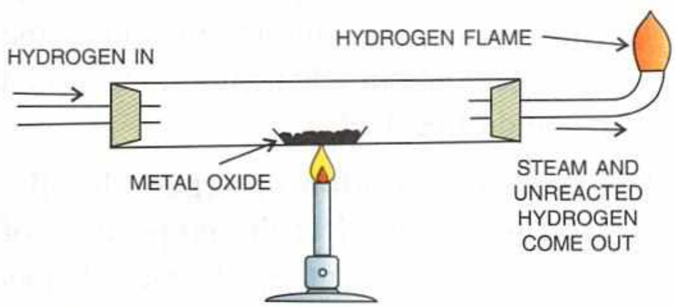


Fig. 7.8 Hydrogen reduces hot metal oxides into their respective metals.

> ***Note :*** The oxides of highly active metals like sodium, potassium, aluminium, calcium and magnesium do not get reduced by hydrogen into their corresponding metals.

TESTS FOR HYDROGEN

1. Ordinary hydrogen burns in air with a characteristic pop sound.
2. Pure hydrogen burns in air or oxygen with a pale blue flame and water is formed upon condensation, which can be tested by sprinkling a few drops of it on white anhydrous copper sulphate. The latter turns into blue hydrated copper sulphate.

USES OF HYDROGEN

1. **For cutting and welding metals :** Hydrogen and oxygen when burnt

together, give a flame known as **oxy-hydrogen flame**. It has a temperature of about 2800°C–3000°C, which is sufficient to melt the metals. The flame is therefore used for cutting and welding of metals.

2. **As a fuel :** Because of its **high heat of combustion**, hydrogen is used as a fuel. Coal gas, water gas ($CO + H_2$) and liquid hydrogen are some significant hydrogen-based fuels. Liquid hydrogen is used as a rocket fuel. Hydrogen can become the mass fuel of the future, replacing the hydrocarbons. In addition, hydrogen is a pollution-free fuel.

3. **For hydrogenation of vegetable oil :** Hydrogen is used in the preparation of *vanaspati ghee* from liquid vegetable fats like groundnut oil, coconut oil, etc. This process is called **catalytic hydrogenation of oils** because it takes place in the presence of finely divided nickel or platinum or palladium acting as a catalyst.

> *Catalytic hydrogenation* is a process by which hydrogen gas is passed through vegetable oils in the presence of a catalyst like Ni, Pt or Pd to convert the oils into vanaspati ghee.

4. **For the manufacture of chemicals :** Hydrogen gas is used extensively in the manufacture of ammonia gas which in turn, is used to produce fertilizers and nitric acid. The gas is also used in the manufacture of hydrochloric acid, methanol, urea, etc.

5. **As a reducing agent in the extraction of metals :** As hydrogen is a good reducing agent, it is used for the extraction of the lesser reactive metals, like copper, lead, tin, *etc*. from their respective oxides. But the use of hydrogen for large-scale extraction of metals from their ores is not common due to the explosive mixture formed by hydrogen and oxygen.

6. **For meteorological purposes :** Earlier, hydrogen-filled balloons equipped with meteorological instruments were sent up in the atmosphere to record temperature, wind speed, air pressure, etc. for weather forecasting. However, due to the inflammable nature of the gas, it has gradually been replaced by helium for filling in weather forecasting balloons.

OXIDATION-REDUCTION OR REDOX REACTION

(a) **Oxidation :** Oxidation is defined as a chemical process that involves :

(i) Addition of oxygen

Element + Oxygen → Oxide

Examples : Carbon reacts with oxygen to form carbon dioxide, hence carbon gets oxidised.

$$\underset{\text{(carbon)}}{C} + \underset{\text{(oxygen)}}{O_2} \xrightarrow{\text{heat}} \underset{\text{(carbon dioxide)}}{CO_2}$$

(ii) Removal of hydrogen

Examples : When hydrogen sulphide reacts with chlorine, it gets oxidised to sulphur due to the removal of hydrogen.

$$\underset{\text{(hydrogen sulphide)}}{H_2S} + \underset{\text{(chlorine)}}{Cl} \rightarrow \underset{\text{(hydrochloric acid)}}{2HCl} + \underset{\text{(sulphur)}}{S}$$

Oxidation is a chemical process that involves addition of oxygen to a substance or removal of hydrogen from a substance.

Oxidising agent : The substance that supplies oxygen or removes hydrogen is called an oxidising agent. *Example :* Oxygen, chlorine, carbon dioxide, nitric acid, etc. The oxidising agent itself gets reduced during the reaction.

(b) **Reduction :** Reduction is the reverse of oxidation. It is a chemical process that involves :

(i) Addition of hydrogen

Examples : When nitrogen reacts with hydrogen under specific conditions, ammonia gas is formed, due to addition of hydrogen to nitrogen. Thus, nitrogen is reduced.

$$\underset{\text{(nitrogen)}}{N_2} + \underset{\text{(hydrogen)}}{3H_2} \rightarrow \underset{\text{(ammonia)}}{2NH_3}$$

(ii) Removal of oxygen

Oxide $\rightarrow$ Element + Oxygen

Examples : When mercuric oxide is heated, removal of oxygen takes place to form mercury, a reduced product.

$$\underset{\text{(Mercuric oxide)}}{2HgO} \xrightarrow{\text{heat}} \underset{\text{(Mercury)}}{2Hg} + \underset{\text{(Oxygen)}}{O_2}$$

Reduction is a chemical process that involves addition of hydrogen to a substance or removal of oxygen from a substance.

Reducing agent : The substance that supplies hydrogen or removes oxygen is called a reducing agent. *Example :* Hydrogen, hydrogen sulphide, carbon, carbon monoxide, etc. The reducing agent itself gets oxidised during the reaction.

Redox reactions : Redox reactions are those in which reduction and oxidation both take place simultaneously, *i.e.* one substance is reduced while the other gets oxidised. Most of the reactions are redox reactions. But double displacement reactions are not redox reactions.

Example :

Reduction [removal of O_2]

$$\underset{\text{Copper oxide}}{CuO} + H_2 \rightarrow \underset{\text{Copper}}{Cu} + H_2O$$

Oxidation [addition of O_2]

In the above reaction, copper oxide gets reduced to copper while hydrogen is oxidised into water. Thus, reduction and oxidation take place simultaneously. (Here, copper oxide acts as an oxidising agent).

Similarly, in redox reaction (i) lead monoxide gets reduced to lead while hydrogen is oxidised into water and in redox reaction (ii) triferric tetroxide is reduced to iron and hydrogen is oxidised into water.

(i) Reduction

$$\underset{\text{Lead monoxide}}{PbO} + H_2 \rightarrow \underset{\text{Lead}}{Pb} + H_2O$$

Oxidation

(ii) Reduction

$$\underset{\text{Triferric tetroxide}}{Fe_3O_4} + 4H_2 \rightarrow \underset{\text{Iron}}{3Fe} + 4H_2O$$

Oxidation

Note : The substance that gets reduced acts as an oxidising agent and the one which gets oxidised acts as a reducing agent.

RECAPITULATION

- Hydrogen is the most abundant element found in the Universe. On the earth, it is found in the form of compounds.
- The chief sources of hydrogen are water, acids and alkalis.
- Hydrogen is prepared by the action of water, acids or alkalis on active metals.
- Electrolysis of water results in the formation of hydrogen and oxygen.
- Hydrogen is a combustible gas but it does not support combustion.
- Hydrogen burns in air to produce water upon condensation.
- Hydrogen acts as a reducing agent and helps in the extraction of metals from their respective oxides.
- Hydrogen burns in air with a pop sound. This helps in testing the gas.
- Hydrogen is used to produce oxyhydrogen flame, in weather forecast balloons and to convert vegetable oils to vanaspati ghee.
- Commercially, hydrogen is prepared by Bosch process.

EXERCISE

1. Fill in the blanks :

(*a*) Hydrogen is .... than air.

(*b*) Hydrogen is .... soluble in water.

(*c*) Hydrogen burns with a .... flame and .... sound is heard.

(*d*) A metal .... hydrogen in the reactivity series gives hydrogen with ....

(*e*) Hydrogen reacts with metal oxides to form .... and ....

(*f*) Oxidation is the removal of .... or addition of ....

(*g*) In redox reactions, oxidation and reduction occur ....

2. Indicate which of the following statements are *true* and which are *false* :

(*a*) Hydrogen molecule is monovalent.

(*b*) The removal of hydrogen from a substance is called reduction.

(*c*) Nitric acid cannot be used to prepare hydrogen by its action on active metals.

(*d*) The reaction between hydrogen and nitrogen to form ammonia is reversible.

(*e*) Zinc can liberate hydrogen from water, acid and alkali solution.

(*f*) Hydrogen is combustible as well as a supporter of combustion.

(*g*) Hydrogen gas is easily liquefiable.

3. Complete and balance the following equations :

(*a*) $H_2 +$ .... $\rightarrow 2HCl$

(*b*) $H_2 + S \rightarrow$ ....

(*c*) $Zn +$ .... $\rightarrow ZnCl_2 + H_2$

(*d*) $CuO +$ .... $\rightarrow Cu +$ ....

(*e*) $Fe + H_2O \rightarrow$ .... $+$ ....

(*f*) $K + H_2O \rightarrow$ .... $+$ ....

4. Give reasons for the following :

(*a*) Hydrogen can be used as a fuel.

(*b*) Though hydrogen is lighter than air it cannot be collected by downward displacement of air.

(*c*) A pop sound is produced when hydrogen gas is burnt.

(*d*) Helium has replaced hydrogen in weather observation balloons.

(e) Nitric acid is not used for the preparation of hydrogen gas.

5. Name the following :
   (a) Two metals which give hydrogen with cold water.
   (b) A metal which liberates hydrogen only when steam is passed over red hot metal.
   (c) The process in which oxygen is added or hydrogen is removed.
   (d) A metallic oxide which can be reduced into metal by hydrogen.

6. (a) Name the chemicals required to prepare hydrogen gas in the laboratory.
   (b) Give a balanced chemical equation for the reaction.
   (c) Draw a neat and well-labelled diagram for the laboratory preparation of hydrogen.
   (d) How is hydrogen gas collected ? Why ?

7. How would you show that hydrogen :
   (a) is a non-supporter of combustion ?
   (b) is lighter than air ?

8. Hydrogen is a good reducing agent. What do you understand by the above statement ? Explain with the help of copper oxide as an example.

9. (a) Name a process by which hydrogen gas is manufactured.
   (b) Give equations for the reactions.
   (c) How is hydrogen separated from carbon dioxide and carbon monoxide ?

10. Match the statements in Column A with those in Column B.

11. State four uses of hydrogen.

12. Define :
   (a) catalytic hydrogenation
   (b) oxidation
   (c) reduction
   (d) redox reaction

12. Multiple Choice Questions

(a) Equal volumes of hydrogen and chlorine are exposed to diffused sunlight to prepare
   (i) hydrogen chloride
   (ii) water
   (iii) sodium hydroxide
   (iv) hydrochloric acid

(b) The metal which reacts with cold water to produce hydrogen is
   (i) magnesium (ii) aluminium
   (iii) calcium (iv) iron

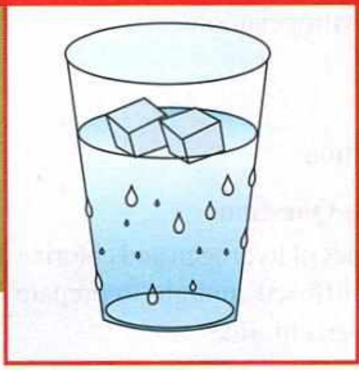
(c) In metal reactivity series, the more reactive metals are at
   (i) top (ii) bottom
   (iii) middle (iv) none

(d) Hydrogen is responsible for producing
   (i) heat and light (ii) hydrogenated oil
   (iii) fertilizers (iv) all of the above

(e) Hydrogen is
   (i) combustible
   (ii) non-combustible
   (iii) supporter of combustion
   (iv) neither supporter nor combustible

(f) Water gas is a mixture of
   (i) carbon monoxide and oxygen
   (ii) carbon monoxide and hydrogen
   (iii) hydrogen and oxygen
   (iv) hydrogen and nitrogen.

Column A	Column B
1. A metal which reacts with cold water to form hydrogen.	1. Reduction
2. A gas which is inflammable and a non-supporter of combustion.	2. Hydrogenation
3. A process in which vanaspati ghee is prepared from vegetable oils.	3. Oxidation
4. The removal of hydrogen or addition of oxygen.	4. Sodium
5. The addition of hydrogen or removal of oxygen.	5. Hydrogen.



Water

8

Theme : Water is one of the most important resources and is a universal solvent. It is important for all living beings — animals, human beings, plants and trees. We use it daily for washing, bathing, drinking and in industries. There are different sources of water such as seas, wells, rivers, lakes, ponds, rain, etc. Water helps in controlling the temperature of the atmosphere.

In this chapter you will learn :

- *Water is an essential constituent of all living things.*
- *There are different sources of water.*
- *Water is a universal solvent.*
- *Meanings of solutions, suspensions and colloids and their differences.*
- *Saturated, unsaturated and supersaturated solutions.*
- *Water of crystallisation, hydrated and anhydrous substances.*
- *Reactivity of metals with water.*
- *Hard and soft water and methods of softening of hard water.*

LEARNING OUTCOMES

The children will be able to :

- describe that water dissolves many substances and it is a universal solvent.
- identify a solution, suspension and colloid on the basis of properties.
- state the differences between saturated, unsaturated and supersaturated solutions.
- describe water of crystallisation.
- write equations of metals with cold water and steam.
- describe hard and soft water.
- discuss the different methods of softening of water.

WATER : AN ESSENTIAL CONSTITUENT OF ALL LIVING THINGS

Water is one of the most essential substances for the existence of life. Since life on earth began in the oceans, and since no living thing can survive without water, it is rightly called the *source of life*.

Water forms a large part of the body mass of all living organisms — 90% of human blood is water. Water has the ability to dissolve a number of substances. Therefore, it serves as the liquid medium in which all reactions within the living body take place.

Fruits and vegetables contain water in them. Even dry-looking substances like wood, peas, beans, gram, *etc*. contain some amount of water.

Table 8.1 : Percentage proportion of water in living things and common eatables

Living things and eatables	*Proportion of water*
Human beings	70%
Elephants	80%
Plants	60%
Leafy vegetables	90%
Potatoes	70%
Tomatoes	95%
Turnips	88%
Water melons	97%
Milk	95%
Eggs	75%
Meat	75%

Activity 1

To show that dry substances contain water.

Take some dry splinters of wood in one test tube, very small pieces of cloth in another test tube, and a few dry bean seeds in the third test tube.

Heat each of them on a flame. What do you observe ? As the substances start getting charred, we see droplets of water getting deposited near the mouth of the test tubes on the inner walls. These droplets are the water vapour released from the dry substances upon heating which then get condensed.

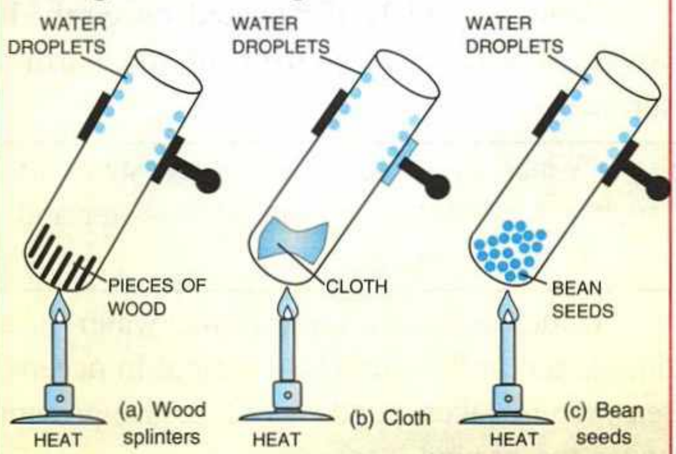


Fig. 8.1 Heating dry substances in test tubes

This activity confirms that even dry substances contain water in them.

Water is important for —

1. everyday uses like cleaning, bathing and washing.
2. drinking in order to satisfy our biological needs.
3. providing medium for all biochemical reactions inside our bodies to take place.
4. regulation of body temperature by the process of sweating and evaporation.
5. growing crops.
6. providing nutrients to aquatic plants and animals.

7. the germination of seeds, growth of plants and in photosynthesis.
8. various manufacturing processes.
9. generating electricity.
10. controlling the earth's climate.

Do You Know ?

Water inside the bodies of animals and humans dissolves minerals, gases and many products of digestion and carries them to a place of need in the body.

SOURCES OF WATER

Water is widely distributed on earth. It covers nearly three-fourths of the earth's surface.

Water exists in all the three states of matter, *i.e.* solid as ice, liquid as water and gas as water vapour.

Under ordinary conditions, water is a liquid, and in this state it is present in oceans, seas, rivers, lakes, ponds, *etc*. It is also present under the ground.

A considerable amount of water is found in the form of ice or snow in polar regions of the world, *viz*. Arctic, Antarctica, and the high mountains. Frost and hail are the other forms of frozen water.

In the atmosphere, water is present in the form of water vapour, which condense to form clouds, fog, mist, *etc*.

Do You Know ?

The electricity produced by using the flow of water from a dam is called hydroelectricity.

Depending on its source and its degree of impurity, natural water is classified into the following types :

1. Ocean and sea water
2. River and lake water
3. Spring and well water
4. Rainwater.

Sources of water

- Surface water
  - Ocean and sea water
  - River and lake water
- Underground water
  - Spring water
  - Well water
- Above surface water
  - Rainwater

1. **Oceans and seas :** Oceans and seas are the largest sources of water, covering nearly 71% of the earth's surface. They contain a very high proportion of dissolved substances, mainly common salt. Therefore, ocean and sea water are the most important sources of salt, but their water is not fit for direct consumption.
2. **Rivers and lakes :** Water in rivers and lakes comes from rain and melting snow. It also contains dissolved impurities like salts, and suspended impurities like clay, sand, twigs and living organisms (including microbes). Rivers and lakes are the most suitable sources of water for domestic and industrial use. But this water needs to be purified before consumption, since it contains impurities and carries germs that can cause diseases.
3. **Springs and wells :** These are sources of underground water. This water is next only to rainwater in its purity. It contains dissolved impurities, but usually has no suspended impurities and germs. The taste

of spring water differs from place to place. Sometimes this water is rich in minerals, with medicinal properties.

4. **Rainwater :** It is the purest form of natural water. It is formed naturally by evaporation followed by condensation of water vapour in the atmosphere. Rainwater dissolves the gases present in the air, and also collects dust particles, as it falls. Some of the important gases dissolved in rainwater are carbon dioxide, oxygen, nitrogen and oxides of nitrogen and sulphur. Rainwater also carries smoke and germs present in the air. Therefore, the first shower of rain contains impurities, but later showers are free from impurities, and therefore safe for drinking.

A major portion of water on the earth is present in the form of snow/ice. Most of the fresh water is frozen in the form of glaciers and polar ice caps.

Activity 2

To show that rainwater does not contain any dissolved impurities but well water contains them.

Take samples of rainwater and well water in two watch glasses. Keep them over two beakers containing water. Heat the beakers. As

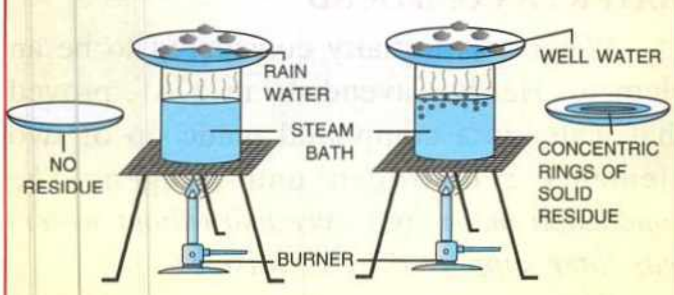


Fig. 8.2

the water in the beakers boils, the samples of water in the two watch glasses evaporate.

What do you observe when both the samples have completely evaporated ?

The watch glass containing rainwater has no residue, but the other containing well water has concentric rings of solid residue. This proves that rain water is purer than well water.

Do You Know ?

— Seas/oceans contain 97·4% of total water available on the earth.

— The process of removal of dissolved salts from sea/ocean water is called desalination.

WATER CYCLE [BALANCE OF WATER IN NATURE]

Water is considered a renewable resource on the earth. It is capable of existing in all the three states of matter. From one state it changes into another rather easily.

$$\text{Ice }(s) \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{Water }(l) \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} \text{Water vapour or steam }(g)$$

(0°C or less) (room temperature) (100°C or more)

This interchange of water from one state to another occurs all the time in nature. *It involves the process of evaporation, condensation and precipitation (rainfall).* This process of *interconversion* helps in balancing the amount of water both in the atmosphere and on the earth's surface, thus supporting both plant and animal lives.

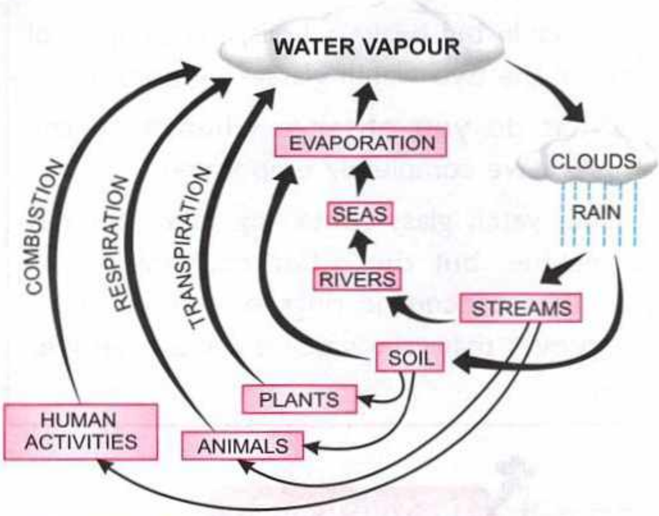


***Fig. 8.3 : Water cycle** (balances water in nature)*

Water is added to the atmosphere by the following natural processes :

1. Evaporation from oceans, seas, rivers, ponds and lakes takes place due to the heat of the sun.
2. Burning of most fuels, mainly fossil fuels.

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

OR

$$\underset{\text{(Butane)}}{2C_4H_{10}} + 13O_2 \xrightarrow{\text{heat}} 8CO_2 + 10H_2O + \text{energy}$$

3. Release of water due to respiration by all living organisms.

$$\underset{\text{(glucose)}}{C_6H_{12}O_6} + 6O_2 \rightarrow 6CO_2 + 6H_2O + \text{Energy}$$

4. Transpiration by plants.

Water that is released into the atmosphere rises up in the form of water vapour which cool down as they rise, and form clouds in the upper regions of the atmosphere. Clouds contain tiny droplets of water, which cool further and these droplets cling to each other and form bigger drops, which fall down as rain. Rainwater gets collected again in seas, rivers, lakes and other water bodies from where it evaporates again. This process goes on and on in a cyclic manner, therefore it is called **water cycle**.

When water vapours cool rapidly well below 0°C, they freeze into snow which falls as snow flakes.

Importance of water cycle

1. Water cycle helps in regulating the climate on the earth.
2. Water cycle makes water available in various forms on the earth.

Do You Know ?

On a humid day, air contains a lot of water vapour, thus reducing the rate of evaporation of surface water from the oceans, seas, rivers, etc. Therefore, during rainy season clothes do not dry quickly.

WATER AVAILABLE FOR HUMAN CONSUMPTION

Water is necessary for all living beings. We cannot stay alive without water.

The earth's surface contains a very large amount of water.

But only 1% of all the water on the earth is suitable for direct use by living organisms. Therefore, this water is very precious to us, and it should be used wisely. The water which is fit for consumption by human beings and other animals is called potable water.

WATER : A COMPOUND

Water was initially considered to be an element. Henry Cavendish, in 1781, proved that water is a compound made up of two elements : hydrogen and oxygen. *The conclusion made by Cavendish (about water) was later supported by Lavoisier.*

Water breaks up into its constituent elements, *viz.* oxygen and hydrogen, when heated to a temperature of 2000°–3500°C or when an electric current is passed through it after adding a small amount of salt or acid in it.

From whatever source water is taken, it has the same chemical composition *i.e.* hydrogen and oxygen gases in 1 : 8 ratio by mass or 2 : 1 by volume.

The molecular formula of water is H_2O and its chemical name is dihydrogen monoxide.

The equation for breaking up of water is,

$$\underset{\text{Water}}{2H_2O} \xrightarrow[\text{current}]{\text{electric}} \underset{\text{Hydrogen}}{2H_2} + \underset{\text{Oxygen}}{O_2}$$

The two gases, hydrogen and oxygen, are identified by their respective chemical tests.

- Hydrogen burns with a pop sound.
- A glowing splinter bursts into a flame in **oxygen**.

PHYSICAL PROPERTIES OF WATER

1. Pure water is a colourless, transparent liquid at room temperature. It is tasteless and odourless too. The water we drink has taste because it contains dissolved minerals and gases.
2. The freezing point of pure water is 0°C and its boiling point is 100°C at normal pressure. Thus, water remains in liquid state over a wide range of temperatures from 0°C to 100°C.

Melting point of ice is also 0°C.

The effect of pressure on the boiling and freezing points of water

The boiling point of water decreases with a decrease in pressure, and increases with an increase in pressure. Therefore, water boils at a lower temperature in the hilly areas, where the atmospheric pressure is lower than in the plains. Also that is why it takes a longer time to cook food in hilly regions.

Do You Know ?

Food is cooked in less time in a pressure cooker because the pressure increases inside the cooker, which also increases the boiling point of water. More heat is required to reach the boiling point, which is sufficient to cook food in a reduced time.

The pressure under a skater's shoes melts the ice, which freezes again when the skater moves on. This results in melting of snow because of an increase in pressure, and the skater glides on water.

The effect of impurities present in water

On freezing point : Any impurity present in water lowers its freezing point.

For example, salt is added to ice to lower its melting point. Such a mixture is called a *freezing-mixture*. The melting point of the freezing mixture is about –15°C.

Similarly, alcohol is added to the water used in car radiators to prevent it from freezing in cold weather.

Why doesn't sea water freeze so easily ?

On boiling point : Any impurity present in water raises its boiling point.

For example : Addition of salt raises the boiling point of water, which helps in cooking food.

3. Water has a **high specific heat**. This means that water needs a large amount

of heat to become hot, and it needs more time to lose a large amount of heat to become cold. In other words, water neither heats up nor cools down quickly.

The specific heat of a substance is the amount of heat required to raise the temperature of a unit mass of that substance by one degree celsius.

Different substances have different specific heat capacities.

Due to high specific heat, water is an excellent cooling agent. It is used as a coolant in car engines, nuclear reactors, machines in factories, etc. It absorbs the heat produced in an engine or a machine or some part of a factory and prevents it from becoming overheated.

Why is water used in room coolers and hot water bags ?

ANOMALOUS EXPANSION OF WATER [RELATIVE VOLUME AND DENSITY OF ICE AND WATER]

The density of a substance is its mass per unit volume, while volume is the space occupied by the mass of the substance. Usually, the solid form of a given substance has greater density, hence it is heavier than the liquid form (for the same volume). *For example :* If you drop a solid iron piece in molten iron, the piece will sink because it is denser than liquid iron.

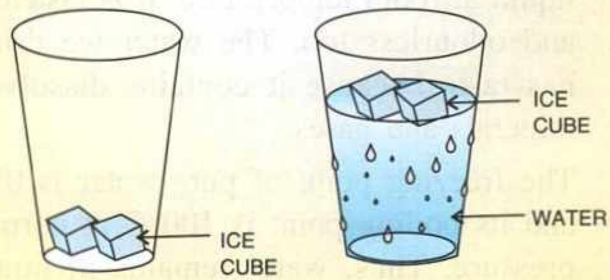
In the case of water, its solid form (ice) is lighter than its liquid form. **This is an anomalous phenomenon** due to anomalous expansion of water.

Generally, when a substance is heated, it expands and its density decreases, and when it is cooled, it contracts and its density increases. Accordingly, when water is cooled, it contracts and its density increases, but only until the temperature reaches 4°C, because on further cooling, water starts expanding, with a decrease in its density, which is an anomalous phenomenon. At 0°C, it becomes ice, and this ice floats on water. *The density of water is, therefore, maximum at 4°C, which is equal to 1 g/cm^3 or 1 kg/L.*

Activity 3

To show that ice is lighter than water.

Take two ice cubes in a glass. Now, add water to it. What do you observe? The ice cubes immediately move to the surface of the water and float.



Freezing lakes : The anomalous expansion of water is a great boon to aquatic life in cold regions. When the temperature falls, the water on the surface of lakes initially cools and contracts. The heavier, colder water sinks, and the surface water again cools. This process continues till the temperature reaches

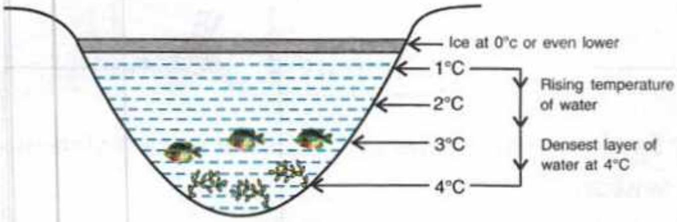


Fig. 8.4 : The topmost layer of water cools and ultimately freezes, while the bottom layers remains liquid at 4°C in lakes and ponds in colder regions of world.

4°C. On further decrease in temperature, the surface water expands and becomes lighter (due to anomalous expansion). Now it does not sink. Therefore, further cooling takes place only at the surface, while the temperature of the lower layers of water does not change. Eventually, the water at the surface changes into ice, which floats over the water below. Ice is a poor conductor of heat and acts like a blanket that protects the bottom layers of the water from the cold air above. This is how fish and other living organisms live inside lakes during winter, even though there is ice at the surface. As the temperature dips further, the thickness of the surface of ice increases.

In the colder parts of the world, water pipes burst when the water inside them freezes in winter. Why ?

Do You Know ?

Large masses of icebergs float on sea water because density of ice is less than that of sea water.

EXERCISE– I

1. Name the four main sources of water.
2. State the importance of water cycle in nature.
3. Why is water very precious for all living beings ?
4. Name the two gases from which water is formed. What is the chemical composition of these two gases in water ? Give the molecular formula of water.
5. What is the effect on boiling point of water when
   (*a*) pressure is increased ?
   (*b*) impurity is added ?
6. Give reasons :
   (*a*) Water is used as a cooling agent.
   (*b*) Water pipes burst in severe winters.
   (*c*) It is difficult to cook in hills as compared to plains.
   (*d*) Ice floats on water.
   (*e*) Sea water does not freeze at 0°C.
7. How does anomolous expansion of water help the aquatic organisms in cold climates ?

8.8 WATER — A UNIVERSAL SOLVENT

Water is often termed as the **universal solvent**, because it dissolves almost all kinds of solids, liquids and gaseous substances. However, the amount of a substance that gets dissolved may vary.

Water can even dissolve the minute particles of the container in which it is kept. It even corrodes our internal body parts if drunk in its purest form (distilled water).

To understand the importance of water as a solvent and its capacity to dissolve substances, it is necessary to understand **solutions**.

"A **solution** is a homogeneous mixture of two or more substances whose composition can be varied." The substances forming solutions are called its components or constituents. One of the components is called a **solvent**, while the other component(s) is called a **solute(s)**.

A **solute** is a subsance that dissolves in other substances. Solute is in a smaller quantity in the solution.

A **solvent** is a medium in which a solute dissolves. It is in a larger quantity in a solution.

Hence,

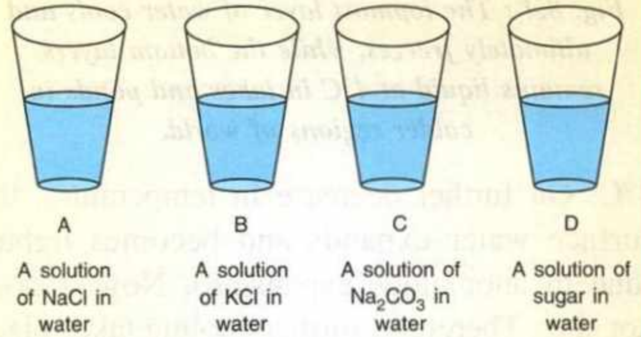
Solution = solute(s) + solvent

Common examples of solutions in our daily life are sugar solution, salt solution, etc. in which sugar and salt are solutes and water is the solvent.

- Solid substances like common salt, sugar, baking soda, washing soda, etc. readily dissolve in water.

Activity 4

To show that different subsances dissolve in water.



Take four glass tumblers and fill half of each of them with water. Mark them as A, B, C and D. Add same quantities of sodium chloride, potassium chloride, sodium carbonate and sugar to A, B, C and D respectively and stir them properly.

What will you observe ? You will observe that in each case the substance dissolves and disappears to give a clear, transparent solution.

Note : The solution made in water is called an **aqueous solution**. While solutions made in any other solvent are called non-aqueous solutions. Some common solvents other than water are alcohol, carbon tetrachloride, etc.

- Liquids like milk, fruit juice, alcohol, vinegar, *etc.* also dissolve in water. Such liquids are called **miscible liquids**.

*The liquids which do not dissolve in water are called **immiscible liquids**, e.g. petrol, diesel, mustard oil, etc.*

- Many gases like oxygen, carbon dioxide and ammonia also dissolve in water.

Do You Know ?

No chemical reaction takes place when a substance (solute) is dissolved in a solvent to prepare its solution. It is a physical change.

Effect of temperature on solubility of a substance (solute) in water :

- Solubility of a solid solute generally increases with an increase in temperature, *e.g.* while making tea, sugar dissolves in hot water easily.
- Solubility of a gas decreases with an increase in temperature. That is why gases dissolved in water can be easily expelled by boiling. Boiled water has a flat taste.

Do You Know ?

Often during summer, fish in shallow ponds die. This is because the water in the pond gets warm due to summer heat, and as a result, the amount of dissolved oxygen in water decreases and the fish in that pond die.

Effect of pressure on the solubility of a substance (solute) in water :

- Pressure has practically no effect on the solubility of a solid (solute) in water.
- In the case of gases, the amount of a gas dissolved in water increases with an increase in pressure and decreases with a decrease in pressure. That is why carbonated drinks (cold drinks, soda water, *etc*.) which contain carbon dioxide are bottled under high pressure as they contain a large amount of gas dissolved in them.

When we open a bottle of a carbonated drink, some of the gas comes out with a hissing sound. Why ?

THE IMPORTANCE OF THE GASES AND MINERALS DISSOLVED IN NATURAL WATER

1. Natural water contains dissolved gases like oxygen, nitrogen and carbon dioxide.
   - Oxygen dissolved in water is the main source of survival for marine life.
   - Carbon dioxide dissolved in water is used by submerged aquatic plants for photosynthesis and in making shells of some marine animals.
   - Nitrogen in water converts into nitrogenous compounds by the action of bacteria and serves as a mode of nourishment for water plants.
2. Some of the salts dissolved in natural water are essential for the proper growth of our bodies. They also add taste to drinking water.

Common salt (sodium chloride) is an essential ingredient of our food. It is added to our food to add taste and for the proper growth and development of our bodies. Sea water contains a large amount of common salt dissolved in it.

CAPACITY OF WATER TO DISSOLVE SUBSTANCES—SATURATED, UNSATURATED AND SUPERSATURATED SOLUTIONS

A solution can have varying compositions of a solute or a solvent or both, *i.e.* different quantities of the same solute can be dissolved in the same volume of water to form solutions.

Is it possible to dissolve any quantity of a solute in the same volume of water ?

This can be understood by the following activity.

Activity 5

Take some water in a glass tumbler. Add some sugar to it and stir properly. You will observe that all the sugar disappears. It means that sugar has been dissolved in the water. Add some more sugar to it. Again it dissolves. Go on adding sugar to the solution formed in the glass, till no more sugar dissolves in it.

This solution is said to be *saturated*, while the solution that can keep dissolving more solute at a particular temperature, is said to be *unsaturated*.

Now, heat the solution. You will observe that the sugar that was not dissolved earlier, now gets dissolved.

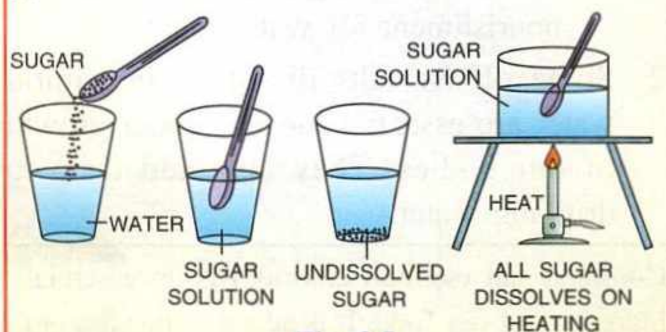


Fig. 8.5

This solution is now said to be *supersaturated*. *This shows that water has a great capacity to dissolve substances.* It also shows that an increase in temperature increases the dissolving capacity of water.

Conclusion : Water can dissolve a substance but only upto a limit.

Saturated solution : A solution that cannot dissolve any more of the solute at a given temperature is called a **saturated solution**.

Unsaturated solution : A solution in which more of the solute can be dissolved at a given temperature is called an unsaturated solution.

Supersaturated solution : A saturated solution that contains more solute than it can hold at room temperature is called supersaturated solution. More solute is dissolved by heating the saturated solution as solubility increases on heating and a supersaturated solution is thus obtained.

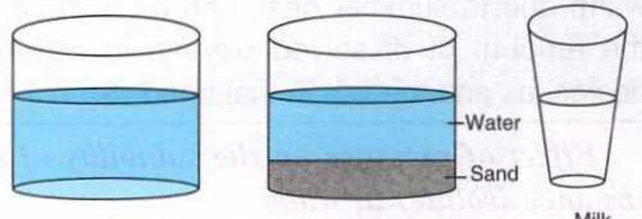
Note : Water dissolves different amounts of different substances under the same conditions.

SOLUTIONS, SUSPENSIONS AND COLLOIDS

You have already studied about, homogeneous and heterogeneous mixtures containing two or more components.

Let us observe the following mixtures :

- Salt dissolved in water
- Sand in water, and
- milk



Salt in water (A solution) Sand in water (A suspension) Milk (A colloid)

You will notice that salt completely dissolves in water to form a clear, transparent homogeneous solution. The particles of salt cannot be seen separately.

Sand does not dissolve in water and settles at the bottom of the container. The two substances, water and sand can be easily distinguished. This is a heterogeneous mixture.

Milk is neither transparent nor its particles settle down. It appears to be homogeneous *i.e.* its particles cannot be distinguished but it is heterogeneous in nature.

Salt and water form a **solution**.

Sand and water is a **suspension**.

Milk is a **colloid**.

Why are they different ?

The difference in a solution, a suspension and a colloid is due to their particle size. A solution has the smallest size of solute particles. Sand has large size of particles due to which it is clearly distinguished from water. The size of the particles of colloids are in between that of a solution and a suspension.

Suspension : A suspension is a heterogeneous mixture in which solid particles, with a size about 10^{-7} m (micron), are dispersed in a medium (mostly liquid medium) that eventually get settled down. *Example :* Chalk in water, mud in water, etc.

Colloid : A colloid is a homogeneous looking heterogeneous mixture in which particles with size ranging from 10^{-10} m to 10^{-7} m are dispersed in a continuous medium. These particles remain suspended throughout the medium and do not settle down. *Example :* Milk, blood, smoke, jellies, butter, ink, soap solution, etc.

Solution : A solution is a homogeneous mixture in which the particle size is less than 10^{-10} m. The solute particles are not visible in it, *e.g.* salt solution.

WATER OF CRYSTALLISATION

When certain solid substances are separated from their aqueous solutions by crystallisation, some water molecules get attached to them. These water molecules give a characteristic shape to the crystals so formed.

> The fixed amount of water which is in a loose chemical combination with one molecule of a substance in its crystal form is called **"water of crystallisation"**.

Such substances which contain water of crystallisation are called **hydrated substances**, *e.g.* $CuSO_4 \cdot 5H_2O$.

The dot in the formula indicates that the water molecule is in a loose chemical combination, and can be removed by heating.

Table 8.2 Differences between solution, suspension and colloids

Solution	Suspension	Colloids
1. Homogeneous.	1. Heterogeneous.	1. Heterogeneous.
2. Particle size less than 10^{-10} m.	2. Particle size greater than 10^{-7} m.	2. Particles size between $10^{-10} - 10^{-7}$ m.
3. Transparent.	3. Opaque.	3. Translucent.
4. Particles of solute are invisible.	4. Solute particles are visible.	4. Solute particles can be seen with the help of a powerful microscope.
5. Particles of solute do not settle.	5. Solute particles settle at the bottom of the container.	5. Solute particles do not settle.
6. Solute particles cannot be filtered. The solution passes easily through a filter paper.	6. Solute particles do not pass through a filter paper.	6. Solute particles pass easily through ordinary filter papers but do not pass through ultra fine filters/nanofilters.

Table 8.3 : Some salts with water of crystallisation

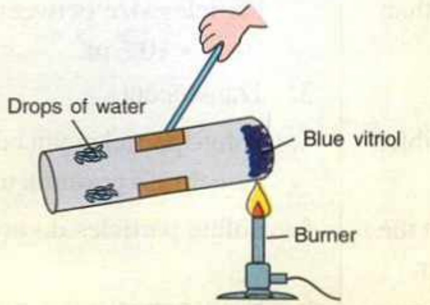
Common name	Chemical name	Formula
1. Blue vitriol	1. Copper sulphate pentahydrate	$CuSO_4 \cdot 5H_2O$
2. Green vitriol	2. Iron (II) sulphate heptahydrate	$FeSO_4 \cdot 7H_2O$
3. White vitriol	3. Zinc sulphate heptahydrate	$ZnSO_4 \cdot 7H_2O$
4. Epsom salt	4. Magnesium sulphate heptahydrate	$MgSO_4 \cdot 7H_2O$
5. Washing soda	5. Sodium carbonate decahydrate	$Na_2CO_3 \cdot 10H_2O$
6. Glauber's salt	6. Sodium sulphate decahydrate	$Na_2SO_4 \cdot 10H_2O$

Since water of crystallisation is loosely held with the salt, on heating upto a certain temperature (100°C), it is driven out of the salt converting the salt from crystalline solid into an anhydrous powder (amorphous solid). This process can be reversed back by adding water to the anhydrous powder. This can be shown by the following activity.

Activity 6

To show that copper sulphate crystals contain water of crystallisation.

Take some blue copper sulphate crystals in a dry test tube. Hold it with a test tube holder and heat it. Hold the tube in a tilted position in such a way that the drops of water formed upon heating may not slip back and crack the test tube at its hot end. What do you observe ?



Showing removal of water on heating

$$CuSO_4 \cdot 5H_2O \xrightarrow{heat} CuSO_4 + 5H_2O$$

Drops of a colourless liquid condense on the upper cooler parts of the test tube, while the blue crystalline solid of copper sulphate turns into white amorphous (powder) solid. This is because the crystals have lost water molecules.

Cool the test tube and trickle the condensed colourless liquid drops back to the white solid. You will observe that the solid again turns into blue. Hence, the colourless liquid is water.

This proves that copper sulphate crystals contain water of crystallisation. It also gives the salt its blue colour.

Crystal : A *crystal* is a homogeneous solid of definite geometrical shape in which particles are arranged symmetrically

Crystallisation : The process by which crystals of a substance are separated by cooling its hot supersaturated solution is called **crystallisation**.

Anhydrous substances : The substances which do not contain water of crystallisation are called **anhydrous substances**.

- They may be crystalline in nature.

  Example : Common salt (NaCl), potassium nitrate (KNO_3), sugar ($C_{12}H_{22}O_{11}$), etc.

- They may be amorphous powdered solids.

  Example : Calcium carbonate ($CaCO_3$), zinc carbonate ($ZnCO_3$), copper oxide (CuO), etc.

Efflorescence : Substances which lose water of crystallisation when exposed to air are called efflorescent substances and the phenomenon is called **efflorescence**.

Washing soda is a white crystalline substance but changes into white powder on exposure to air.

Another example is Glauber's salt ($Na_2SO_4 \cdot 10H_2O$).

Deliquescence : There are substances which absorb water when exposed to atmosphere, become moist and lose their crystalline form. They are water soluble, hence, ultimately turn into solutions by absorbing water. Such substances are called deliquescent and the phenomenon is called deliquescence.

Example : Caustic soda ($NaOH$), magnesium chloride ($MgCl_2$) calcium chloride ($CaCl_2$), Iron (III) chloride ($FeCl_3$), etc.

> On exposure to air, table salt ($NaCl$) turns moist and ultimately forms a solution especially during rainy season because it contains impurities like magnesium chloride and calcium chloride which are deliquescent. Sodium chloride is not deliquescent.

Hygroscopy : Certain substances absorb moisture (water vapour) from the atmosphere without dissolving in it. Such substances are called hygroscopic substances and the phenomenon is called hygroscopy.

Example : Concentrated sulphuric acid (H_2SO_4), phosphorus pentoxide (P_2O_5), quicklime (CaO), anhydrous calcium chloride ($CaCl_2$), silica gel, etc.

Such substances are used as **drying agents** to remove moisture from other substances without chemically reacting with them.

Silica gel pouches are very commonly used to absorb moisture and to keep things dry. They are often kept in unused water bottles, with camera lenses, shoe boxes, loose tablets, etc. to keep them dry. These pouches are ideal to reuse throughout, in places at home where there is excess of moisture.

Chemical properties of water

1. Pure water is neutral to litmus which means that no change in the colour of blue or red litmus solution is observed when treated with water.
2. *Action of heat or electrolysis :* Water is a highly stable compound. However, when heated above 2000°C or when electric current passes through water, it decomposes into hydrogen and oxygen gases.

$$2H_2O \xrightarrow[\text{or electrolysis}]{2000°C} 2H_2 + O_2$$

3. *Action of water on elements :*

   A. **On metals :** Water reacts with active metals like sodium, potassium, magnesium, calcium, iron, etc. under different conditions to produce hydrogen gas.

$$\text{Potassium} + \text{Water} \xrightarrow[\text{fast}]{\text{very}} \text{Potassium hydroxide} + \text{Hydrogen}$$

$$2K + 2H_2O \longrightarrow 2KOH + H_2$$

Sodium + Water (cold) $\xrightarrow{\text{fast}}$ Sodium + Hydrogen hydroxide

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2(g)$$

Calcium + Water (cold) $\xrightarrow{\text{slow}}$ Calcium + Hydrogen hydroxide

$$Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2(g)$$

Magnesium + Steam $\longrightarrow$ Magnesium + Hydrogen oxide

$$\underset{\text{(steam)}}{Mg + H_2O} \longrightarrow MgO + H_2(g)$$

Iron (hot) + Steam $\longrightarrow$ Triferric + hydrogen tetraoxide

$$\underset{\text{(hot)}}{3Fe} + \underset{\text{(steam)}}{4H_2O} \rightleftharpoons Fe_3O_4 + 4H_2(g)$$

The reaction of metals with water under different conditions indicates the reactive nature of metals and enables us to arrange the metals on the basis of their reactivity which is called metal activity series. Here, metals are arranged in decreasing order of their reactivity, *i.e.* most reactive metals are at the top and least reactive at the bottom of the series.

Metal activity series (of some common metals)

Decreasing activity ↓

K
Na
Ca
Mg
Al
Zn
Fe
Pb
(H)
Cu
Hg
Ag
Au
Pt

B. **On non-metals :**

(i) Non-metals, like carbon (coke), react with steam to produce water gas, which is an important industrial fuel.

Coke (red hot) + Steam $\xrightarrow{1000°C}$ Water gas.

$$\underset{\text{(red hot)}}{C} + H_2O \xrightarrow{1000°C} \underset{\text{Water gas}}{[CO + H_2]}$$

(ii) Chlorine gas dissolves in water to produce chlorine water, which contains hydrochloric acid and releases oxygen gas on exposure to light.

$$Cl_2 + 2H_2O \xrightarrow{\text{sunlight}} 2HCl + O_2$$

4. *Action of water on compounds :*

A. **On metallic oxides :** Water dissolves metallic oxides like sodium oxide, potassium oxide, *etc.* to produce metallic hydroxides, also known as alkalis. Soluble bases are called alkalis.

Sodium oxide + Water $\longrightarrow$ Sodium hydroxide (alkali)

$$Na_2O + H_2O \longrightarrow 2NaOH$$

Potassium oxide + Water $\longrightarrow$ Potassium hydroxide (alkali)

$$K_2O + H_2O \longrightarrow 2KOH$$

Calcium oxide + Water $\longrightarrow$ Calcium hydroxide (weak alkali)

$$CaO + H_2O \longrightarrow Ca(OH)_2$$

Magnesium oxide + Water $\longrightarrow$ Magnesium hydroxide

$$MgO + H_2O \longrightarrow Mg(OH)_2$$

B. **On non-metallic oxides :** Oxides like carbon dioxide, sulphur dioxide, nitrogen dioxide, *etc.* dissolve in water to produce acidic solutions.

Carbon dioxide + Water $\longrightarrow$ Carbonic acid

$$CO_2 + H_2O \longrightarrow H_2CO_3$$

Sulphur dioxide + Water $\longrightarrow$ Sulphurous acid

$$SO_2 + H_2O \longrightarrow H_2SO_3$$

Test for Water

1. Water can be tested by its boiling point (100°C) and freezing point (0°C) if it is pure.
2. When a few drops of water are added to anhydrous copper sulphate powder, its colour changes from white to blue.

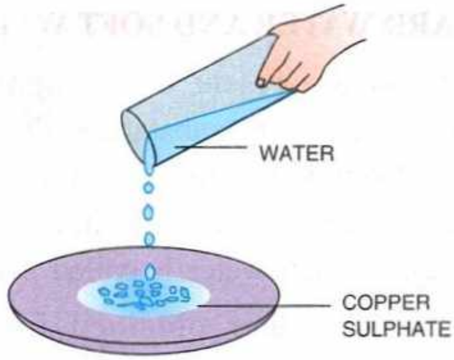


Fig. 8.6 : Adding water to white copper sulphate salt turns the salt blue

3. When a few drops of water are added to blue anhydrous cobalt chloride, its colour changes to pink.

EXERCISE – II

1. Explain the terms :
   (*a*) Solution (*b*) Solute
   (*c*) Solvent

2. What is meant by
   (*a*) Unsaturated (*b*) Saturated and
   (*c*) Supersaturated solutions.

3. How do the solubility of a solid and a gas gets affected by
   (*a*) Increase in temperature ?
   (*b*) Increase in pressure ?

4. Differentiate between :
   (*a*) Solution and suspension
   (*b*) Suspension and colloid

5. Define 'water of crystallisation'. Give *two* examples with formulae.

6. Give two examples for each of the following :
   (*a*) Hydrated substances
   (*b*) Crystalline anhydrous substances
   (*c*) Drying agents
   (*d*) Deliquescent substances
   (*e*) Efflorescent substances
   (*f*) Colloids
   (*g*) Solvents other than water.

7. What do you observe when :
   (*a*) Blue vitriol is heated ?
   (*b*) Washing soda is exposed to air ?
   (*c*) Blue litmus solution is added to water ?

8. Give reason :
   (*a*) Silica gel pouches are kept in unused water bottles.
   (*b*) Table salt becomes moist during rainy season.
   (*c*) On opening a bottle of a cold drink, a fizz sound is heard.

9. Give balanced chemical equations for the reaction of water with
   (*a*) Sodium (*b*) Iron
   (*c*) Carbon dioxide (*d*) Sodium oxide

10. What is metal activity series ?

11. Name the gas produced when
   (*a*) steam is passed over hot coke.
   (*b*) chlorine is dissolved in water and exposed to sunlight.
   (*c*) a piece of calcium is added to water.
   (*d*) when fossil fuel is burnt.

8.9 HARD WATER AND SOFT WATER

The water present in different natural sources has different substances dissolved in it. It has been found that water drawn from certain sources forms a lather with soap rather easily. Such water is called **soft water**.

Whereas, water obtained from some other sources does not easily form a lather with soap, rather it forms a white sticky scum or a precipitate. This water is called **hard water**.

Hard water is of two types :

(i) ***Temporary hard water :*** Water which has bicarbonates of calcium and magnesium dissolved in it is temporary hard water. This kind of hardness can be easily removed by boiling.

WATER
On the basis of lather formation with soap

- **Soft Water** (forms lather easily with soap
- **Hard Water** (does not form lather easily with soap)
  - **Temporary hard water** (Due to bicarbonates of calcium and magnesium) Hardness can be removed by simple physical method, *i.e.* by boiling
  - **Permanent hard water** (Due to chlorides and sulphates of calcium and magnesium) Hardness can be removed by chemical method, *i.e.* by adding sodium carbonate

(ii) ***Permanent hard water :*** Water which has sulphates and chlorides of calcium and magnesium dissolved in it is called permanent hard water. This hardness cannot be removed by boiling.

Table 8.4 Differences between hard water and soft water

Hard water	Soft water
1. Hard water does not form lather with soap easily.	1. Soft water easily forms lather with soap.
2. It is not good for making solutions and washing clothes.	2. It is good for making solutions and washing clothes.
3. Hard water is formed due to the presence of bicarbonates, chlorides and sulphates of calcium and magnesium.	3. Soft water does not contain bicarbonates, chlorides and sulphates of calcium and magnesium.
4. Slightly hard water is fit for drinking but too much hard water is not fit for drinking.	4. Soft water is not fit for drinking purposes.

DISADVANTAGES OF HARD WATER

1. It is not fit for laundry. It results in the wastage of soap.
2. It leaves a substance at the bottom of the container when it is boiled. This substance is called scale or fur which damages the container. Hence, it is not good for boilers and in industries.
3. Hard water is not suitable for leather and textile industries. Clothes do not take dyes uniformly while printing them using hard water.
4. Hard water is not very good for cooking.

REMOVAL OF HARDNESS OF WATER

The hardness of water is removed to make it consumable in homes, laundries, factories, *etc*. When dissolved salts of calcium and magnesium are removed from water, it becomes soft. This is called softening of water. The following are some of the methods used to remove the hardness of water.

1. **Boiling :** This method helps to remove only the temporary hardness of water. When temporary hard water is boiled, the bicarbonates of calcium and magnesium break up to form their respective insoluble carbonates. These can be filtered out so that water becomes soft.

$$\underset{\text{(soluble calcium bicarbonate)}}{Ca(HCO_3)_2} \xrightarrow{\text{boiled}} \underset{\text{(insoluble calcium carbonate)}}{CaCO_3} + \underset{\text{(water)}}{H_2O} + \underset{\text{(carbon dioxide)}}{CO_2(g)}$$

$$\underset{\text{(soluble magnesium bicarbonate)}}{Mg(HCO_3)_2} \xrightarrow{\text{boiled}} \underset{\text{(insoluble magnesium carbonate)}}{MgCO_3} + \underset{\text{(water)}}{H_2O} + \underset{\text{(carbon dioxide)}}{CO_2(g)}$$

2. **By adding sodium carbonate (washing soda) :** The permanent hardness of water is removed when such water is treated with a small quantity of sodium carbonate. It reacts with the soluble chlorides and sulphates of calcium and magnesium to form their respective insoluble carbonates. These can be removed by filtration and then the water becomes soft. Sodium sulphate or sodium chloride formed after the reaction does not affect the soap.

$$\underset{\text{(sodium carbonate)}}{Na_2CO_3} + \underset{\text{(calcium sulphate)}}{CaSO_4} \longrightarrow \underset{\text{(insoluble calcium carbonate)}}{CaCO_3\downarrow} + \underset{\text{(soluble sodium sulphate)}}{Na_2SO_4}$$

$$\underset{\text{(sodium carbonate)}}{Na_2CO_3} + \underset{\text{(magnesium chloride)}}{MgCl_2} \longrightarrow \underset{\text{(insoluble magnesium carbonate)}}{MgCO_3\downarrow} + \underset{\text{(soluble sodium chloride)}}{2NaCl}$$

Activity 7

Take some tap water in beaker A. Add a pinch of magnesium chloride to it and stir so that it gets dissolved. Now the tap water has become hard water. Pour half of that water in another beaker B.

(i) Add some soap solution in beaker A and stir.

(ii) Add some washing soda ($Na_2CO_3 \cdot 10H_2O$) in beaker B and stir. A white insoluble solid ($CaCO_3$) is formed. Allow it to settle and then filter it. Now add some soap solution to this water and stir.

What are your observations ?

You observe that in beaker A no lather is formed but in beaker B, lather is formed.

This is because beaker A contains hard water but the water in beaker B has become soft on adding washing soda.

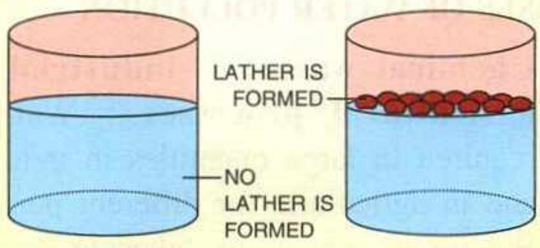


Fig. 8.7

Do You Know ?

Soap is a compound which is present in many washing cakes, but detergents are different compounds from soap. Detergents act in the same way as a soap but they do not form scum with hard water. They do not contain the soap compound in them. Hence they are also called as "soapless soap."

WATER POLLUTION

Water pollution is the contamination of water with substances which makes it unfit for many types of uses, *i.e.* undesirable changes in the physical, chemical and biological conditions of water that make it unfit for human consumption, is called water pollution.

The pollution of fresh water sources is one of the most serious environmental problems faced by the world today. The water present in rivers, ponds, lakes and streams, comes from rain and the melted snow of the mountains. As it flows down the plains, it picks up many dissolved and suspended impurities. This water further increases in its impurity with the addition of the waste products from homes, agricultural lands and industries. These waste products are very harmful. They are called **pollutants,** because they make the water impure and polluted.

CAUSES OF WATER POLLUTION

1. **Chemical wastes of industrial and agricultural processes :** Water is required in large quantities in industries and in agriculture for different purposes. But often, care is not taken to keep the water fit for consumption. Most of the industries dump their wastes in rivers and even into the sea without prior treatment.

   Fertilizers, pesticides, insecticides and other chemicals used in cultivated fields, get dissolved in water when it flows across these fields. Acidic water from mines and acid rain also pollute the water in rivers and seas. These chemicals are dangerous for aquatic life, and they make the water unfit for human consumption.

2. **Thermal wastes from nuclear and thermal power plants :** Power plants use large amounts of water for cooling purposes. They discharge the resultant hot water, often containing chemicals, into water streams. This results in an increase in the temperature of water, which is injurious for fish and other aquatic life forms.

3. **Sewage and garbage :** The discharge of sewage and garbage into river water is another major cause of water pollution. Water becomes unsafe for drinking and other uses due to people urinating, defecating or washing in it.

Water borne diseases : Different kinds of germs which cause diseases grow in polluted water. Almost two-thirds of all the diseases in India are water borne such as,

Fig. 8.8 Pictures showing water pollution

typhoid, hepatitis, cholera, diarrhoea, dysentery, etc.

PREVENTION OF WATER POLLUTION

1. We should spread awareness and make sanitary facilities available in rural areas and city slums so as to discourage people from defecating in the open.
2. Domestic sewage should be treated before being discharged into rivers.
3. The solid matter separated from sewage can be used to generate biogas, an important alternative fuel.
4. The waste products of industries should be treated before they are discharged into rivers and other water bodies.
5. We should wisely use substances like detergents, pesticides, polythene, *etc*. They are non-biodegradable and they pollute the environment, including water, to an undesirable extent.
6. Washing and cleaning of utensils, clothes, bathing of animals and human beings should be avoided in or near the sources of water like rivers, lakes, *etc*.
7. Trees and plants must be planted along the banks of rivers.
8. Put a covering on the well.
9. Purification of water bodies should be carried out from time to time.
10. Dead bodies of animals should be disposed off in a hygienic way.

EFFECT OF WATER SCARCITY ON PLANTS

You have already learnt that water is needed by plants for various purposes. If water is not available to plants for long, they will dry out and the earth will lose its green character.

Human beings are dependent on plants. Plants provide food, oxygen, wood and many other useful products. Moreover, trees help to bring rain. If there are no trees, droughts will become more frequent. Therefore, without plants, there will be no life on earth.

WATER MANAGEMENT

Water is a precious natural resource. In order to make water easily available to us in future, we have to conserve it today. A few ways by which water can be conserved are as follows :

1. Construct dams and reservoirs to control flood and collect water.
2. Do not leave taps running while brushing teeth, shaving or washing clothes.
3. Recycle water in industries and use it as many times as possible, before disposing it.
4. Plant trees to slow down the flow of rainwater on land and to increase seepage of water into soil.
5. Practise rainwater harvesting.
6. Repair leaking taps and pipes at the earliest.
7. Recycle polluted water and use it for irrigation.
8. Use biodegradable fertilizers and pesticides.
9. Farmers should use better methods of irrigation like drip irrigation in which water is supplied to plants drop by drop instead of filling the entire field with water.

RECAPITULATION

- Water is the source of life for all living organisms.
- In nature, water occurs in all the three states of matter, *i.e.* ice, liquid and vapour, but mostly in liquid form.
- Water found in nature comes from :
  (i) oceans and seas (ii) rivers and lakes (iii) springs and wells (iv) rain
- Rainwater is the purest form of water. Sea water is very impure, as it contains a very high proportion of dissolved salts.
- Potable water should be free from suspended impurities and harmful germs, but it should contain some dissolved minerals and gases for taste and health purposes.
- Water is a compound with the molecular formula H_2O.
- The boiling point and the freezing point of water are affected by change in pressure and presence of impurities.
- 0°C is also called triple point because water exists in all its three states at this temperature.
- The specific heat of water is higher than that of any other liquid. Consequently, it is used as a cooling agent.
- Water has minimum volume and maximum density at 4°C. It starts expanding below 4°C. This is called anomalous expansion of water.
- Water is an excellent solvent. The gases dissolved in water have biological importance. They enable aquatic life to sustain itself.
- Depending upon the size of the particles of a substance mixed with solvents, mixtures can be homogeneous like solutions, or heterogeneous like suspensions and colloids.
- Solution is unsaturated if it dissolves more solute and saturated if no more solute is dissolved in the solution at a given temperature.
- A solution is supersaturated if it contains more solute than it can hold at a given temperature.
- Some substances contain water molecules in loose combination with them in their crystals. These water molecules are called water of crystallisation.
- All crystalline substances do not contain water molecules in their chemical structure.
- Anhydrous substances do not contain water of crystallisation, they are used as drying agents.
- Metals react with cold water, hot water and steam depending on their reactivity.
- Water may form lather with soap easily, or it may not do so easily. Depending on this, water may be 'soft' or 'hard'. Hardness of water can be removed by boiling or by chemical treatments.
- Water pollution is a serious problem. It is necessary to control pollution and conserve water.

EXERCISE – III

1. Define :
(a) Soft water *(b)* Hard water

2. *(a)* Name the compounds responsible for
(i) temporary hardness
(ii) permanent hardness of water

(b) Suggest one method for the removal along with the reactions for
(i) temporary hardness
(ii) permanent hardness of water

3. What are the main causes of water pollution ? How can it be controlled ?

4. Name three water borne diseases.
5. Give reasons :
   (a) Alcohol is mixed with water and is used in car radiators.
   (b) Icebergs float on ocean water.
   (c) Carbonated drinks are bottled under high pressure.

OBJECTIVE TYPE QUESTIONS

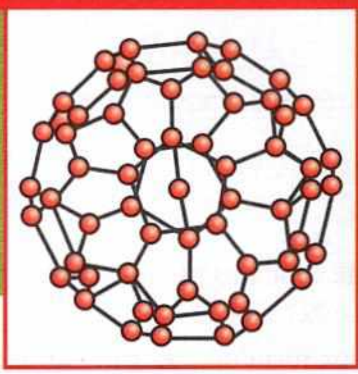
1. Fill in the blanks :
   (a) Water has .... density and .... volume at 4°C.
   (b) Freezing mixture contains .... and ....
   (c) The solubility of a gas in water .... with rise in temperature and .... with rise in pressure.
   (d) .... is the purest form of natural water.
   (e) Use of excessive .... by farmers causes water pollution.
   (f) Boiling removes the .... hardness of water.
   (g) Water turns the colour of anhydrous copper sulphate ....
   (h) The sticky substance formed when soap is added to hard water is ....
2. Match the statements in column A with that in column B.

Column A	Column B
(a) deliquescent substance	(i) sodium bicarbonate
(b) hygroscopic substance	(ii) magnesium chloride
(c) efflorescent substance	(iii) conc. sulphuric acid
(d) substance causing temporary hardness	(iv) washing soda

3. Give one word/words for the following statements :
   (a) Water fit for human consumption ....
   (b) The harmful substances dissolved in water ....
   (c) The change of states of water from one form to another ....
   (d) The gaseous form of water found in air ....
   (e) A mixture of common salt and ice ....
   (f) A substance which does not contain water in its chemical structure ....
   (g) A property due to which a substance absorbs water without dissolving ....
   (h) Water molecules in loose chemical combination with other substances ....

MULTIPLE CHOICE QUESTIONS

1. Two gases found dissolved in natural water are
   (a) oxygen and carbon dioxide
   (b) hydrogen and oxygen
   (c) sulphur dioxide and hydrogen
   (d) chlorine and ammonia
2. Temporary hardness of water can be removed by
   (a) filtering (b) boiling
   (c) loading (d) none of the above
3. The ultimate source of all water on the earth is
   (a) oceans and seas
   (b) spring and wells
   (c) rivers and lakes
   (d) rain
4. Colloids have the particle size ranging between
   (a) 10^{-7} to 10^{-10} m (b) 10^{-10} to 10^{-12} m
   (c) 10^{-7} to 10^{-5} m (d) 10^{-12} to 10^{-15} m



Carbon And Its Compounds

9

Theme : Products such as paper, wooden furniture, soaps, food items are used extensively in daily life activities. These are made up of carbon as one of their elements. The fuel that is used in cars and trucks is also made of carbon. It is a constituent of all plants and animals. In fact, a large number of compounds are made up of carbon. It is a very versatile element. Children will learn the importance of carbon and some of its compounds.

In this chapter you will learn :

- *Allotropy and allotropes of carbon.*
- *Crystalline and amorphous allotropes — diamond, graphite, coal, coke, etc.*
- *Uses of diamond, graphite, coke, coal, soot.*
- *Laboratory preparation, properties and uses of carbon dioxide.*
- *Physical, chemical and acidic properties of carbon dioxide.*
- *Fire extinguishers.*
- *Properties of carbon monoxide as a reducing agent and its poisonous nature.*

LEARNING OUTCOMES

The children will be able to :

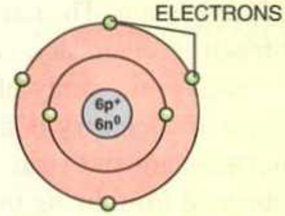
- explain the term allotropy and describe different allotropes of carbon.
- state the properties of graphite and diamond.
- prepare carbon dioxide in the laboratory.
- describe the uses of carbon dioxide.
- demonstrate different reactions of carbon dioxide with lime water and litmus solutions.

Symbol of carbon : C, *Atomic number* : 6, *Mass number* : 12, *Valency* : 4

INTRODUCTION

Carbon is the very basis of life. All living things whether plants or animals and organic matter (which is derived from the dead remains of plants and animals and their waste products) contain carbon. **Therefore it is not surprising that carbon is one of the most widely distributed elements on the earth.** In fact it is the third most important element, after oxygen and hydrogen, for the existence of life on the earth.

The name *carbon* is derived from the Latin word *carbo* (meaning *coal*).



Orbital structure of carbon. It has four electrons in its valence shell, i.e. the valency of carbon is 4

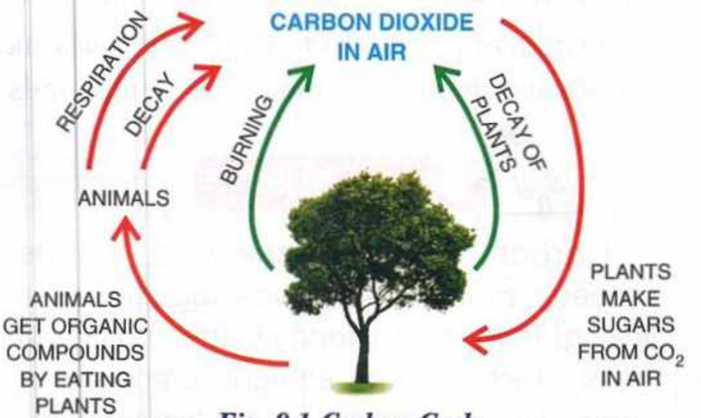


Fig. 9.1 Carbon Cycle

OCCURRENCE

Carbon occurs in the earth's crust in free as well as in combined state.

In the *free state,* it occurs as coal, diamond and graphite.

In the *combined state,* carbon occurs in :

(i) the atmosphere as carbon dioxide gas.

(ii) natural water as dissolved carbon dioxide.

(iii) natural gas and petroleum.

(iv) food nutrients like carbohydrates (starch, sugar), fats, proteins, vitamins, *etc.*

(v) carbonates and bicarbonates such as chalk, limestone, marble ($CaCO_3$)*, calamine ($ZnCO_3$), washing soda ($Na_2CO_3 \cdot 10H_2O$) and baking soda ($NaHCO_3$). Shells of many aquatic animals are made up of carbonates.

(vi) Natural fibres like cotton, silk, etc.

Note : Carbon is a unique non-metal with widely different forms and properties. Mankind has known carbon in the form of charcoal and soot ever since the discovery of fire.

Fig. 9.2 Different forms of carbon

* *$CaCO_3$ is the chemical formula for all the three substances, chalk, marble and limestone.*

Although carbon constitutes only 0·03% of the earth's crust, it forms an enormously large number of compounds. The number is so large that an entire branch of chemistry, called **organic chemistry**, is devoted to the study of carbon and its compounds. Organic chemistry is the term given to this branch because most of these carbon compounds are obtained from living organisms.

In fact, for the convenience of study, all compounds are classified into two classes :

(i) organic compounds

(ii) inorganic compounds.

All organic compounds essentially contain carbon as a constituent.

A few examples of useful organic compounds are starch, wax, vinegar, alcohol, dyes, detergents, soaps, plastics, clothing materials like nylon, polyester, silk, wool and cotton, as well as paper, polythene, perfumes, disinfectants and many medicines.

Table 9.1 : Names of some common compounds containing carbon

Organic Compounds		
Class	**Name**	**Chemical Formulae**
Hydrocarbons	Methane	CH_4
	Butane	C_4H_{10}
	Ethene	C_2H_4
	Ethyne	C_2H_2
	Benzene	C_6H_6
	Naphthalene	$C_{10}H_8$
Alcohols	Methyl alcohol	CH_3OH
	Ethyl alcohol	C_2H_5OH
Acids	Acetic acid (vinegar)	CH_3COOH
	Formic acid	$HCOOH$
	Oxalic acid	$(COOH)_2$
Sugars	Glucose	$C_6H_{12}O_6$
	Sucrose	$C_{12}H_{22}O_{11}$

Inorganic Compounds	
Name	**Chemical Formulae**
Carbon dioxide	CO_2
Carbon monoxide	CO
Calcium carbonate	$CaCO_3$
Sodium carbonate	Na_2CO_3
Copper carbonate	$CuCO_3$
Zinc carbonate	$ZnCO_3$
Sodium bicarbonate	$NaHCO_3$
Calcium bicarbonate	$Ca(HCO_3)_2$

In addition, petroleum is a mixture of carbon compounds found in a liquid state. Also, vegetable oils are the liquid forms of carbon compounds.

Oxides of carbon (CO_2, CO), carbonates and bicarbonates belong to inorganic compounds.

Do You Know ?

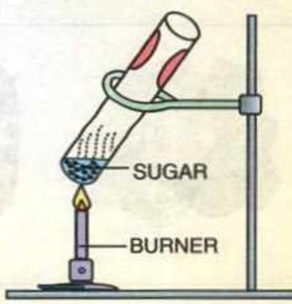
Carbon atoms possess a unique property to link themselves together (self linking) to form very long chains of different sizes which may be straight, branched or cyclic. This property is called 'catenation'.

Activity 1

To show the presence of carbon in sugar.

Take some sugar in a test tube. Heat it for some time. You will observe that sugar first melts, then turns brown, and finally gets charred and turns black. This black substance is carbon. This experiment proves that sugar contains carbon.

Most of the substances that leave a black residue after heating contain carbon. *e.g.* coal, wood, paper, etc.



Sugar turns black on heating

A. Allotropy in carbon

ALLOTROPY AND ALLOTROPES OF CARBON

Allotropy is defined as the phenomenon due to which an element exists in two or more forms in the same physical state with identical chemical properties but with different physical properties. Such forms of an element are known as its allotropic forms or *allotropes*. Allotropes contain the same kind of atoms but due to the differences in the arrangements of the atoms, they give different looks and have different physical properties.

Carbon exhibits allotropy. The chart given below outlines the different allotropes of carbon. The allotropes of carbon are divided into two types :

(i) crystalline allotropes

(ii) amorphous or non-crystalline allotropes

These two broad types are sub-divided into a number of specific allotropes.

All these forms of carbon differ in their physical properties, but when burnt in the presence of oxygen, they all produce carbon dioxide with the release of heat.

$$\text{Diamond/graphite/coal/coke} + \text{Oxygen} \xrightarrow{\text{heat}} \text{Carbon dioxide} + \text{Heat}$$

$$C + O_2 \xrightarrow{\text{heat}} CO_2 + \text{Heat}$$

The difference in their physical properties is due to their different structures.

> Phosphorus and sulphur also exhibit allotropy.
>
> Allotropes of
>
> (1) *Phosphorus* : Yellow phosphorus, red phosphorus, white phosphorus.
>
> (2) *Sulphur* : Rhombic, monoclinic, plastic, colloidal and flower of sulphur.

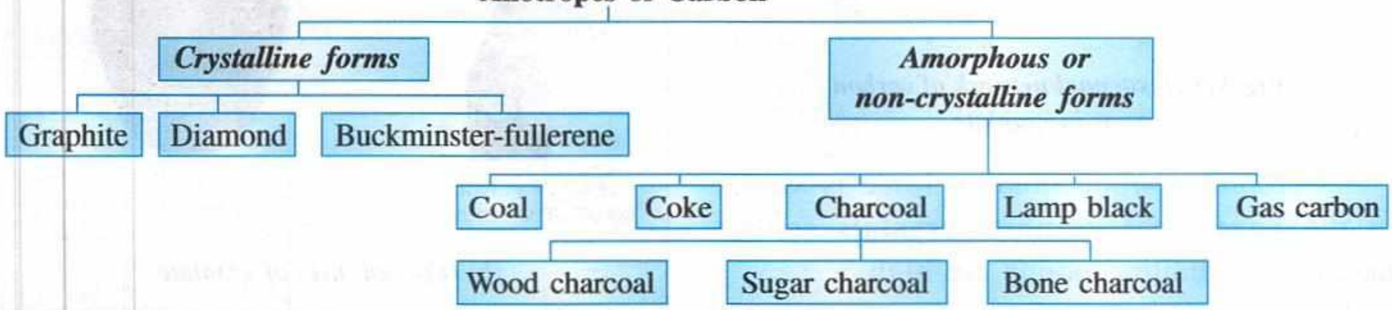
> **WHAT ARE CRYSTALS ?**
>
> A crystal is a homogeneous solid in which particles (atoms, molecules or ions) are arranged in a definite pattern due to which they have a definite geometrical shape with a plane surface.
>
> *For example* : Sodium chloride and sugar both are crystalline with a definite cubical shape for their crystals.

CRYSTALLINE ALLOTROPES OF CARBON

Graphite

Graphite is a natural crystalline form of carbon. Natural graphite is extensively found as the mineral *plumbago* (a mineral of lead) in both Sri Lanka and Siberia. In India, graphite is found in Jammu and Kashmir, Bihar, Odisha, Rajasthan and West Bengal.

Allotropes of Carbon



Preparation of graphite : Pure graphite is prepared artificially by heating **powdered coke** mixed with a little sand and Iron (III) oxide in an electric furnace to a temperature of about 3000°C.

$$\underset{\text{(Sand)}}{SiO_2} + \underset{\text{(Coke)}}{3C} \xrightarrow[\text{ferric oxide}]{3000°C} 2CO + \underset{\text{(Silicon carbide)}}{SiC}$$

$$SiC \xrightarrow{\text{heat}} Si + C\ \text{(graphite)}$$

Structure of graphite : In a graphite molecule, each carbon atom is linked with three neighbouring carbon atoms, thus forming a flat *hexagonal arrangement of atoms.* These hexagonal groupings of carbon atoms are arranged as layers or sheets piled one on top of the other. The layers are held together by weak forces such that they can slide over one another. That is why graphite is soft and slippery and can be used as a lubricant in machines and in pencil leads. Also, in a graphite molecule, one valence electron of each carbon atom remains free, thus making graphite a good conductor of electricity.

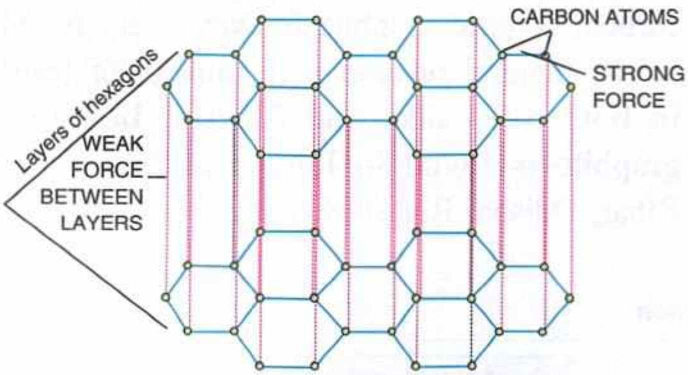


Fig. 9.3 Hexagonal network of carbon atoms in graphite

Carbon – carbon binding force is strong in the hexagonal graphite crystals so the melting and boiling points are high.

Properties of graphite :

1. Graphite is a greyish black, opaque substance, with a metallic lustre.
2. Its density is 2·39 g/cm^3.
3. It is stable to heat and has a very high melting point of 3700°C.
4. It is soft and greasy to touch.
5. It leaves a black mark on paper.

Activity 2

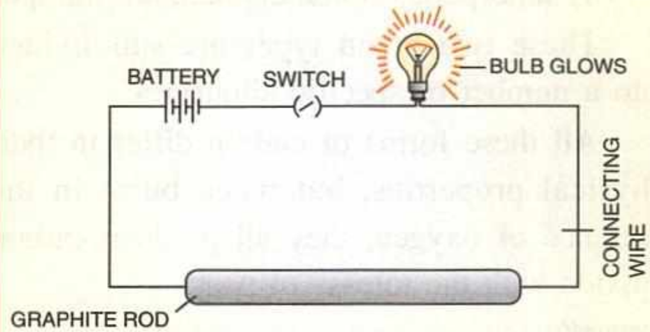
To show that graphite is a good conductor of electricity.

Take a graphite rod. Connect it to a battery, a bulb and a switch with the help of connecting wires. Now close the circuit with the help of the switch.

What do you observe?

The bulb starts glowing.

This is because graphite conducts electricity.



(a) An electric circuit (using graphite rod)

Graphite is used in many other ways too.

LEAD OF PENCILS IS MADE UP OF GRAPHITE

GRAPHITE CRUCIBLES ARE USED TO MELT METALS AND TO KEEP MOLTEN METALS

(b) Different uses of graphite

6. It is a good conductor of heat and electricity.

Uses of graphite :

1. With petroleum jelly to form **graphite grease** (a lubricant).
2. For making the electrodes of electric furnaces.
3. For making crucibles for melting metals due to its high melting point.
4. For making carbon brushes for electric motors.
5. For making pencil leads because it leaves a black mark on paper. Pencil leads are made by mixing graphite with variable quantities of clay. Pencil leads can be sharpened easily because layers of graphite crystals can slide over one another due to weak forces between them.
6. It is used in nuclear reactors as a moderator to slow down the speed of neutrons.
7. For making artificial diamond.

Diamond

Diamond is perhaps the purest form of carbon. It occurs in all shapes and sizes. Diamonds are found in South Africa, Brazil, Namibia, Russia, Australia, U.S.A. and India. In India, diamonds are found at *Goleconda* in Karnataka and at *Panna* in M.P.

Formation of natural diamonds : Natural diamonds are formed by the action of high pressure and temperature on the carbon present in the earth at depths of 150 km or so. They are mostly brought to the surface by volcanic eruptions. Diamond bearing rocks are called *kimberlite rocks,* after the Kimberley Mines in South Africa.

Preparation of artificial diamonds : Synthetic or artificial diamonds are made from graphite. Graphite is subjected to very high temperature (about 3000°C) and pressure in absence of air. The diamonds produced under such conditions are rather small.

Value of diamonds : The value of a diamond as a gem depends upon :

(i) its weight

(ii) the impurities present in it.

The weight of a diamond is expressed in terms of carats [1 carat = 0·2 g].

Some famous natural diamonds are :

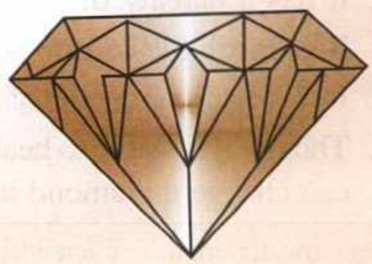
(i) KOHINOOR
(ii) PITT DIAMOND] mined in India

(iii) CULLINAN (the biggest diamond ever found, mined in South Africa).

Colourless, transparent diamonds are the costliest because they have negligible impurity. The value of a diamond decreases with an increase in the impurities present in it.

Do You Know ?

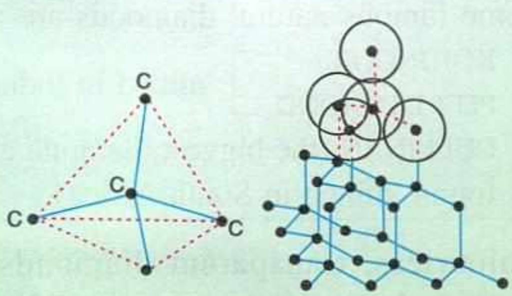
It is the presence of small traces of metallic oxide and salt that imparts distinct colours to diamonds. These coloured diamonds are called gems. Diamond gems of the darker colours, *i.e.* red, pink and blue, are extremely rare and therefore very valuable. Diamonds can also be grey, yellow, brown, green and orange, and even black. Black diamonds have copper oxide present in them as an impurity. They are not used as gems, but they have important industrial uses.



A diamond sparkles due to the reflection of light from its several cut surfaces.

THE STRUCTURE OF DIAMOND

A diamond is a giant molecule. The number of valence electrons in a carbon atom is four. Each carbon atom is linked with four neighbouring carbon atoms, thus forming a rigid tetrahedral structure. It is this strong bonding that makes diamond the hardest naturally occurring substance. Since they have no free or mobile electrons, diamonds do not conduct electricity. The basic tetrahedral unit of a diamond crystal is repeated infinitely forming a three dimensional molecule. The shape of the crystal is octahedral.



Properties of diamond :

1. Pure diamond is transparent and colourless. Impurities impart colours to diamonds.
2. It is the hardest naturally occurring substance. Black diamonds are the hardest of all.
3. It is of a brittle nature.
4. It has a refractive index of 2·5, which is the highest value for a natural substance. Therefore, diamonds sparkle.
5. It has a density of 3·5 g/cm^3.
6. It is insoluble in any solvent.
7. It is a bad conductor of electricity.
8. Though it is stable to heat, prolonged heating can change a diamond into graphite.

The raw diamonds extracted from earth do not shine much. To make them shine, they are first cut and then polished. The country Belgium is the trade centre of costly diamonds in the world today.

Fig. 9.4 Diamond Jewellery

Uses of diamonds :

1. Because of its brilliant shine, pure diamond is used in jewellery as a gem.
2. Impure diamond (black diamond) is used:
   (i) for cutting and drilling rocks, glass or other diamonds.
   (ii) as tip heads in deep boring drills.
   (iii) as bearings in watches.
   (iv) as needles for long-playing record players.
   (v) for making radiation-proof windows for space satellites (so as to prevent the entry of harmful radiation).

Table 9.2 : Differences between diamond and graphite

Diamond	*Graphite*
1. Pure diamond is colourless and transparent.	1. Graphite is greyish black, opaque and shiny.
2. It is the hardest naturally occurring substance.	2. It is soft and greasy to touch.
3. It has high density, *i.e.* 3·5 g/cm^3.	3. It has a comparatively low density, *i.e.* 2·39 g/cm^3.
4. It is a bad conductor of electricity.	4. It is a good conductor of electricity.
5. It burns in air at 900°C to produce carbon dioxide.	5. It burns in air at 700°C to produce carbon dioxide.

Fullerenes

Fullerenes are the third crystalline form of carbon. Though they were discovered only recently, they have been found to exist in interstellar dust as well as in the geological formations of the earth.

In fullerenes, many carbon atoms are held together in a cage-like structure. The number of carbon atoms vary between 30-900. In the most common fullerene, called *buckminster fullerene* or *buckyball,* 60 carbon atoms are arranged in a spherical structure. Buckminster fullerene is denoted as C_{60}. It is named after **Richard Buckminster Fuller**, an American architect. There are larger and smaller fullerenes, made of 32, 50, 70 and 76 carbon atoms respectively. However, in fullerenes the cluster of carbon atoms exists as unlinked particles, unlike diamond and graphite where the carbon atoms are held together in a crystalline pattern.

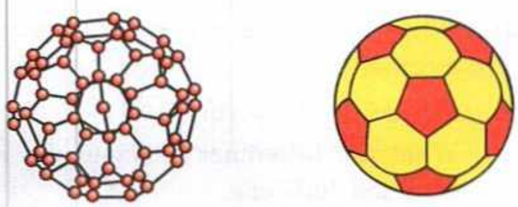


Fig. 9.5 The structure of buckminster fullerene. The structure has hexagons and pentagons, just as there are in a football.

Properties of Fullerenes :

1. The colour of fullerenes vary according to the number of carbon atoms present in them.
2. They are soluble in organic solvents.
3. They have specific gravity ranging from 1·8 to 2·1.
4. Chemically, fullerenes are more active than diamond and graphite. On heating upto ±1000°C, their cage like structure breaks.

Uses :

1. They act as insulators.
2. Some of the compounds of fullerene are used as **superconductors**.

EXERCISE – I

1. Fill in the blanks :

(*a*) .... is present in both living and non-living things.

(*b*) The tendency of an element to exist in two or more forms but in the same physical state is called ....

(*c*) .... and .... are the two major crystalline allotropes of carbon.

(*d*) .... is the hardest substance that occurs naturally.

(*e*) The name 'carbon' is derived from the Latin word ....

2. Choose the correct alternative :

(*a*) In combined state, carbon occurs as
(*i*) coal (*ii*) diamond
(*iii*) graphite (*iv*) petroleum

(*b*) A crystalline form of carbon is
(*i*) lamp black (*ii*) gas carbon
(*iii*) sugar (*iv*) fullerene

(*c*) Graphite is not found in
(*i*) Bihar (*ii*) Maharashtra
(*iii*) Odisha (*iv*) Rajasthan

(*d*) Diamond is used for
(*i*) making the electrodes of electric furnaces.

(ii) making crucibles for melting metals.
(iii) cutting and drilling rocks and glass.
(iv) making carbon brushes for electric motors.

(e) Carbon forms innumerable compounds because
(i) it has four electrons in its outermost shell.
(ii) it behaves as a metal as well as a non-metal.
(iii) carbon atoms can form long chains.
(iv) it combines with other elements to form covalent compounds.

3. Write *'true'* or *'false'* against the following statements :
(a) Carbon constitutes 0·03% of the earth's crust.
(b) Graphite is the purest form of carbon.
(c) Coloured diamonds are costlier than colourless and transparent diamonds.
(d) Graphite has layers of hexagonal carbon bondings.
(e) Diamond is insoluble in all solvents.

4. Define the following terms :
(a) Allotropy (b) Carat
(c) Crystal (d) Catenation

5. State the terms :
(a) Substances whose atoms or molecules are arranged in a definite pattern.
(b) Different forms of an element found in the same physical state.
(c) The property by which atoms of an element link together to form long chain or ring compounds.

6. Name the following :
(a) The hardest naturally occurring substance.
(b) A greyish black non-metal that is a good conductor of electricity.
(c) The third crystalline form of carbon.

7. Answer the following questions :
(a) Why is graphite a good conductor of electricity but not diamond ?
(b) Why is diamond very hard ?
(c) What are fullerenes ? Name the most common fullerene.
(d) What impurity is present in black diamond ?

8. Give *two* uses of :
(a) graphite (b) diamond

9. Write three differences between graphite and diamond.

B. Amorphous Forms of Carbon

INTRODUCTION

The word *amorphous* means *lacking in form or shape*. Accordingly, amorphous substances have no particular shape or structure. Their atoms or molecules are not arranged in any regular geometrical pattern. Some common amorphous substances are anhydrous copper sulphate, sulphur powder, talcum powder, wheat flour, etc.

The amorphous forms of carbon include coal, coke, charcoal, lamp black (soot) and gas carbon. They are derived from different sources. The amorphous forms of carbon are usually not pure (except sugar charcoal). Coal occurs in nature while the other forms are obtained as the black residue of compounds rich in carbon when they are heated in the absence of air.

> **Note :** The amorphous forms of carbon contain nitrogen, hydrogen, oxygen and sulphur as impurities.

COAL

Coal is one of the cheapest fossil fuels. It is a hard black solid. Rich deposits of coal are found in Russia, China, USA, Germany, South Africa and Australia. In India large deposits of coal are found in Jharkhand, West Bengal, Odisha and Madhya Pradesh.

Formation of coal : The formation of coal took millions of years. Coal was formed by the bacterial decomposition of ancient vegetable matter buried under successive layers of the earth. Under the action of high temperature and pressure, and in the absence of air, the decayed vegetable matter converted into coal through a series of steps as shown in Fig. 9.6. With each successive layer of coal formed, the amount of carbon present in the deposit increased and the level of impurities decreased. **This process is known as carbonization** and the most active episode of carbonization took place in the carboniferous era, *i.e.* 270 million years ago.

Types of coal

Carbonization of vegetable matter over millions of years resulted in four varieties of coal, which differ in their carbon content.

(i) **Peat :** Peat is the first stage in the formation of coal. Therefore, it is light brown in colour and contains only about 50 to 60% carbon. So peat is the most inferior form of coal.

(ii) **Lignite :** It is the second stage in the formation of coal. It contains more than 60% carbon. It is also brown in colour, but it is harder than peat.

(iii) **Bituminous coal :** Bituminous coal is the third stage in the formation of coal. There are high, medium and low varieties of bituminous coal, with carbon content being 90%, 80% and 70–75% respectively.

Bituminous coal is the most common variety of coal and is also known as household coal. It is black and hard. On heating it, both volatile and non-volatile materials are given out.

(iv) **Anthracite :** It is the purest variety of coal. It is the final stage in the formation of coal. Its carbon content varies between 92-98%. It is hard, dense and black. It is difficult to ignite, but once ignited it burns with a lot of heat and for a very long time. It is found at only a few places in the world.

Uses of coal

1. Coal is used as both a domestic and an industrial fuel, *i.e.* in homes, thermal power stations, steam engines and furnaces.

PLANTS $\xrightarrow[\text{pressure}]{\text{heat}}$ PEAT [light brown; 50– 60 % carbon] $\xrightarrow[\text{pressure}]{\text{heat}}$ LIGNITE [brown, 60–70% C] $\xrightarrow[\text{pressure}]{\text{heat}}$ BITUMINOUS [black, 80–90%C] $\xrightarrow[\text{pressure}]{\text{heat}}$ ANTHRACITE [black, 90–98%C]

Fig. 9.6 Formation of coal

2. It is used to prepare coke, coal gas and coal tar.
3. It is used in the manufacture of synthetic petrol.
4. Coal is also an important requirement for the manufacture of some fertilizers, drugs, synthetic textiles and perfumes.
5. It is used as a source of organic compounds such as benzene, naphthalene, aniline, *etc*.

Destructive distillation of coal

Destructive distillation of coal produces coke, coal tar, coal gas and ammoniacal solution.

$$\text{Coal} \xrightarrow[\text{distillation}]{\text{destructive}} \text{Coke} + \underset{\text{tar}}{\text{Coal}} + \underset{\text{gas}}{\text{Coal}} + \underset{\text{solution}}{\text{Ammoniacal}}$$

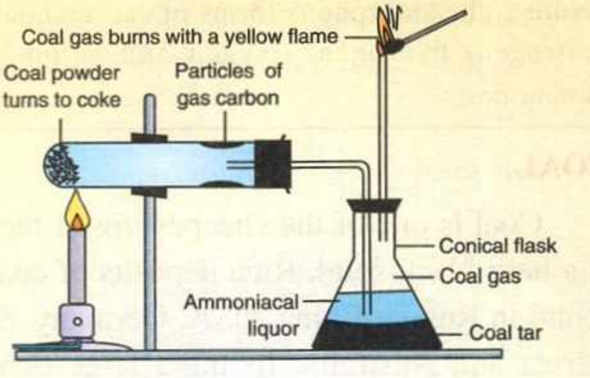
Important definitions

1. ***Carbonization :*** The process of the slow conversion of vegetable matter into carbon-rich substances is called carbonization.
2. ***Destructive distillation :*** When a substance is heated in the absence of air, the process is called destructive distillation. It results in the decomposition of the substance into its constituent substances.

Activity 3

Experiment : *To study the products obtained by destructive distillation of coal.* (To be shown by the teacher.)

Procedure : Place some powdered coal in a hard-glass test tube. Heat it strongly till it changes into coke that lies in the test tube as a grey porous residue.



Destructive distillation of coal.

Observation : Dark brownish black vapours are given out. Some carbon present in the vapours solidifies near the mouth of the test tube and is called *gas carbon*.

When the above vapours are passed through water contained in a conical flask, a large proportion of the vapours condenses to form two distinct layers. The lower layer consists of a black, thick liquid, which is called *coal tar.* The upper layer contains *ammoniacal liquor.* A colourless gas bubbles out of flask and it is commonly called *coal gas*. It burns with a yellowish flame when it is ignited.

Coke : Coke is an amorphous form of carbon, which is derived from coal. It is a black, porous, solid substance that burns without smoke. It contains about 98% carbon. It is a good reducing agent and a bad conductor of heat and electricity. Coke is prepared by the destructive distillation of coal.

Coke is of two types :

Hard coke : It is a light lustrous substance used in industrial furnaces.

Soft coke : It is a black and porous form of coke. It ignites easily and is used in household furnaces.

Uses of coke :

1. Coke is used as a smokeless industrial and domestic fuel.

2. It is used extensively as a reducing agent for the extraction of metals from their ores.
3. It is used in the manufacture of water gas ($CO + H_2$), producer gas ($CO + N_2$) and artificial graphite.
4. It is also used to prepare metallic carbides, *viz.* calcium carbide.

Coal tar : Coal tar is a foul smelling, dark brown liquid. On fractional distillation, it furnishes a number of useful organic compounds that are used for making dyes, drugs and explosives.

Coal gas : Purified coal gas is used as an industrial and a household fuel.

Ammoniacal solution : It is used in the preparation of artificial fertilizers.

CHARCOAL

When a solid organic substance is subjected to destructive distillation, a greyish black porous solid is produced, which is named as ***charcoal***. Charcoal is obtained from a variety of sources of plant and animal origin. Each type is named after the source from which it is obtained. The three main types of charcoal are ***wood charcoal, bone charcoal and sugar charcoal***.

Wood charcoal

Wood charcoal is prepared when wood is heated in a limited supply of air. It is a black, porous, brittle solid. Locally, wood charcoal is prepared by piling logs of wood one above the other with a gap in the centre of the pile. The pile is covered with wet clay to prevent the entry of air. A few holes are left at the bottom of the pile. The wood is set on fire. When the fire dies out, a greyish black, brittle, porous solid is left behind, which is wood charcoal.

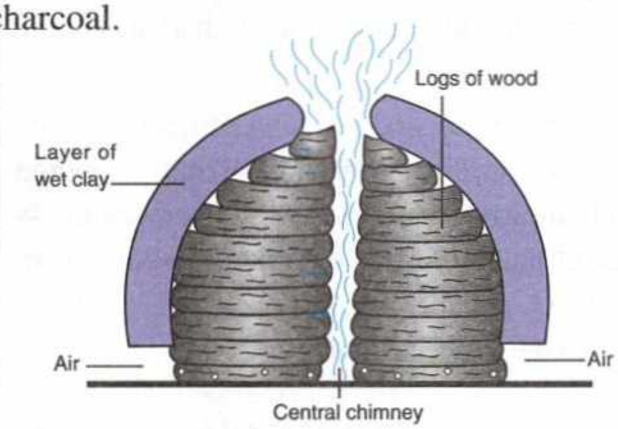
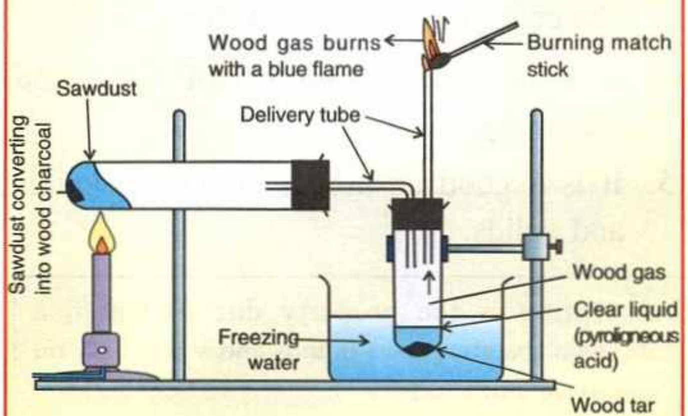


Fig. 9.7 Local preparation of wood charcoal.

Activity 4

To show how to prepare wood charcoal in the laboratory. (To be shown by the teacher).

Procedure : Take some sawdust in a hard glass test tube. Then fit a right-angled delivery tube to it with the help of a one-bored cork. Dip the other end of the delivery tube in another test tube held in a trough of cold water with the help of a clamp (see in the figure below). Fix



Preparation of wood charcoal

another delivery tube into the second test tube. Heat the sawdust strongly. Bring a burning match stick near the end of the second delivery tube. The gas that is liberated burns with a blue flame. Keep heating the sawdust till it chars to a black mass.

Now the second test tube contains a liquid. The lower layer of the liquid is dark coloured and is known as **wood tar**. The upper, colourless layer of liquid, is known as **pyroligneous acid**. The black residue left in the first test tube is **wood charcoal**.

$$\text{Wood} \xrightarrow[\text{distillation}]{\text{destructive}} \underset{\text{charcoal}}{\text{Wood}} + \underset{\text{tar}}{\text{Wood}} + \underset{\text{acid}}{\text{Pyroligneous}} + \text{Wood gas.}$$

Physical properties of wood charcoal :

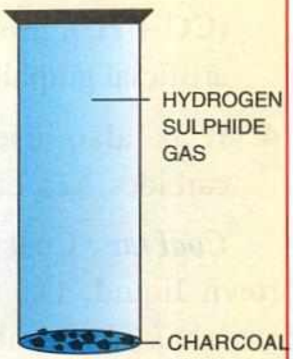
1. Wood charcoal is a soft, black, porous solid.
2. It is brittle and tasteless.
3. Though it is heavier than water, it floats on it, since it is porous and has the capacity to hold air in its pores. When the air is removed from the pores of the wood charcoal by boiling water, the charcoal gradually settles down.
4. It is a bad conductor of heat and electricity.
5. It is a good adsorbent of gases, liquids and solids.

Adsorption is the property due to which a substance absorbs gases, liquids and solids only on its surface.

Activity 5

To show that wood charcoal is a good adsorbent.

Procedure : Take a gas jar and fill it with hydrogen sulphide gas. It smells like rotten eggs. Put some charcoal pieces in the jar and close the lid. Now shake the jar and leave it still for a few minutes. Then remove the lid. You will notice that there is no smell in the jar. This proves that charcoal is a good adsorbent of gases.

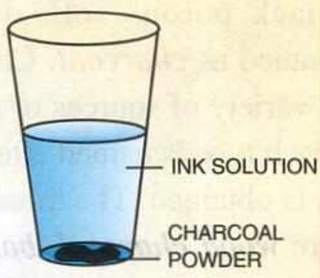


A jar containing a gas with a foul smell

Activity 6

To show that wood charcoal adsorbs colouring matter.

Procedure : Take a glass tumbler, fill it half with water and pour in a few (2-3) drops of ink. Colourless water turns blue, meaning that an ink solution has been formed. Now add a small quantity of charcoal powder to the ink solution and stir well.



Wood charcoal adsorbs colouring matter from its solution.

After a while you will observe that the blue colour of the ink solution fades away. This proves that wood charcoal adsorbs ink from the ink solution, *i.e.* coloured matter from its solution.

Activated Charcoal : The adsorption capacity of wood charcoal is increased by heating it upto 900°C by steam, because it opens up the pores and increases its capacity to hold more substances. This charcoal is called **activated charcoal.**

Uses of wood charcoal :

1. Wood charcoal is used as a fuel because it burns at low temperatures and produces no smoke.
2. It is used extensively as a reducing agent in the extraction of metals from their respective metallic oxides.
3. It is used to prepare water gas which is an important industrial fuel.
4. It is an important constituent of gun powder.
5. Due to its high adsorbing capacity, wood charcoal is used :
   (a) in military and industrial gas masks to adsorb harmful gases (poisonous and foul smelling).
   (b) in the form of tablets by persons suffering from indigestion and gastric problems. The tablets adsorb the stomach gases and thus relieve gas pressure.
   (c) to decolourise sugar syrup and refine fats and oils.
   (d) in making water purifier filters and sieves.

Sugar charcoal

Sugar charcoal is the purest form of amorphous carbon. It is prepared by heating cane sugar or glucose in the absence of air.

$$\text{cane sugar} \xrightarrow{\text{heat}} \text{sugar charcoal} + \text{water}$$

$$C_{12}H_{22}O_{11} \xrightarrow{\text{heat}} 12C + 11H_2O$$

Sugar charcoal can also be prepared by the dehydration of cane sugar or glucose in the presence of concentrated sulphuric acid. The acid removes the water, leaving behind carbon.

Uses : 1. Sugar charcoal is used mostly as a reducing agent to extract metals from their respective oxides. 2. It is also used to decolourise coloured solutions. 3. It is used to prepare artificial diamonds.

$$\text{Sugar charcoal} \xrightarrow{3000° - 3500°\text{C}} \text{Artificial diamond}$$

Activity 7

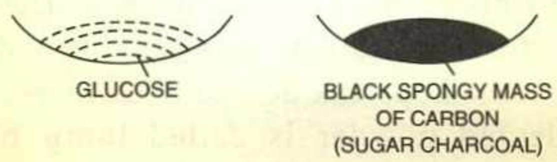
To show the formation of charcoal. (To be demonstrated by the teacher.)

Take some glucose, a white crystalline solid in a watch glass. Add a few drops of concentrated sulphuric acid to it.

What do you observe?

The white solid has turned into a black spongy mass of carbon.

The black mass is nothing but sugar charcoal.



Glucose turned into sugar charcoal

Bone charcoal

Destructive distillation of bones produces bone charcoal along with bone oil and an organic compound pyridine. Bone charcoal contains mainly calcium phosphate, *i.e.* its carbon content is rather limited (10-12%). The carbon content of bone charcoal is separated by treating the latter with hydrochloric acid, which dissolves the calcium phosphate. Carbon is then filtered out of the solution and in this form it is called **bone black** or **ivory black**.

Uses of bone charcoal :

1. Bone charcoal is extensively used to decolourise cane sugar in the process of manufacturing sugar.
2. It is also used in the manufacture of a large number of phosphorus compounds.
3. Ivory black is used as a black pigment in artistic paintings (for making charcoal pencils) because it is the deepest black.
4. It is used to filter aquarium water.
5. It is used to remove excess of fluoride from water which causes tooth decay.

LAMP BLACK (SOOT)

Lamp black is one of the amorphous forms of carbon. It contains 98-99% carbon. It is prepared by heating carbon-rich substances like turpentine oil or kerosene oil in a limited supply of air. The oil burns with a sooty flame that contains large amounts of free carbon. The black smoke is collected in the form of black powder over damp blankets kept inside the chambers. The collected powder is called **lamp black** or **soot**. It is a light, black powder that is used in our country to prepare an eye liner, commonly known as **kajal**.

Uses :

1. Lamp black is used in making black shoe polish, carbon paper, printing ink, black paint, *etc.*
2. It is also used in the manufacture of tyres and gun powder.

Activity 8

Preparation of kajal (lamp black) from vegetable oil.

Take some mustard oil or ghee in an earthen lamp or a bowl. Put a cotton wick in it. Light the wick and hold a metal plate over the flame.

Black powder gets deposited on the surface of the plate. This is *kajal*, a commonly used cosmetic for eye lids.

Preparing kajal at home.

GAS CARBON

Gas carbon is prepared by destructive distillation of coal or when petroleum products are heated at high temperatures in a closed container.

Carbon particles are deposited on the walls of the container which is **gas carbon**.

It is a grey solid. It is also a good conductor of heat and electricity (similar to graphite).

Uses : 1. It is used for making electrodes of dry cells.

2. It is used to make the lining of electrolytic cells.

3. It is used to make carbon rods for arc lamps.

Table 9.3 : Showing uses of allotropes of carbon

Name of allotrope	Uses
1. **Diamond**	• In jewellery as a gem due to its brilliant shine, for cutting and drilling rocks, glass, etc., as a bearing in watches, for making radiation-proof windows for space satellites.
2. **Graphite**	• As a lubricant, for making pencil leads, for making electrodes of electric furnaces, for making crucibles to keep molten metals, in nuclear reactors as moderators to slow down the speed of neutrons, for making artificial diamonds, etc.
3. **Fullerene**	• As insulators and for making superconductors.
4. **Coal**	• As domestic and industrial fuel, to prepare coke, coal gas and coal tar, in the manufacture of synthetic petrol, drugs, textile, perfumes and some fertilizers, as a source of organic compounds.
5. **Coke**	• As a smokeless fuel, as a reducing agent in the extraction of metals, in the manufacture of water gas, producer gas, artificial graphite and metallic carbides.
6. **Wood charcoal**	• Wood charcoal is used as a fuel, as a reducing agent, in preparing water gas, gun powder, to decolourise sugar syrup and in refining fats and oils, in gas masks to adsorb harmful and poisonous gases, in the form of tablets to remove gastric problems.
7. **Sugar charcoal**	• As a reducing agent to extract metals, to decolourise coloured solutions and for making artificial diamonds.
8. **Bone charcoal**	• To decolourise cane sugar, in the manufacture of sugar, in making black paints, as filters to remove impurities from water and in making phosphorous compounds.
9. **Lamp black (soot)**	• In making black shoe polishes, carbon paper, black paints, printing inks, etc.
10. **Gas carbon**	• In making electrodes of dry cells and carbon rods in arc lamps.

EXERCISE – II

1. Fill in the blanks :
   (a) .... is formed when charcoal is burnt in a limited supply of air.
   (b) Coal is a .... form of carbon.
   (c) .... is the most inferior form of coal.
   (d) Wood charcoal is a .... conductor of heat and electricity.
   (e) .... is used in making black shoe polish.

2. Choose the correct alternative :
   (a) Anthracite is
   (i) an inferior type of coal
   (ii) a superior type of coal
   (iii) a cheapest form of coal
   (iv) none of the above
   (b) Destructive distillation of coal yields
   (i) coal tar
   (ii) coal gas
   (iii) coke
   (iv) all of the above
   (c) Lamp black is
   (i) an amorphous form of carbon
   (ii) a crystalline form of carbon
   (iii) a pure form of carbon
   (iv) a cluster of carbon atoms
   (d) The process by which decayed plants slowly convert into coal is called
   (i) petrification
   (ii) carbonization
   (iii) carbonification
   (iv) fermentation
   (e) The purest form of the amorphous carbon is
   (i) wood charcoal
   (ii) sugar charcoal
   (iii) bone charcoal
   (iv) lamp black

3. Write *'true'* or *'false'* against the following statements :
   (a) Charcoal is a good adsorbent.
   (b) Coke is obtained by destructive distillation of sugar.
   (c) Activated charcoal is a good conductor of electricity.
   (d) Wood charcoal is an important constituent of gun powder.
   (e) Coal gas is used in the preparation of artificial fertilizers.

4. Define the following :
   (a) Carbonization (b) Adsorption
   (c) Bone black

5. Name the following :
   (a) Substances whose atoms or molecules are not arranged in a geometrical pattern.
   (b) The best variety of coal.
   (c) The purest form of amorphous carbon.
   (d) An amorphous form of carbon that contains about 98% carbon.
   (e) Mixture of carbon monoxide and hydrogen.

6. Answer the following questions :
   (a) What is destructive distillation ? What are the products formed due to the destructive distillation of coal ?
   (b) Why is wood charcoal used in water filters and gas masks ?
   (c) How is wood charcoal made locally ? What other substances are formed in the process ?
   (d) How many carbon atoms are there in Buckminster fullerenes ?

7. (a) Describe the formation of coal.
   (b) Name four types of coal with the percentage of carbon present in each.
8. Name the products formed when :
   (a) wood is burnt in the absence of air.
   (b) bone is heated in the absence of air.
   (c) diamond is burnt in air at 900°C.
   (d) graphite is subjected to high pressure and 3000°C temperature.
9. Give *two* uses for each of the following :
   (a) coal (b) coke
   (c) wood charcoal (d) sugar charcoal
   (e) bone charcoal (f) lamp black

C. Carbon dioxide

Molecular formula : CO_2

Relative molecular mass = 44 amu

One molecule of carbon dioxide contains one atom of carbon and two atoms of oxygen.

DISCOVERY

Carbon dioxide gas is one of the most important constituents of air. It was discovered by *Van Helmont* in 1630 by burning charcoal in air.

In 1775, *Joseph Black* obtained carbon dioxide by the action of dilute acids on metal carbonates. Later, *Antoine Lavoisier* studied the gas and named it *acid carbonique.* He found that the gas was an oxide of carbon, which when dissolved in water produces an acidic solution. Later on, it was named carbon dioxide.

OCCURRENCE

In nature, carbon dioxide occurs in free state as well as in its combined state.

(i) In free state :

- About 0·03% – 0·04% (by volume) of carbon dioxide is present in air.
- It is also present in the rocks of volcanic regions.
- In natural water, carbon dioxide is present in a dissolved state as it is fairly soluble in water.

(ii) In combined state :

It occurs as metallic carbonates and bicarbonates in the earth's crust and also in sea shells. The chief minerals containing CO_2 are dolomite ($MgCO_3.CaCO_3$), limestone ($CaCO_3$), *etc.*

Note : All life-forms are based on carbon-containing molecules like proteins, carbohydrates, fats, nucleic acid and vitamins. The exoskeletons and endoskeletons of various animals are also formed from carbonate salts.

PREPARATION OF CARBON DIOXIDE

Carbon dioxide can be prepared by any of the following methods :

(i) By burning carbon or its compounds :

The burning of carbon (charcoal, coke) *or carbon compounds produces carbon dioxide.*

(a) carbon + oxygen $\xrightarrow{\text{burning}}$ carbon dioxide + heat

$$C + O_2 \longrightarrow CO_2 + \text{heat}$$

(b) methane + oxygen $\xrightarrow{\text{burning}}$ carbon dioxide + water

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

(ii) By heating metal carbonates and bicarbonates :

Metallic carbonates and bicarbonates decompose on strong heating to produce carbon dioxide.

(a) Copper carbonate $\xrightarrow{\text{heating}}$ copper (II) oxide + carbon dioxide

$$CuCO_3 \xrightarrow{\text{heating}} CuO + CO_2$$

(b) Sodium bicarbonate $\xrightarrow{\text{heating}}$ sodium carbonate + water + carbon dioxide

$$2NaHCO_3 \xrightarrow{\text{heating}} Na_2CO_3 + H_2O + CO_2$$

(iii) By the action of dilute acids on Metallic carbonates and bicarbonates :

Metal carbonates and bicarbonates react with dilute acids to produce carbon dioxide.

(a) Calcium carbonate + dilute hydrochloric acid → calcium chloride + water + carbon dioxide

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2(g)$$

(b) Sodium carbonate + dilute sulphuric acid → sodium sulphate + water + carbon dioxide

$$Na_2CO_3 + H_2SO_4 \rightarrow Na_2SO_4 + H_2O + CO_2(g)$$

(c) Sodium carbonate + dilute hydrochloric acid → sodium chloride + water + carbon dioxide

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$$

(d) Sodium bicarbonate + dilute hydrochloric acid → sodium chloride + water + carbon dioxide

$$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2(g)$$

(e) Baking soda* + vinegar (acetic acid) → sodium acetate + water + carbon dioxide

$$NaHCO_3 + CH_3COOH \rightarrow CH_3COONa + H_2O + CO_2$$

* The chemical name of baking soda is sodium bicarbonate ($NaHCO_3$). It is used to make food items light and spongy.

LABORATORY PREPARATION OF CARBON DIOXIDE

Carbon dioxide is prepared in the laboratory by the action of dilute acid on metal carbonate.

Chemicals required :

(i) Marble chips ($CaCO_3$)

(ii) Dilute hydrochloric acid

Chemical equation for the reaction

Calcium carbonate + dilute hydrochloric acid → calcium chloride + water + carbon dioxide

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2(g)$$

Procedure : A Woulfe's bottle* is taken and some marble chips are placed in it. Through the mouth of the bottle, a thistle funnel is introduced with the help of a rubber stopper. Through the second mouth of the Woulfe's bottle, a delivery tube is fitted. The other end of the tube is put in a gas jar.

Now, dilute hydrochloric acid is poured into the bottle through the thistle funnel so that the marble chips are completely immersed in the acid and the stem of the thistle funnel dips in the acid.

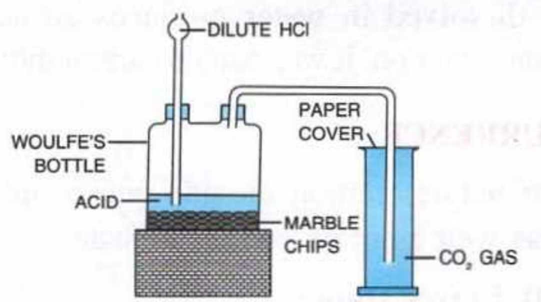


Fig. 9.8 : Laboratory preparation of carbon dioxide.

* Woulfe's bottle is a double mouth glass apparatus used in the chemistry laboratory to prepare gases.

A vigorous chemical reaction takes place and carbon dioxide gas is released with brisk effervescence. First, a few bubbles of carbon dioxide are allowed to escape, as they might contain air and acid as impurities. Then, the gas is allowed to pass through the delivery tube into a gas jar, where it is collected by *upward displacement of air*, since CO_2 is heavier than air.

Why is calcium carbonate preferred ?

Calcium carbonate is preferred to other metallic carbonates to prepare carbon dioxide because it is cheap and easily available.

Why is sulphuric acid not used ?

Dilute sulphuric acid also reacts with calcium carbonate, just like hydrochloric acid does. But it is not used because the calcium sulphate which is formed during the reaction is insoluble in water. It covers the marble chips and stops the reaction.

$$\underset{\text{(calcium carbonate)}}{CaCO_3} + \underset{\text{(dilute sulphuric acid)}}{H_2SO_4} \rightarrow \underset{\text{(calcium sulphate)}}{CaSO_4} + \underset{\text{(water)}}{H_2O} + \underset{\text{(carbon dioxide)}}{CO_2}$$

Why is carbon dioxide not collected over water?

Because carbon dioxide is highly soluble in water.

PROPERTIES OF CARBON DIOXIDE

Physical properties

1. **Nature :** Carbon dioxide is a colourless, odourless gas, with a faint acidic (sour) taste.
2. **Density :** Under ordinary conditions, carbon dioxide is 1·5 times heavier than air.
3. **Solubility :** It is fairly soluble in water. Solubility increases with an increase in pressure.
4. **Liquefaction and solidification into dry ice :** Carbon dioxide can be liquefied under pressure. When liquid CO_2 is cooled to –78°C, under normal pressure, it changes into a snow-white solid, called *dry ice*. Dry ice has nothing in common with ice formed from water, except that it is a solid with snow-white colour.

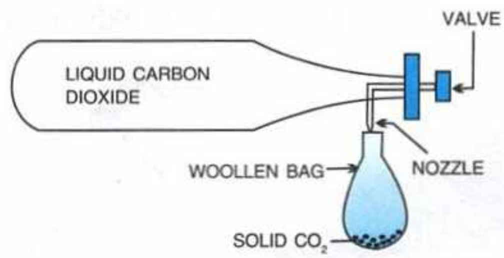


Fig. 9.9 : To show how CO_2 can be solidified

Dry ice, directly changes into gaseous carbon dioxide, from its solid state, *i.e.*, without passing through its liquid state. This is because dry ice is a sublimable substance. It has a cooling effect.

5. **Physiological action :** Carbon dioxide is a non-poisonous gas, but an excess of this gas can cause suffocation. A person may die in an atmosphere of carbon dioxide due to lack of oxygen.

Do You Know ?

Aquatic plants use the carbon dioxide dissolved in water to prepare their food during photosynthesis.

Chemical properties

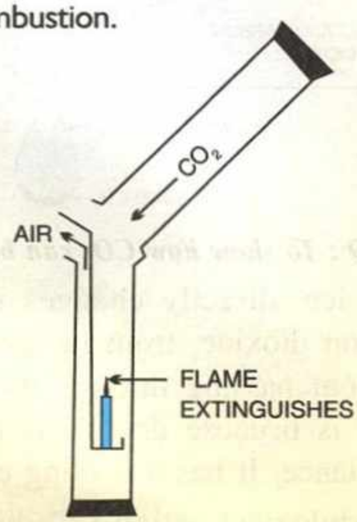
1. **Combustibility :** Carbon dioxide is neither combustible nor a supporter of combustion.

Activity 9

To show that carbon dioxide is heavier than air and a non-supporter of combustion.

Take a jar containing air. Introduce a burning candle in it. Tilt a jar containing carbon dioxide over the jar with burning candle. You will observe that the candle flame goes off. This is because carbon dioxide being heavier, moves downwards, spreads over the flame and pushes the air up.

This also proves that carbon dioxide does not support combustion.



However, some metals like sodium, potassium and magnesium continue to burn in presence of carbon dioxide (see reaction 5).

2. **Action on litmus paper :** Carbon dioxide turns moist blue litmus red. This shows that the gas is acidic in nature.

3. **Reaction with water :** Carbon dioxide dissolves in water to give carbonic acid (a weak acid).

$$\underset{\text{(carbon dioxide)}}{CO_2} + \underset{\text{(water)}}{H_2O} \longrightarrow \underset{\text{(carbonic acid)}}{H_2CO_3}$$

4. **Reaction with alkalis :** Carbon dioxide reacts with alkalis to produce salt and water.

$$\underset{\text{(sodium hydroxide)}}{2NaOH} + CO_2 \longrightarrow \underset{\text{(sodium carbonate)}}{Na_2CO_3} + H_2O$$

$$\underset{\text{(potassium hydroxide)}}{2KOH} + CO_2 \longrightarrow \underset{\text{(potassium carbonate)}}{K_2CO_3} + H_2O$$

When excess of carbon dioxide is passed through alkalis, a soluble bicarbonate is obtained.

$$\underset{\text{(sodium carbonate)}}{Na_2CO_3} + CO_2 + H_2O \longrightarrow \underset{\text{(sodium bicarbonate)}}{2NaHCO_3}$$

$$\underset{\text{(potassium carbonate)}}{K_2CO_3} + CO_2 + H_2O \longrightarrow \underset{\text{(potassium bicarbonate)}}{2KHCO_3}$$

Potassium hydroxide (caustic potash) or sodium hydroxide (caustic soda) are used to remove carbon dioxide from a mixture of gases because small amount of these alkalis can absorb large amount of carbon dioxide. However, potassium hydroxide is preferred as it has better absorbing capacity.

Carbon dioxide when passed through lime water [$Ca(OH)_2$] turns it milky. This is due to the formation of insoluble calcium carbonate. But when an excess of the gas is passed through the solution, the milkiness disappears. This is due to the formation of a soluble bicarbonate.

$$\underset{\text{(lime water)}}{Ca(OH)_2} + CO_2 \longrightarrow \underset{\text{(calcium carbonate, insoluble)}}{CaCO_3} + H_2O$$

$$CaCO_3 + H_2O + CO_2 \longrightarrow \underset{\text{(calcium bicarbonate, soluble)}}{Ca(HCO_3)_2}$$

This reaction is used as a chemical test for carbon dioxide.

5. **Reaction with metals :** Magnesium metal burns in presence of carbon dioxide to give magnesium oxide and carbon (black).

$$\text{Magnesium} + \text{Carbon dioxide} \xrightarrow{\text{heat}} \text{Magnesium oxide} + \text{Carbon}$$

$$2Mg + CO_2 \xrightarrow{\text{heat}} 2MgO + C$$

6. **Reaction with metallic oxides (basic oxides) :** Carbon dioxide reacts with metallic oxides to produce metal carbonates.

(a) $$\text{Sodium oxide} + \text{Carbon dioxide} \longrightarrow \text{Sodium carbonate}$$

$$Na_2O + CO_2 \longrightarrow Na_2CO_3$$

(b) $$\text{Magnesium oxide} + \text{Carbon dioxide} \longrightarrow \text{Magnesium carbonate}$$

$$MgO + CO_2 \longrightarrow MgCO_3$$

7. **Reaction with non-metals :** Carbon dioxide reacts with red hot coke to produce carbon monoxide, a highly poisonous gas.

$$\text{Carbon dioxide} + \text{Coke} \longrightarrow \text{Carbon monoxide}$$

$$CO_2 + C \longrightarrow 2CO$$

TESTS FOR CARBON DIOXIDE

1. Carbon dioxide turns moist blue litmus paper into pinkish red.
2. If a burning candle or a smouldering match stick is introduced in a jar containing carbon dioxide, it extinguishes. This proves that the gas is a non-supporter of combustion.
3. When carbon dioxide is passed through lime water, the latter turns milky and when excess of gas is passed through the solution, milkiness disappears. This is an excellent test for carbon dioxide.

USES OF CARBON DIOXIDE

1. **In aerated drinks :** Carbon dioxide is used in the manufacture of aerated drinks. The gas is dissolved in water under pressure to give **soda water**.

   When the bottle is opened, the pressure is released and the bottled gas escapes with a brisk effervescence that adds a *fizz* to the drink. Common soft drinks are formed by dissolving carbon dioxide in a sugar solution.

2. **Refrigeration and preservation of food stuffs : Solid carbon dioxide (dry ice)** is used as a coolant and a refrigerant in ships for preserving food articles like fruits, vegetables, meat, etc. which otherwise perish easily. Also, if grains are kept in an atmosphere of carbon dioxide, they remain in good condition for a long period, without allowing insect attacks.
3. **In hospitals :** A mixture of 5% carbon dioxide and 95% oxygen, called *carbogen,* is used for artificial respiration. It is given to patients suffering from gas poisoning, pneumonia, drowning, etc.
4. **In the manufacture of fertilizers (urea) :** Urea is an important nitrogenous fertilizer. It is prepared when carbon dioxide and ammonia are heated at 150°C under very high pressure.
5. **In the preparation of chemicals :** Carbon dioxide is used to manufacture chemicals like washing soda (sodium carbonate) and baking soda (sodium bicarbonate).
6. **In the baking industry (cooking processes) :** Carbon dioxide is used to make the dough rise and become light.

Baking powder is a mixture of baking soda ($NaHCO_3$), potassium hydrogen tartarate and starch. When it is added to the dough, the ingredients of baking soda react to release carbon dioxide. As the gas rises through the dough, spaces are formed, thus making the dough porous and the cakes and bread light, soft and spongy.

7. **In photosynthesis :** Carbon dioxide is used in photosynthesis, a natural process to prepare food by green plants. The food prepared is also used by animals.

   Photosynthesis helps in maintaining the balance of carbon dioxide and oxygen in the atmosphere.

8. **In extinguishing a fire :** When a substance burns with a flame, the heat, light and often the smoke so produced, constitute fire.

   For a fire to occur, presence of a combustible substance and a supporter of combustion are required. If these two conditions are not satisfied, a fire will not be produced. This is the principle used for fire fighting purposes, on which fire extinguishers are constructed.

   Carbon dioxide helps to extinguish fires because :

   (i) it neither burns nor does it help in burning.

   (ii) being heavier than air, it insulates the burning substance by cutting off the supply of oxygen, which is a supporter of combustion.

FIRE-EXTINGUISHERS

Fire extinguishers are devices in which carbon dioxide is produced in different forms to be used as an extinguishing agent.

Some common types of fire extinguishers are :

(i) soda-acid fire extinguishers

(ii) foam-based fire extinguishers

(iii) liquid carbon dioxide fire extinguishers

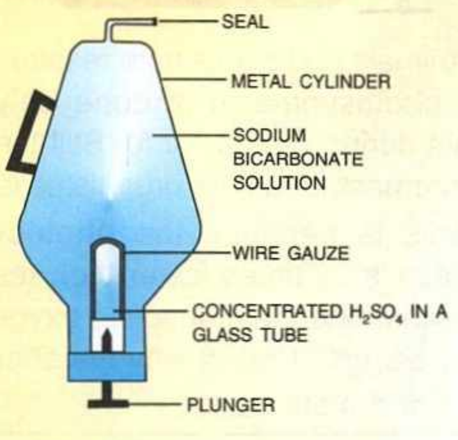
(i) Soda-acid fire extinguishers : It contains sodium bicarbonate and sulphuric acid in separate chambers. The extinguisher consists of a metallic cylinder filled up to two-thirds with a saturated solution of sodium bicarbonate. A sealed glass bottle containing concentrated sulphuric acid is kept inside the cylinder. When the apparatus has to be used, the cylinder is inverted and made to hit the floor. As a result, the glass bottle with the acid breaks and the two chemicals come in contact to react and produce carbon dioxide gas. The gas comes out through the nozzle with a great force. It spreads over the fire and cuts off the supply of oxygen to it and hence the fire is extinguished.

$$\underset{\text{(Sodium bicarbonate)}}{2NaHCO_3} + \underset{\text{(sulphuric acid)}}{H_2SO_4} \rightarrow \underset{\text{(sodium sulphate)}}{Na_2SO_4} + 2H_2O + 2CO_2(g)$$

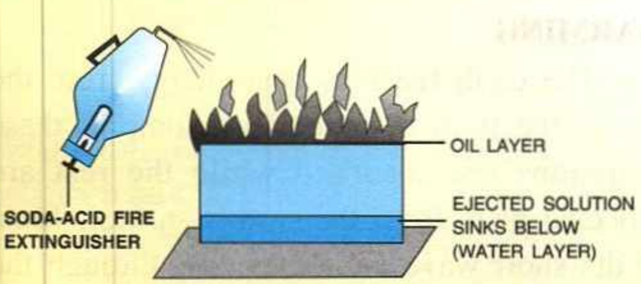
Soda acid fire extinguisher cannot be used in case of an oil fire because the solution ejected by this extinguisher is heavier than oil, so it sinks below the oil and does not work.

(ii) Foam-type fire extinguishers : It contains aluminium sulphate and sodium bicarbonate in two separate chambers. When needed, the two chemicals are made to mix with each other and react to produce carbon dioxide and aluminium hydroxide which come out in the form of foam through the nozzle. It is used to extinguish oil-fed fires because the

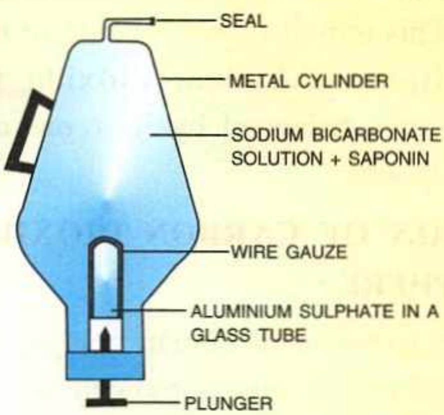
SODA-ACID FIRE EXTINGUISHER	FOAM-TYPE FIRE EXTINGUISHER



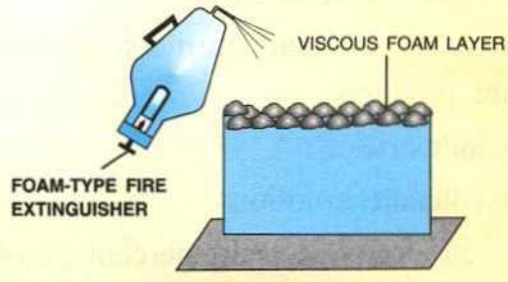
SODA-ACID-TYPE FIRE EXTINGUISHER



A SODA-ACID TYPE FIRE EXTINGUISHER CANNOT FIGHT AN OIL FIRE



FOAM-TYPE FIRE EXTINGUISHER



A FOAM-TYPE FIRE EXTINGUISHER CAN FIGHT AN OIL FIRE

foam covers the oil as well as cuts off the air supply to the fire.

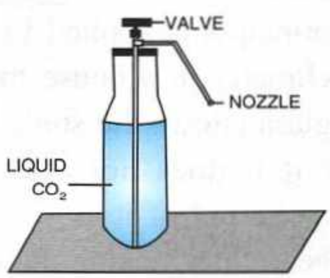
$$6NaHCO_3 + Al_2(SO_4)_3 \rightarrow 3Na_2SO_4 + 2Al(OH)_3\downarrow + 6CO_2\uparrow$$

(Sodium bicarbonate)	(Aluminium sulphate)	(Sodium sulphate)	(Aluminium hydroxide)	(Carbon dioxide)

Note : • A chemical *saponin* is used to produce foam.

• Sulphuric acid is replaced by aluminium sulphate.

(iii) Liquid carbon dioxide fire extinguishers : It is a modern and improvised type of fire extinguisher, in which liquid carbon dioxide is stored in a steel cylinder under pressure.



LIQUID CARBON DIOXIDE FIRE EXTINGUISHER

On opening the valve of the cylinder, pressure falls and liquid carbon dioxide solidifies into white snow (dry ice). It can be used to put out both oil-fed fires and electrical fires.

Why the soda-acid and the foam-type of extinguishers cannot be used for fighting electrical fires ?

In both of these fire extinguishers, the solutions are prepared in water, which conducts electricity. As a result, it might generate an electric shock leading to short-circuit and another fire.

CARBON DIOXIDE CYCLE

The process by which the amount of carbon dioxide is maintained in the atmosphere is called **carbon dioxide cycle.**

In the atmosphere, the percentage of carbon dioxide is about 0·03% - 0·04% by volume. This remains almost constant because the addition of carbon dioxide to the atmosphere is balanced by its removal from the atmosphere.

ADDITION OF CARBON DIOXIDE TO ATMOSPHERE :

- By respiration of human beings, animals and plants in which oxygen is used up and carbon dioxide is released.
- By combustion of fuels.
- By decay of dead animals, plants and plant products.
- By industries.
- By volcanic eruptions.
- By sea water when the percentage of gas increases in water.

REMOVAL OF CARBON DIOXIDE FROM ATMOSPHERE :

- By photosynthesis.
- By dissolution of carbon dioxide in water when the percentage of gas increases in the atmosphere.

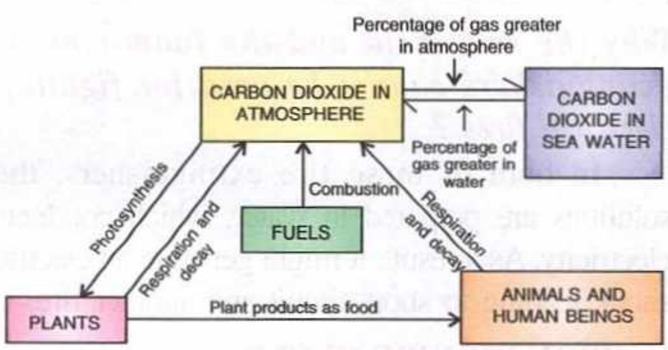


Fig. 9.10 : Carbon dioxide Cycle

Do You Know ?

Animals and plants both respire 24 hours and photosynthesis is done only by the plants during the day time. Still the oxygen requirement for all the organisms is fulfilled.

This is because the photosynthesis process is 28 times faster than respiration. It means, one plant supplies oxygen for 27 living beings. That is why we should grow more and more trees.

GREENHOUSE EFFECT AND GLOBAL WARMING

The earth receives heat energy from the sun in the form of radiations. Some of these radiations are absorbed while the rest are reflected back from the earth's surface. Most of the short wave radiations pass through the air but most of the long wave reflected radiations are trapped by the carbon dioxide gas present in air, making the earth hot. This is a blessing, for it keeps the earth warm.

The trapping of the earth's radiated energy by carbon dioxide present in the air, so as to keep the earth warm, is called the 'greenhouse effect'.

This effect is known as the "greenhouse effect" because, in colder regions, this principle is applied to grow plants of warmer climates in a house made of glass walls. The glass allows the sun's heat to enter the house, but it does not allow this trapped heat to move out of it as glass is a bad conductor of heat, thus making the house sufficiently warm inside to help grow tropical plants. The glass house is popularly called a greenhouse, and therefore we get the name "*greenhouse effect.*"

Why does a closed car parked in the sun become unbearably hot inside ?

However, due to air pollution, the total amount of carbon dioxide in air has increased. Besides that, there are other greenhouse gases in air due to pollution, such as nitrous oxide, ozone, chlorofluorocarbons, methane, etc. All these gases have created a thermal blanket around the earth trapping more heat and increasing the greenhouse effect. This causes further warming of the earth known as **global warming**.

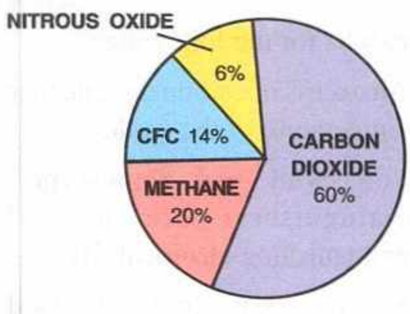


Fig. 9.11 : The percentage contribution of different gases to global warming

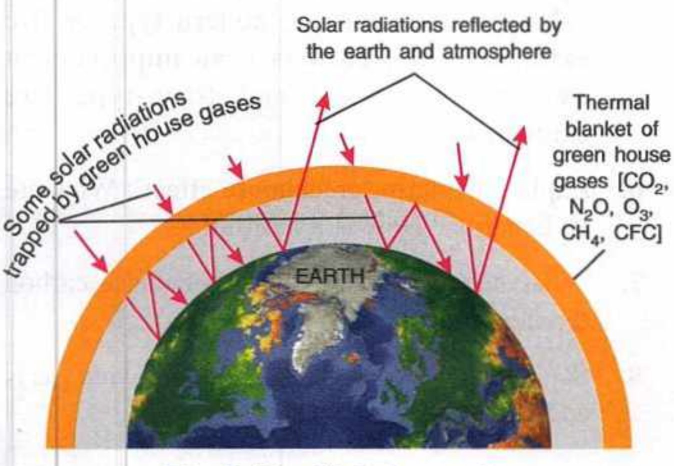


Fig. 9.12 : Global warming

CAUSE OF INCREASED PERCENTAGE OF CARBON DIOXIDE IN THE ATMOSPHERE

- Deforestation or cutting down of vegetation. (As we know that plants help in the removal of carbon dioxide from air by utilising it during photosynthesis). Thus, human activities contribute to the accumulation of CO_2 in the atmosphere.
- Air pollution due to industries, burning of fuels, use of chemical weapons, *etc.*

EFFECTS OF GLOBAL WARMING

Global warming has serious consequences.

- According to scientists, the average temperature of the earth has risen by 0·5°C in the past 100 years.
- Ice in polar regions may melt causing floods in coastal regions and islands. The Gangotri glacier in the Himalayas has already started melting.
- A disturbance in ecological balance may be caused. It could result in wide ranging effects on rainfall patterns, agriculture, forests, etc.
- Extinction of many species of plants and animals may be caused.
- This may result in a change in the pattern of crop cultivation (crop cycle).

STEPS TO BALANCE CARBON DIOXIDE IN THE ATMOSPHERE

Since global warming will cause an unbalanced ecological system, serious efforts should be made to balance the percentage of carbon dioxide in the atmosphere. Some of them are :

- Growing more trees and plants.
- Reducing the consumption of fossil fuels.
- Using smokeless sources of energy, like solar energy, biogas, *etc*.
- Using filters in the chimneys of factories and power houses.

Do You Know ?

Many countries have reached an agreement to reduce the emission of greenhouse gases. The Kyoto Protocol is one such agreement.

EXERCISE – III

1. (*a*) Name the chemicals required for the preparation of carbon dioxide in the laboratory.

(*b*) How will you collect the gas ?

(*c*) Write the balanced chemical equation for the above reaction.

(*d*) Draw a labelled diagram for the preparation of CO_2 in the laboratory.

(*e*) Why is sulphuric acid not used for the preparation of carbon dioxide in the laboratory ?

2. Write the balanced chemical equations for the preparation of carbon dioxide by :

(*a*) heating calcium carbonate.

(*b*) the action of acetic acid on sodium bicarbonate.

(*c*) the action of dilute sulphuric acid on sodium bicarbonate.

(*d*) the action of dilute hydrochloric acid on sodium carbonate.

3. What happens when :

(*a*) a lit splinter is introduced into a jar containing carbon dioxide ?

(*b*) moist blue litmus paper is placed in a jar containing carbon dioxide ?

(*c*) carbon dioxide is passed through lime water first in small amounts and then in excess ?

(*d*) a baking mixture containing baking powder is heated ?

(*e*) a soda water bottle is opened ?

4. Give reasons for the following :

(*a*) An excess of carbon dioxide increases the temperature of the earth.

(*b*) Soda-acid and foam-type of fire extinguishers are not used for extinguishing electrical fires.

(*c*) Solid carbon dioxide is used for the refrigeration of food.

5. What is a fire extinguisher ? What is the substance used in the modern type of fire extinguishers ? How is it an improvement over the soda-acid and foam-type fire extinguishers ?

6. Explain the term 'greenhouse effect'. What are its benefits ? How is it harmful ?

7. What steps should be taken to balance carbon dioxide in the atmosphere ?

8. State three ways by which carbon dioxide gas is added into the atmosphere.

D. Carbon Monoxide — A Compound of Carbon

INTRODUCTION

Molecular formula : CO, *Relative molecular mass :* 28

As the name suggests, carbon monoxide is an oxide of carbon containing one atom of carbon and one atom of oxygen in its molecules.

OCCURRENCE

Carbon monoxide occurs in coal gas, volcanic gases, tobacco fumes, chimney gases and in the exhaust gases of automobiles (in trace amounts).

FORMATION OF CARBON MONOXIDE AND ITS ADDITION TO THE ATMOSPHERE

Mostly carbon monoxide is formed when a large amount of carbon or its compounds is burnt in a limited supply of air or oxygen. In other words, *carbon monoxide is a product of incomplete burning or combustion of carbon and fuels.*

$$\underset{\text{(carbon)}}{2C} + \underset{\text{(oxygen)}}{O_2} \xrightarrow{\text{heat}} \underset{\text{(carbon monoxide)}}{2CO}$$

Incomplete burning of fuels : Domestic and industrial ovens running on coal, coke or charcoal produce some carbon monoxide. In an oven, where there is usually free supply of air, the carbon at the lower end burns to produce carbon dioxide. This carbon dioxide passes through the heated layers of coke placed above in the oven, reacts with it and thus gets reduced to carbon monoxide.

$$C + O_2 \rightarrow CO_2 + \text{heat}$$
$$CO_2 + C \rightarrow 2CO$$

The carbon monoxide so produced burns with a pale blue flame at the top of the oven forming carbon dioxide again.

$$2CO + O_2 \rightarrow 2CO_2$$

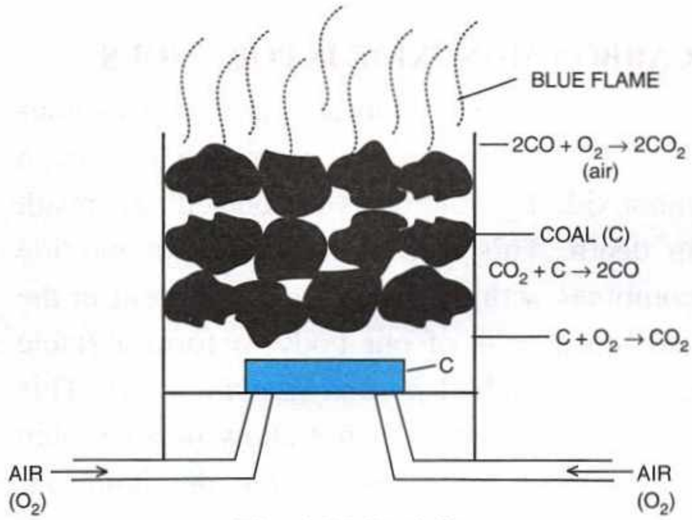


Fig. 9.13 Coal fire

Exhaust gases of automobiles :

In automobile engines, mostly petrol or diesel is used as a fuel. Both contain carbon and hydrogen. They burn within the engine in a limited supply of air, thus forming a little bit of carbon monoxide too.

> *Note :* A fuel is a substance that burns in air to produce usable energy, mostly in the form of heat along with some light energy. The other products are oxides. *Example :* Coal, wood, petroleum, diesel, cooking gas, etc.

Properties of carbon monoxide :

1. Carbon monoxide is a colourless,

tasteless and neutral gas, but it has a faint odour. It is a highly poisonous gas.

2. It is very sparingly soluble in water.
3. It is a thermally stable gas, *i.e.* it does not decompose even at high temperatures.
4. It is a combustible gas, and is a non-supporter of combustion. In air, it burns with a blue flame to form carbon dioxide. (See Fig. 9.13)

CARBON MONOXIDE IS POISONOUS

Carbon monoxide is a highly poisonous gas. If air containing even 0·5% carbon monoxide by volume is inhaled, it can result in death. This is because carbon monoxide combines with the haemoglobin present in the red blood cells of our body to form a stable compound called *carboxyhaemoglobin*. This stable compound does not allow haemoglobin to absorb or carry oxygen, thus depriving our body cells of oxygen. This causes paralysis of the respiratory organs and results in death due to suffocation (asphyxiation).

HARMFUL EFFECTS OF CARBON MONOXIDE, PRECAUTIONS AND REMEDIES FOR ITS POISONING

1. *Sleeping in a closed room with a coal fire burning*

It is very dangerous to sleep in a room where coal or wood is burning and its doors and windows are closed. Owing to the limited supply of air in such a room, carbon monoxide is produced. Since the gas is colourless and has a barely detectable smell, people sleeping in the room do not feel its presence and run the risk of CO poisoning. That is why CO is also called a 'silent killer'.

2. *Starting a car engine inside a closed garage*

Carbon monoxide is almost always present in a small amount in the exhaust gases of automobiles, owing to incomplete combustion of motor fuels. Therefore, it is not advisable to start a car engine inside a closed garage. More of carbon monoxide could thus be produced inside a closed area, increasing the risk of poisoning.

Remedies for carbon monoxide poisoning

(i) The victim should be immediately brought out into the open.

(ii) The victim should be given artificial respiration with *carbogen* (a mixture of 95% oxygen and 5% carbon dioxide) to restore normal breathing.

(iii) At places where carbon monoxide concentration is on the higher side, people should wear gas masks that absorb carbon monoxide and oxidise it to carbon dioxide. Such masks are made of *hopcolite* (a mixture of metallic oxides).

REDUCING ACTION OF CARBON MONOXIDE

Carbon monoxide is a strong reducing agent. It reduces the oxides of the moderately or less active metals to their respective metals and itself gets oxidised to carbon dioxide in the process, *e.g.* :

(i) $\underset{\text{(cupric oxide, black)}}{CuO} + CO \xrightarrow{\text{heat}} \underset{\text{(copper, red/brown)}}{Cu} + CO_2$

(ii) $\underset{\text{(lead monoxide)}}{PbO} + CO \xrightarrow{\text{heat}} \underset{\text{(lead)}}{Pb} + CO_2$

(iii) $\underset{\text{(ferric oxide)}}{Fe_2O_3} + 3CO \xrightarrow{\text{heat}} \underset{\text{(iron)}}{2Fe} + 3CO_2$

Because of its reducing action, carbon monoxide is used in the extraction of pure metals from their corresponding ores. It is widely used in industries in the metallurgy of iron from its ore.

Note : Traditional cooking methods, using firewood or charcoal, especially in rural areas causes a lot of indoor air pollution due to the smoke produced. This leads to more than a million premature deaths in the country every year.

To overcome this, the Government of India has initiated a scheme called **"Pradhan Mantri Ujjwala Yojna"** which recognises the importance of clean cooking energy. It aims to provide Liquified Petroleum Gas (LPG) connections to poor households thereby reducing indoor air pollution.

EXERCISE – IV

1. Fill in the blanks :
   (*a*) .... is formed when carbon is burnt in a limited supply of air or oxygen.
   (*b*) Carbon monoxide burns in air with a .... flame to form carbon dioxide.
   (*c*) Carbon monoxide is a product of .... combustion.
   (*d*) A mixture of 95% oxygen and 5% carbon dioxide is called ....
   (*e*) Carbon monoxide is used as a .... in the extraction of pure metals from their corresponding ores.

2. Match the following :

	Column A	Column B
1.	A product of incomplete burning	(a) Hopcolite
2.	Nature of carbon monoxide	(b) Combustible gas
3.	A compound formed by the combination of haemoglobin and carbon monoxide	(c) Carboxyhaemoglobin
4.	A mixture of metallic oxides	(d) Carbon monoxide
5.	Carbon monoxide	(e) Highly poisonous

3. How is carbon monoxide gas formed ?
4. State the poisonous nature of carbon monoxide.
5. Give *two* uses of carbon monoxide.
6. Why is carbon monoxide called the silent killer ?
7. Explain the reducing action of carbon monoxide.
8. Write *two* remedies for carbon monoxide poisoning.
9. Complete the reactions and balance them.
   (*a*) $CuO + CO \rightarrow$
   (*b*) $Fe_2O_3 + CO \rightarrow$

RECAPITULATION

- Carbon occurs in all living matter, substances derived from living matter (food and fuels), in the earth's crust and in the atmosphere.
- Carbon forms so many compounds because carbon can form long chains. Most compounds of carbon are studied under organic chemistry.
- Graphite, diamond and fullerenes are crystalline allotropes of carbon.
- Diamond is the hardest naturally occurring substance known.
- Fullerenes are discovered only recently.
- Amorphous carbon has different forms : coal, coke, charcoal, lamp black and gas carbon. All are not pure.
- Coal, coke and charcoal are used as fuels. Charcoal can be activated by heating it upto 900°C which increases its adsorbing capacity.
- Coke and charcoal are good reducing agents.
- Bituminous coal is the most common variety of coal.
- Graphite and gas carbon are the two forms of carbon which are good conductors of electricity.
- A fuel is a substance which produces usable heat energy or other forms of energy.
- Most of the fuels produce energy when they burn in air.
- Carbon dioxide is present in air upto 0·04% of the total volume of the atmosphere.
- Carbon dioxide is prepared by burning charcoal or any other fuel in the presence of oxygen.
- In the laboratory, carbon dioxide is prepared by the action of dilute acids on metallic carbonates and bicarbonates.
- Carbon dioxide is a colourless, odourless gas. It is heavier than air and is highly soluble in water.
- Carbon dioxide is neither a supporter of combustion nor does it burn itself. Therefore, it is used as the main anti-fire agent in fire extinguishers.
- Fire extinguisher is a device in which carbon dioxide gas is produced for extinguishing fire.
- Carbon dioxide turns lime water milky.
- Carbon dioxide is an acidic oxide that gives an acidic solution when dissolved in water – carbonic acid.
- Carbon dioxide causes greenhouse effect.
- Carbon monoxide gas is an oxide of carbon formed due to incomplete combustion of carbon and carbon derivatives.
- CO is poisonous and becomes fatal if inhaled in sufficient quantity.
- CO also acts as a reducing agent in the extraction of metals from their oxides.

GLOSSARY

Adsorption : The process due to which substances (gas, liquid or solid) adhere to a surface.

Allotropes : Different forms of an element in the same physical state with identical chemical properties but different physical properties.

Allotropy : A phenomenon due to which an element exists in two or more forms in the same physical state.

Alloy : A homogeneous solid mixture of two or more metals or metals and non metal

Amorphous : The powdered solid in which particles are not arranged in regular geometrical pattern.

Atom : The building block of matter, may or may not have independent existence *i.e.* the smallest unit of matter (element).

Atomic number : The number of protons present in an atom.

Biogas : A mixture of gases obtained by degradation and decomposition of animal and plant matter.

Carbonization : The process of slow conversion of vegetable matter into carbon-rich substances.

Catalyst : Substance which changes the rate of a chemical reaction without undergoing any change itself.

Catalytic hydrogenation : A process in which hydrogen gas is passed through oil in the presence of a catalyst to change it into solid (ghee).

Catenation : The property due to which same kind of atoms link together to form long chains.

Chemical bond : The attractive force which holds the atoms together in a molecule.

Chemical change : A change in which new substances are formed.

Colloid : A homogeneous looking heterogeneous mixture in which particle size ranges from 10^{-10} m to 10^{-7} m dispersed in a continuous medium.

Combination reaction : A reaction in which two or more substances combine chemically to produce a single product.

Combustible : Substances that burn easily in air and produce energy.

Combustion : Burning of a substance in air to produce heat and light energy.

C.N.G. : Compressed Natural Gas (methane is the main constituent).

Crystalline : The solid in which particles are arranged in a regular geometrical pattern due to which they have definite shape.

Decomposition reaction : A reaction in which a substance decomposes to produce two or more new substances.

Deliquescence : The phenomenon due to which some substances absorb moisture, dissolve in it and turn into solution.

Destructive distillation : Heating of a substance in the absence of air or oxygen.

Displacement reaction : A reaction in which a more reactive element displaces a less active element from its compound.

Distillation : The process of converting a liquid into its vapour by boiling and then condensing the vapour into the liquid state on cooling.

Double displacement reaction : A reaction in which two compounds exchange their radicals to produce two new compounds in solution.

Efflorescence : The phenomenon due to which some crystalline substnaces loose water of crystallisation and become anhydrous.

Electrodes : The solid conductor through which electric current enters and leaves the electrolyte.

Electrolysis : The process in which electricity is passed through a substance in its aqueous or molten state to bring about some chemical change.

Electrolyte : The substance in molten or aqueous solution state through which electricity passes during electrolysis.

Electronic configuration : Arrangement of electrons in different shells/orbits of an atom surrounding the nucleus.

Electrons : The negatively charged subatomic particles with negligible mass which revolve round the nucleus of an atom.

Endothermic reaction : Reaction in which the energy is absorbed from the surroundings.

Enzymes : Complex organic compounds acting as catalysts in biochemical reactions.

Exothermic reaction : Reaction in which energy is given out to the surrounding in the form of heat and light.

Fire extinguisher : A device used to extinguish fire.

Flame : A zone of combustion of gaseous substances.

Fractional distillation : A process of distillation employed to separate the components of a homogeneous liquid mixture on the basis of different boiling points of the components.

Fuel : A substance used to produce usable heat energy.

Fullerene : The third recently discovered crystalline form of carbon.

Green house effect : Trapping of the earth's radiated energy by carbon dioxide present in air to keep the earth warm.

Hygroscopic substances : Substances that absorb moisture from other substances but do not turn into solution and can act as drying agents.

Ignition temperature : The minimum temperature at which a substance begins to burn.

Inflammable : A substance that burns with a flame and has low ignition temperature.

Ions : Charged particles formed due to loss or gain of electrons from atoms.

Isotopes : Atoms with same atomic number but different mass number.

L.P.G. : Liquefied petroleum gas. [Contains isobutane and butane].

Mass number : Sum of protons and neutrons present in the nucleus of an atom.

Metal activity series : A series of metals in which they are arranged in the descending order of their reactivity.

Molecules : The smallest particle of matter with independent existence.

Neutrons : The subatomic particles present in the nucleus of an atom with unit mass but no charge.

Non-combustible : A substance that does not burn in air or oxygen.

Non-electrolyte : A substance that does not conduct electricity in liquid state.

Nucleons : The particles present in the nucleus of an atom. [Protons and neutrons are collectively called as nucleons].

Nucleus : The central, dense part of an atom containing protons and neutrons which is responsible for the total mass of an atom.

Oxidation : The process in which a substance combines with oxygen.

Oxidising agent : A substance which helps in oxidation.

Petroleum : A dark coloured, viscous mixture of hydrocarbons obtained from earth's crust.

Physical change : A change in which no new substance is formed.

Precipitate : Insoluble solid formed due to chemical combination of two solutions.

Promoter : It increases the efficiency of a catalyst.

Protons : Positively charged particles with unit mass present in the nucleus of an atom.

Radicals : A single atom or group of atoms which is charged and behaves as a single unit.

Redox reaction : A reaction in which reduction and oxidation take place simultaneously.

Reducing agent : A substance which helps in reduction.

Reduction : The process in which a substance combines with hydrogen.

Shells/Orbits : The fixed circular paths along which electrons revolve round the nucleus of an atom.

Valence electrons : The number of electrons present in the outermost shell of an atom.

Valency : It is the combining capacity of an element. It is equal to the number of electrons lost, gained or shared by an atom during chemical combinations.