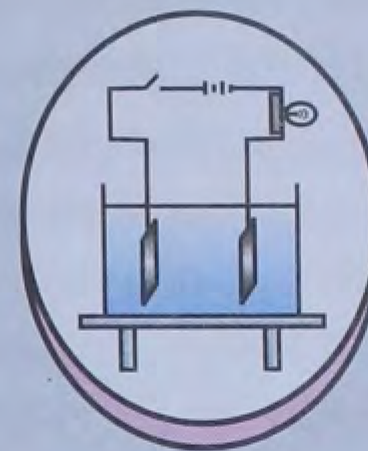
3 TRANSFORMATION OF SUBSTANCES



In This Chapter You Will Learn :

- *Change of state – physical changes*
- *Chemical reactions involving energy changes*
- *Boiling and melting points of pure substances*
- *Chemical equations and simple calculations*
- *Characteristics of chemical reactions*
- *Types of chemical reactions*
- *Distillation*
- *Electrolysis and its applications*

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. In this chapter you will learn about the different types of changes and transformation that substances and their derivatives can undergo.

A. Physical Change, Chemical Change and Chemical Reactions

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

1. Physical change 2. Chemical change

3.1 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.

3.1.1 Characteristics of 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. 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*

***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.

(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 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 masses 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 ice, sublimation of iodine, 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, etc., are just a few common examples to name.

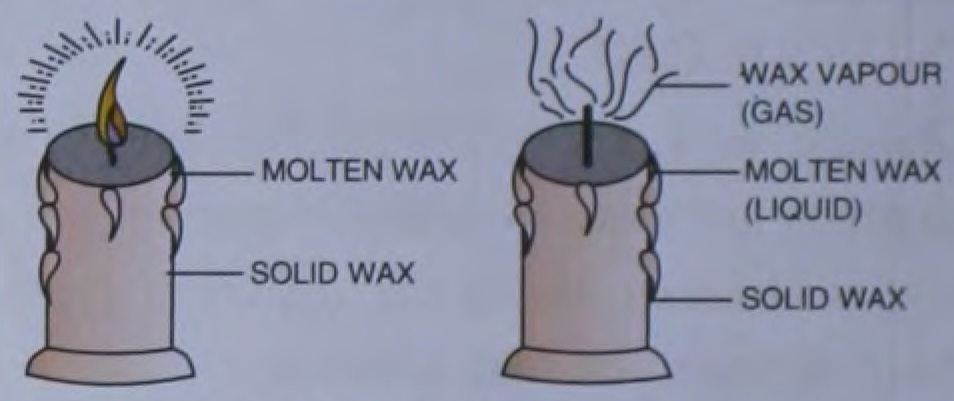


Fig. 3.1 Change of states of wax.

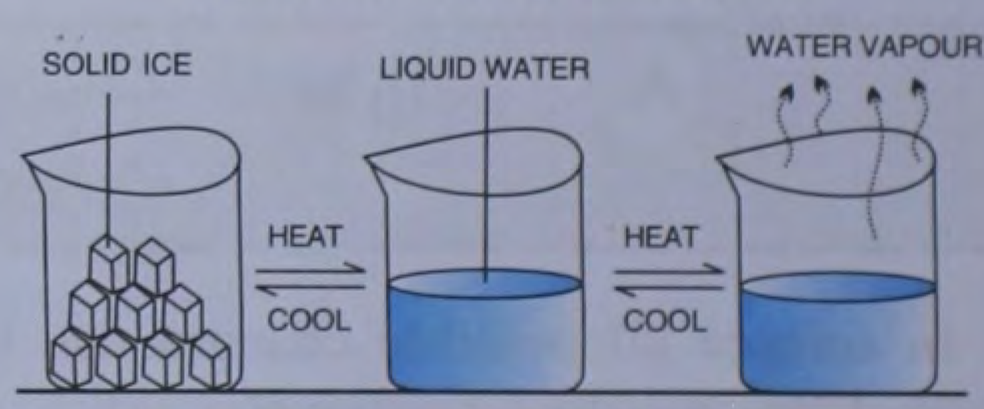
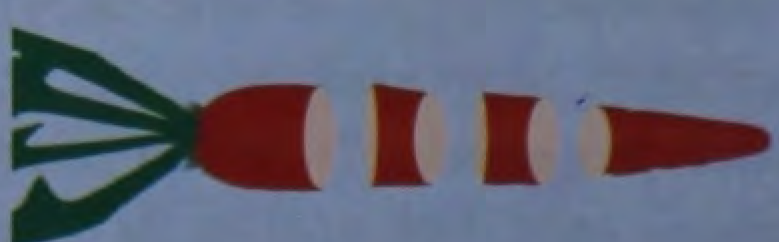


Fig. 3.2 Interconversion of states of ice.

Activity 1

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 temporary 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*. might 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.

3.2 CHEMICAL CHANGE

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

Characteristics of chemical change :

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

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

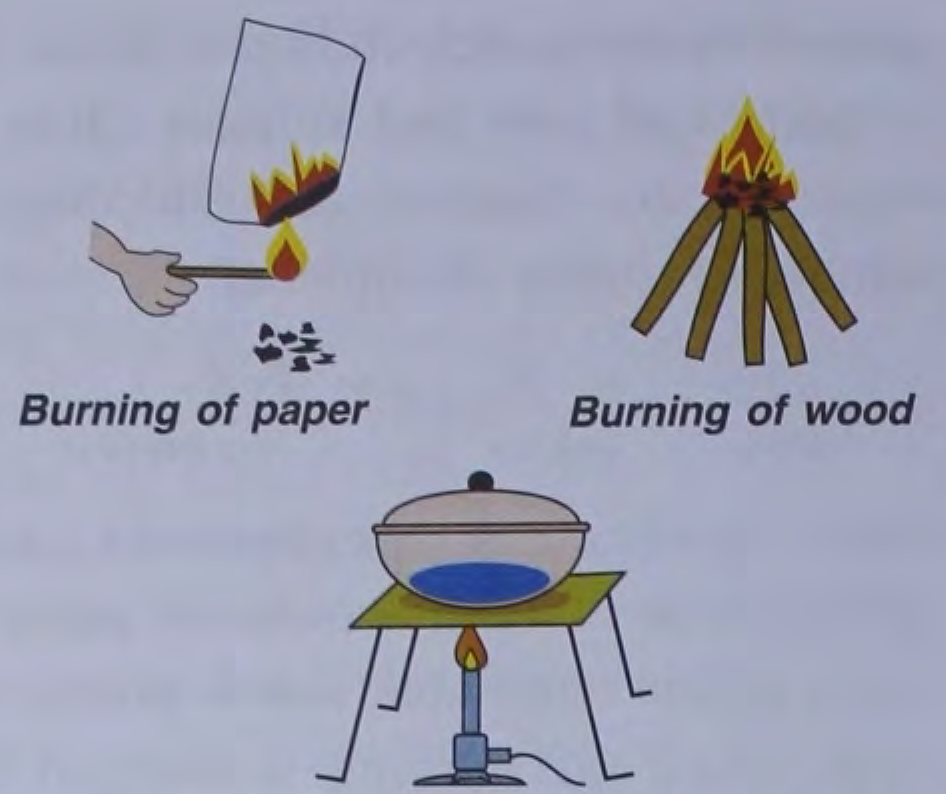
2. **The change is permanent and irreversible :** When a piece of paper is burnt a new substance *ash* is produced. Even when the burning is stopped the ash cannot be changed back into paper. This shows that the formation of the ash from paper is a permanent and irreversible change.

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 the 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.

4. **Exchange of energy takes place :** When wood 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.



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

Fig. 3.3 To show chemical change.

The characteristics of chemical change can be summarized as follows :

1. New substance(s) is/are formed.
2. The composition of the original substance completely changes.
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, *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 turns into vapour to produce a flame. New substances CO_2 and H_2O vapour are formed alongwith the evolution of light and heat energy. This shows a chemical change.

When some of the molten wax drops to 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.

3.3 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.

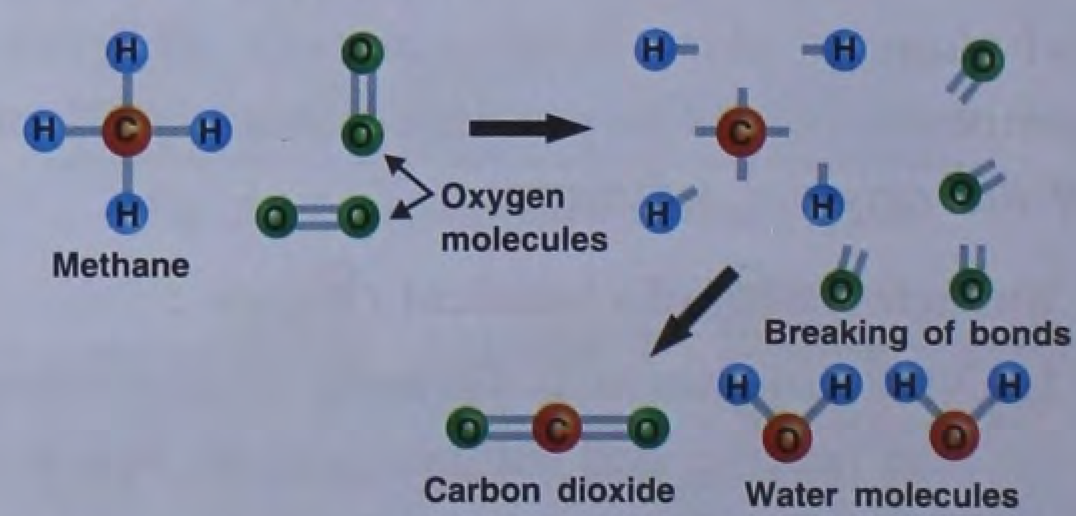


Fig. 3.4 A diagrammatic representation of breaking of bonds and formation of new bonds.

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	→	$NaCl$	+	H_2O
[sodium hydroxide]		[hydrochloric acid]		[sodium chloride]		[water]
(Reactants)				(Products)		

Sodium hydroxide and dilute HCl are reactants and sodium chloride and water are products.

3.3.1 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	→	$ZnSO_4$	+	H_2
[zinc]		[dil. sulphuric acid]		[zinc sulphate]		[hydrogen]

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

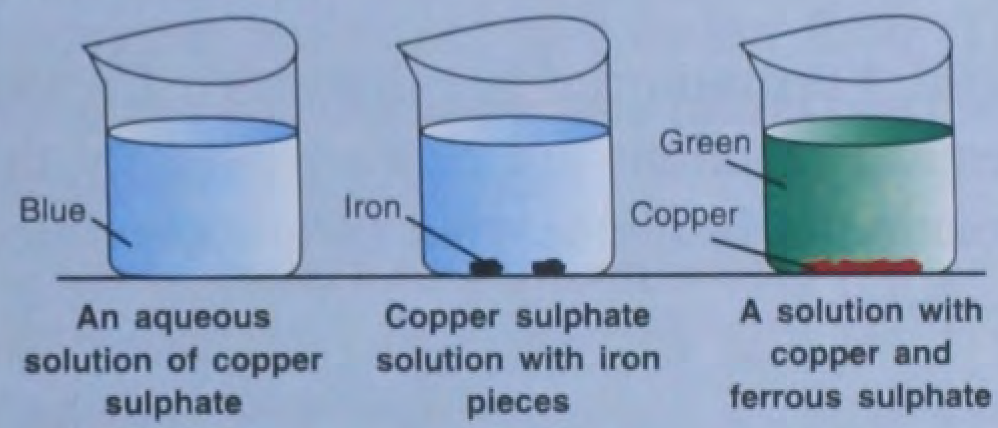
Na_2SO_3	+	$2HCl$	→	$2NaCl$	+	H_2O	+	SO_2
[sodium sulphite]		[dil. hydro-chloric acid]		[sodium chloride]		[water]		[sulphur dioxide]

Effervescence : The formation of gas bubbles in a liquid during a reaction is called *effervescence*. In reactions (a) and (c), one of the reactants is a liquid. In such cases, *i.e.* 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 2

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 few hours. What do you observe ?

....

Why has the colour changed ?

....

What conclusion do you draw ?

....

Give 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	+	$CuSO_4(aq)$	→	$FeSO_4$	+	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***.

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.

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(ppt) + NaNO_3(aq)$$

[silver nitrate solution] [sodium chloride solution] [silver chloride] (white ppt) [sodium nitrate solution]

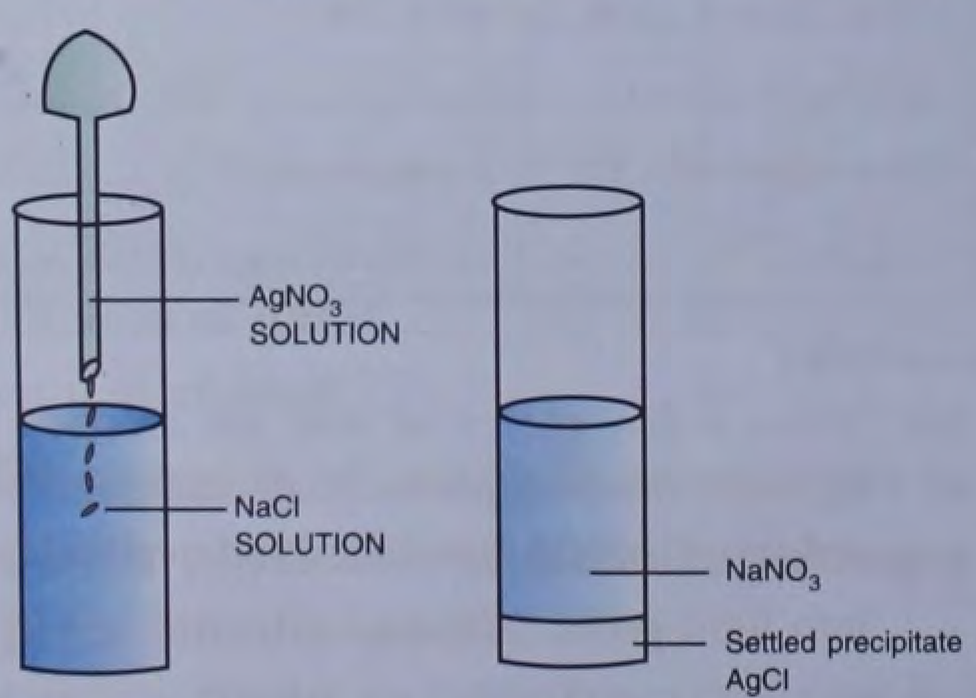


Fig. 3.3 Precipitation of silver chloride.

(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 calcium bicarbonate solution is heated, a white precipitate, calcium carbonate, is produced.

$$Ca(HCO_3)_2 \xrightarrow{heat} CaCO_3 + CO_2 + H_2O$$

[calcium bicarbonate solution] (while ppt.) [calcium carbonate] [carbon dioxide] [water]

(iv) Change in energy : 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 given out, 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 *exothermic reaction.* It causes rise in temperature.

Examples

(a) When carbon burns in oxygen to form carbon dioxide, a lot of heat is produced.

$$C + O_2 \rightarrow CO_2 + Heat$$

[carbon] [oxygen] [carbon dioxide]

(b) When water is added to quicklime a lot of heat energy is produced [alongwith alkaline calcium hydroxide (slaked lime)], which boils water.

$$\underset{\text{[quicklime]}}{CaO} + H_2O \rightarrow \underset{\text{[slaked lime]}}{Ca(OH)_2} + \text{Heat}$$

Respiration, rusting, burning of coal, petrol, kerosene, etc., are some common exothermic reactions.

2. ***Endothermic reaction*** **:** A chemical reaction in which heat is absorbed is called *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.

$$\underset{\text{[nitrogen]}}{N_2} + \underset{\text{[oxygen]}}{O_2} + \underset{(3000°C)}{\text{heat}} \rightarrow \underset{\text{[nitric oxide]}}{2NO\,(g)}$$

(b) Similarly, decomposition of calcium carbonate into carbon dioxide and calcium oxide when it is heated to a temperature of about 1000°C is an endothermic reaction.

$$\underset{\text{[calcium carbonate]}}{CaCO_3\,(s)} + \underset{(1000°C)}{\text{heat}} \rightarrow \underset{\text{[calcium oxide]}}{CaO\,(s)} + \underset{\text{[carbon dioxide]}}{CO_2\,(g)}$$

(v) 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 ends up with gaseous products, and vice versa.

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

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

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.

3.3.2 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 come into contact with each other.

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

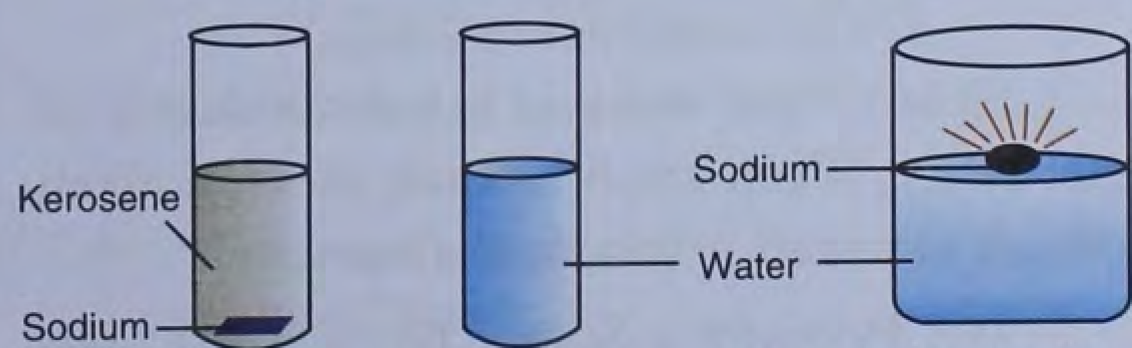
$$2Na + \underset{\text{[cold]}}{2H_2O} \xrightarrow{\text{vigorous}} 2NaOH + H_2$$

(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 into contact with each other at a faster rate.

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

$$NaCl\,(aq) + AgNO_3\,(aq) \rightarrow \underset{\text{(White precipitate)}}{AgCl\,(\downarrow)} + NaNO_3\,(aq)$$

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



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

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

A mixture of iron and sulphur

Iron and sulphur react on heating to form a new substance iron sulphide

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

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

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

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

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

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

(v) Electricity : Some chemical reactions occur only when electricity is passed through the reactant.

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

$$2H_2O \xrightarrow[\text{current}]{\text{electric}} 2H_2 + O_2$$

Note : Pure water is a bad conductor of electricity. When small amount of acid, alkali or salt is added 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*.

$$N_2 + 3H_2 \underset{\text{200 atm}}{\overset{450°C}{\rightleftharpoons}} 2NH_3$$

* ***This reaction takes place in the presence of iron as catalyst.***

(vii) Catalyst : 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.

$$N_2 + 3H_2 \underset{\substack{450°C \\ \text{Mo (promoter)}}}{\overset{\substack{\text{Iron (catalyst)} \\ 200-900\text{ atm}}}{\rightleftharpoons}} 2NH_3 + \text{Heat}$$

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.

(b) *Negative catalyst* : When a catalyst decreases the rate of a chemical reaction it is known as negative catalyst.

Examples : Phosphoric acid acts as a negative catalyst to decrease the rate of the 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 bodies of living beings.

For example : The process of digesting the food is a series of chemical reactions that are accelerated by enzymes.

Amylase, present in saliva, breaks starch into sugar.

Pepsin, in the stomach, breaks proteins into amino acids.

Trypsin and *lipase,* in the small intestine, help to break down starch, fat and protein into glucose, fatty acids and amino acids respectively.

Enzymes present in the living cells of yeast help in the conversion of *maltose* (a carbohydrate) into glucose.

$$\text{Maltose + Water} \xrightarrow{\text{Maltase (enzyme)}} \text{Glucose}$$

The enzyme *zymase,* which is present in yeast cells, helps in the fermentation of glucose into alcohol.

$$\text{Glucose} \xrightarrow{\text{Zymase}} \text{Ethyl alcohol} + 2CO_2$$

EXERCISE – I

1. Fill in the blanks :
   (a) The process of a liquid changing into a solid is called ....
   (b) The temperature, at which a solid changes into a liquid, is called ....
   (c) A change, which alters the composition of a substance, is known as a .... change.
   (d) There is no change in the .... of the substance during a physical change.
   (e) The reaction in which energy is evolved is called ....
2. Define :
   (a) a physical change.
   (b) a chemical change.
3. Classify the following as a physical or a chemical change :
   (a) Drying of wet clothes
   (b) Manufacture of salt from sea water
   (c) Making of curd from milk
   (d) Butter getting rancid
   (e) Growth of a tree
   (f) Rusting of iron
   (g) Boiling of water
   (h) Burning of paper
4. Give *one* example each which illustrates the following characteristics of a chemical reaction:
   (a) evolution of a gas
   (b) change of colour
5. Which of the following is not a physical change ? Why ?
   (a) Freezing of water.
   (b) Powdered sulphur when heated gently.
   (c) Magnetisation of a piece of iron.
   (d) Burning of a piece of magnesium wire in air.
6. Give reasons to show that the following are chemical changes:
   (a) Burning of magnesium in air.
   (b) Adding of sodium to water.
7. How do the following help in bringing about a chemical change ?
   (a) pressure (b) light
   (c) catalyst (d) heat.
8. Define exothermic and endothermic reactions. Give two examples in each case.

9. Give reason :
   (a) Freezing of water to ice and evaporation of water are physical changes.
   (b) A glass tube broken into pieces is a physical change.
   (c) Burning of a candle is both a physical and a chemical change.
10. State the effect of :
   (a) an endothermic reaction
   (b) an exothermic reaction on the surroundings.
11. (a) Define the term catalyst.
   (b) What are (i) positive catalysts and (ii) negative catalysts ? Support your answer with one example for each them.
   (c) Name three biochemical catalysts found in the human body.
12. 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) calcium bicarbonate solution is heated ?
13. 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) $Zn + \underset{(dil)}{H_2SO_4} \rightarrow$
   (f) $\underset{(aq.)}{FeSO_4} + \underset{(aq.)}{NaOH} \rightarrow$
   (g) $Pb(NO_3)_2 \xrightarrow{heat}$
   (h) $Ca(HCO_3)_2 \xrightarrow{heat}$

B. Types of Chemical Reactions

3.4 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 decomposition reaction
5. Oxidation-reduction reaction (or redox reaction)

3.4.1 Combination reaction

A reaction in which two or more substances combine to form a single substance is called *combination reaction*.

$$A + B \rightarrow AB$$

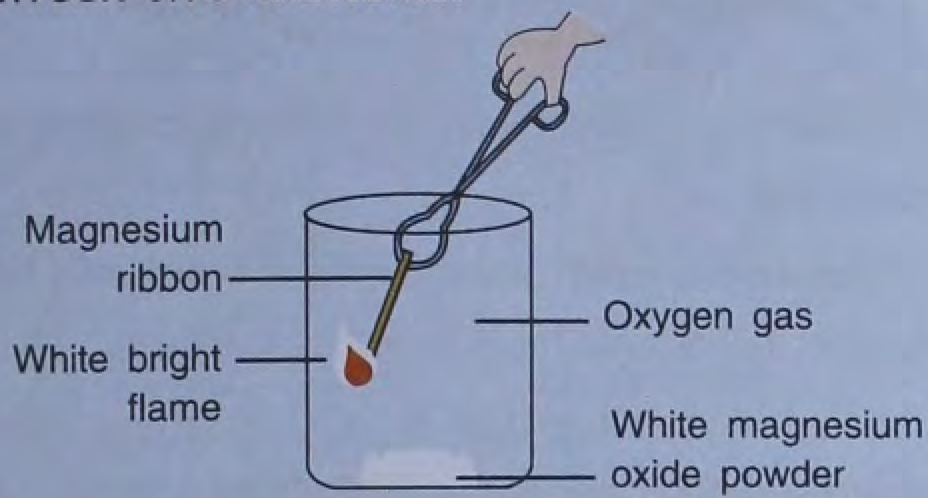
In the above reaction the combination of substances A and B takes place to give a molecule of a new substance AB

In combination reactions :

(i) two elements can combine to form a compound; [This reaction is also called **synthesis**]

Activity 3

To demonstrate a combination reaction between two elements.



Take a magnesium ribbon and hold it with a tong 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.

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

(b) Similarly, carbon burns in oxygen to form a gaseous compound, carbon dioxide.

$$\underset{\text{[carbon]}}{C} + \underset{\text{[oxygen]}}{O_2} \xrightarrow{\text{heat}} \underset{\text{[carbon dioxide]}}{CO_2} + \text{Heat}$$

(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.

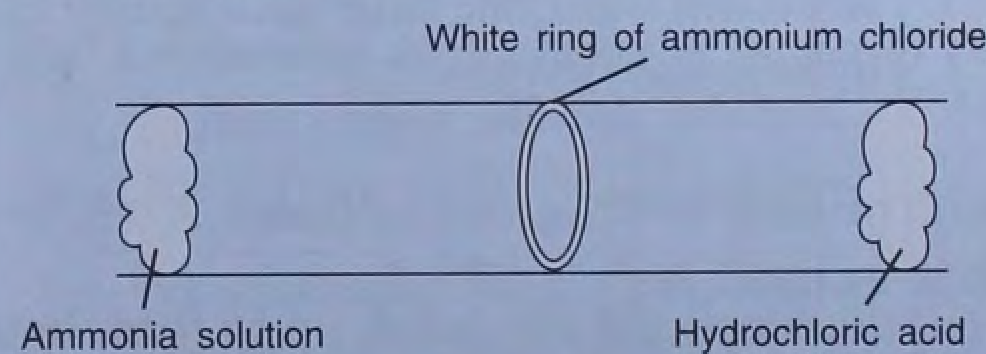
$$\underset{\text{[ammonia]}}{NH_3\,(g)} + \underset{\text{[hydrogen chloride]}}{HCl\,(g)} \rightarrow \underset{\text{[ammonium chloride]}}{NH_4Cl\,(s)}$$

Activity 4

To demonstrate 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 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) $$\underset{\text{[phosphorus trichloride]}}{PCl_3} + \underset{\text{[chlorine]}}{Cl_2} \xrightarrow{\text{heat}} \underset{\text{[phosphorus pentachloride]}}{PCl_5}$$

(ii) $$\underset{\text{[sulphur dioxide]}}{SO_2} + \underset{\text{[water]}}{H_2O} \rightarrow \underset{\text{[sulphurous acid]}}{H_2SO_3}$$

(iii) $$\underset{\text{[nitrogen pentoxide]}}{N_2O_5} + \underset{\text{[water]}}{H_2O} \rightarrow \underset{\text{[nitric acid]}}{2HNO_3}$$

3.4.2 Decomposition reaction

A reaction in which a compound breaks

up due to the application of heat into two or more simple 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.

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.

$$2H_2O\ (l) \xrightarrow[\text{current}]{\text{electric}} 2H_2\ (g) + O_2\ (g)$$

(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 compounds.

Example : The compound calcium carbonate decomposes on strong heating to form two compounds, calcium oxide and carbon dioxide.

$$\underset{\text{[calcium carbonate]}}{CaCO_3\ (s)} \xrightarrow[100°C]{heat} \underset{\text{[calcium oxide]}}{CaO\ (s)} + \underset{\text{[carbon dioxide]}}{CO_2\ (g)}$$

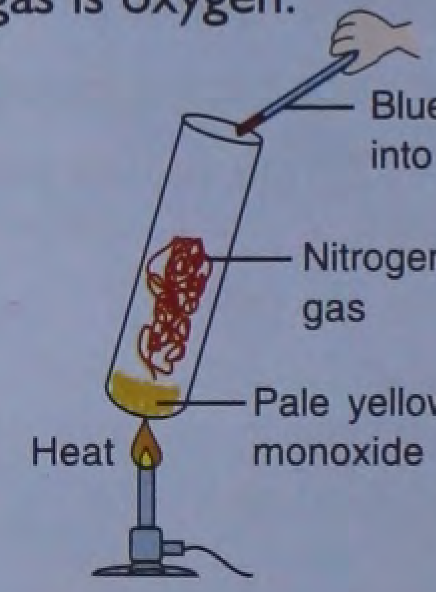
Activity 5

To demonstrate thermal decomposition of a compound.

Take some lead nitrate 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 and the colourless gas is oxygen.



$$2Pb(NO_3)_2 \xrightarrow{heat} 2PbO + 4NO_2 + O_2$$

More examples of a decomposition reaction are :

(i) $$\underset{\text{[potassium chlorate]}}{2KClO_3\ (s)} \xrightarrow{heat} \underset{\text{[potassium chloride]}}{2KCl\ (s)} + \underset{\text{[oxygen]}}{3O_2\ (g)}$$

(ii) $$\underset{\text{[copper hydroxide]}}{Cu(OH)_2\ (s)} \xrightarrow{heat} \underset{\text{[copper oxide]}}{CuO\ (s)} + \underset{\text{[water]}}{H_2O\ (l)}$$

(iii) $$\underset{\text{[copper carbonate]}}{CuCO_3\ (s)} \xrightarrow{heat} \underset{\text{[copper oxide]}}{CuO\ (s)} + \underset{\text{[carbon dioxide]}}{CO_2\ (g)}$$

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

Action of heat on metal nitrates to show decomposition reaction on the basis of metal activity series

Metal Activity series (Decreasing reactivity ↓)	Reaction on heating
K, Na	$2KNO_3 \xrightarrow{\Delta} 2KNO_2 + O_2$
Ca, Mg, Al, Zn, Fe, Pb, [H], Cu	$2Ca(NO_3)_2 \xrightarrow{\Delta} 2CaO + 4NO_2 + O_2$ $2Cu(NO_3)_2 \xrightarrow{\Delta} 2CuO + 4NO_2 + O_2$
Hg, Ag	$2Hg(NO_3)_2 \xrightarrow{\Delta} 2Hg + 4NO_2 + O_2$
Pt, Au	

(i) Potassium nitrate decomposes to give potassium nitrite and oxygen, only due to greater activity of potassium, and so does sodium.

(ii) Nitrates of metals, calcium to copper, decompose to give respective oxides, nitrogen dioxide and oxygen.

(iii) While nitrates of mercury and silver decompose to give respective metal, nitrogen dioxide and oxygen.

(iv) Platinum and gold do not form nitrates.

3.4.3 Displacement reaction

A reaction in which a more active element displaces a less active element from a compound is called   . In such reactions one constituent of the reactant molecule is replaced by another.

$$AB + C \rightarrow CB + A$$

Here C displaces A from the molecule AB, since C is chemically more active as compared to A.

(i) A more reactive metal displaces a less reactive metal from its salt solution.

Example : Zinc being more reactive than copper displaces copper from copper sulphate solution the blue copper sulphate solution turns colourless due to formation of zinc sulphate and a reddish brown copper is deposited on zinc.

Zn	+	$CuSO_4$ (*aq*)	→	$ZnSO_4$(*aq*)	+	Cu
[zinc]		[blue]		[colourless]		[reddish brown]

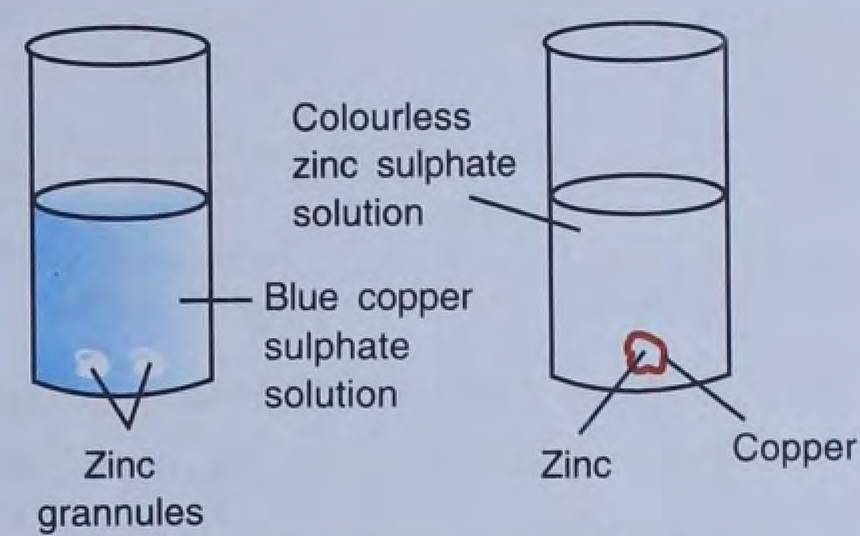


Fig. 3.4 Showing displacement of copper by zinc from copper sulphate solution.

(ii) A metal more reactive than hydrogen displaces hydrogen gas from an acid or an alkali or from water.

Examples

(a) Zinc being an active metal, it displaces hydrogen from dilute hydrochloric (or dilute sulphuric) acid.

Zn	+	2HCl (dil)	→	$ZnCl_2$	+	H_2
[zinc]		[hydrochloric acid]		[zinc chloride]		[hydrogen]
Zn	+	H_2SO_4 (dil)	→	$ZnSO_4$	+	H_2
[zinc]		[sulphuric acid]		[zinc sulphate]		[hydrogen]

(b) Similarly zinc displaces hydrogen from alkalis like sodium hydroxide.

Zn	+	2NaOH	→	Na_2ZnO_2	+	H_2
[zinc]		[sodium hydroxide]		[sodium zincate]		[hydrogen]

(c) Zinc reacts with steam too to give hydrogen gas and zinc oxide.

$$\underset{\text{[zinc]}}{Zn} + \underset{\text{[steam]}}{H_2O} \xrightarrow{\text{heat}} \underset{\text{[zinc oxide]}}{ZnO} + \underset{\text{[hydrogen]}}{H_2}$$

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].

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

$$\underset{\text{[potassium]}}{2K} + \underset{\text{[cold water]}}{2H_2O} \xrightarrow[\text{reaction}]{\text{vigorous}} \underset{\text{[potassium hydroxide]}}{2KOH} + \underset{\text{[hydrogen]}}{H_2}$$

(iii) A more reactive non-metal displaces a less reactive non-metal from the latter's solution.

Example : Chlorine being more active than bromine, it displaces bromine from sodium bromide solution.

$$\underset{\text{[sodium bromide]}}{2NaBr} + \underset{\text{[chlorine]}}{Cl_2} \rightarrow \underset{\text{[sodium chloride]}}{2NaCl} + \underset{\text{[bromine]}}{Br_2}$$

> Displacement reactions help us to distinguish between more reactive and less reactive elements.

3.4.4 Double decomposition reaction

A chemical reaction in which two compounds in their aqueous state exchange their ions to form new compounds is called a *double decomposition reaction* or a *double displacement reaction.*

$$AB + CD \rightarrow CB + AD$$

Here AB and CD are the two reactant molecules. They exchange their ions (or radicals) to form two new molecules, CB and AD. Double decomposition reactions are of two types :

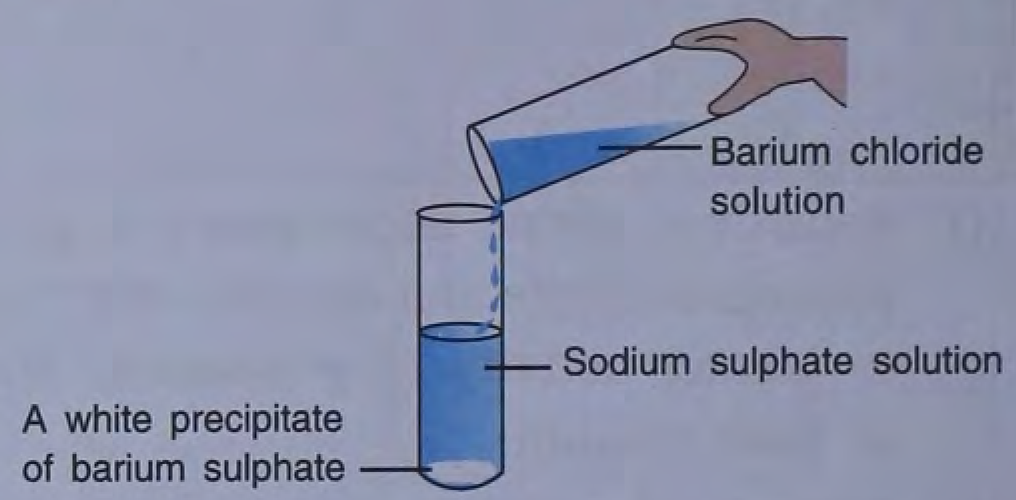
(a) Precipitation reaction

(b) Neutralization reaction

(a) *Precipitation reaction* : A chemical reaction in which two compounds in their aqueous state react to form an insoluble salt (a precipitate) as one of the products is known as a *precipitation reaction.*

Example : When barium chloride solution reacts with sodium sulphate solution, an insoluble white precipitate, barium sulphate, and a soluble salt, sodium chloride, are obtained.

$$\underset{\text{[barium chloride]}}{BaCl_2(aq)} + \underset{\text{[sodium sulphate]}}{Na_2SO_4(aq)} \rightarrow \underset{\text{[barium sulphate white precipitate]}}{BaSO_4\downarrow^*} + \underset{\text{[sodium chloride]}}{2NaCl(aq)}$$



Some more examples of precipitation reaction are :

(i) $\underset{\text{[silver nitrate]}}{AgNO_3(aq)} + \underset{\text{[hydrochloric acid]}}{HCl} \rightarrow \underset{\text{[silver chloride] (white ppt.)}}{AgCl(\downarrow)} + \underset{\text{[nitric acid]}}{HNO_3(aq)}$

(ii) $\underset{\text{[copper sulphate]}}{CuSO_4(aq)} + \underset{\text{[sodium hydroxide]}}{2NaOH(aq)} \rightarrow \underset{\text{[copper hydroxide] (blue ppt.)}}{Cu(OH)_2(\downarrow)} + \underset{\text{[sodium sulphate]}}{Na_2SO_4(aq)}$

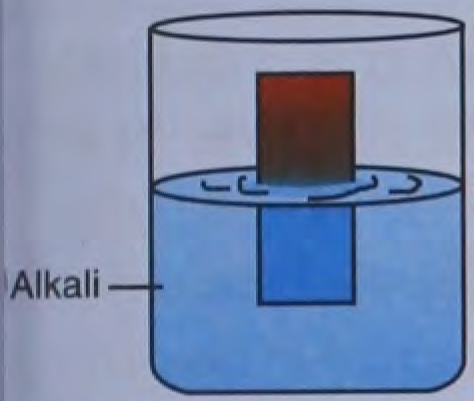
(iii) $\underset{\text{[lead nitrate]}}{Pb(NO_3)_2} + \underset{\text{[ammonium hydroxide]}}{2NH_4OH} \rightarrow \underset{\text{[lead hydroxide] (white ppt.)}}{Pb(OH)_2(\downarrow)} + \underset{\text{[ammonium nitrate] (aq)}}{2NH_4NO_3}$

* ***A precipitate is indicated by an arrow pointing downwards (↓).***

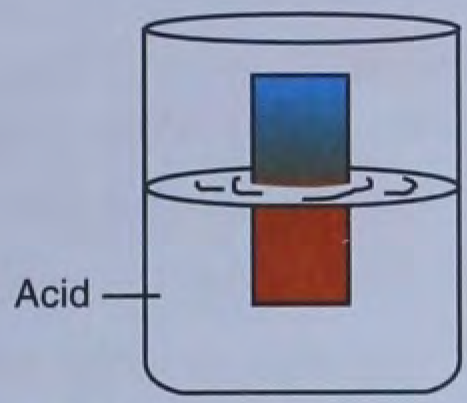
(iv) $CaCl_2(aq) + 2NaOH(aq) \rightarrow Ca(OH)_2(\downarrow) + 2NaCl(aq)$
[calcium chloride] [sodium hydroxide] [calcium hydroxide] (white ppt.) [sodium chloride]

(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 *neutralization reaction*.

Acid + *Base* → *Salt* + *Water*



Red litmus changed into blue in an alkaline solution



Blue litmus changed into red in an acid solution

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]

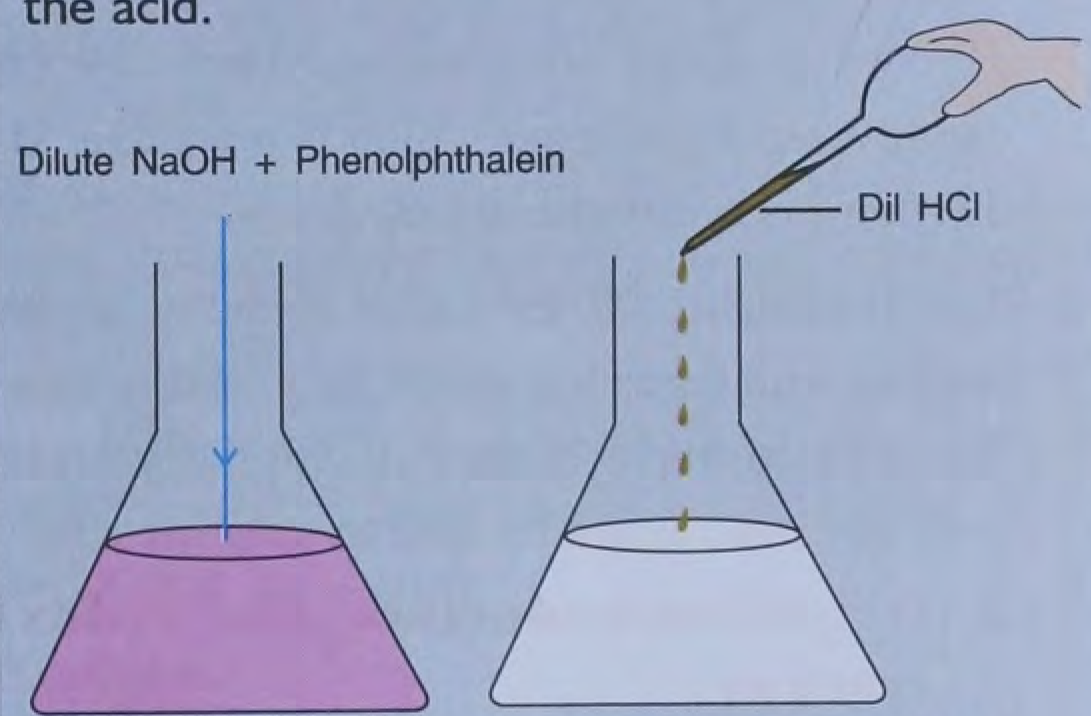
(ii) $CuO(s) + 2HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + H_2O$
[copper oxide] [nitric acid] [copper nitrate] [water]

All metallic oxides and metallic hydroxides are called *bases*, whereas those bases that dissolve in water are known as alkalis.

Activity 6

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 truns 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 has turned into an acidic solution. The base is completely neutralized by the acid.



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.

COLOUR IN

Indicator	Acidic solution	Basic solution
1. Litmus	Red	Blue
2. Methyl orange	Red or pink	Yellow
3. Phenolphthalein	Colour less	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. Many toothpastes contain bases to neutralize the acids formed in the mouth.
3. **Insect sting :** When a bee stings it injects an acidic liquid 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.

3.4.5 Oxidation–reduction (or redox) reaction

(a) *Oxidation* : Oxidation is defined as a chemical process that involves :

(1) *Addition of oxygen* :

Element + Oxygen → Oxide

Examples

(a) Carbon reacts with oxygen to form carbon dioxide. Carbon gets oxidised.

$$\underset{\text{[carbon]}}{C} + \underset{\text{[oxygen]}}{O_2} \xrightarrow{\text{heat}} \underset{\text{[carbon dioxide]}}{CO_2}$$

(b) Magnesium reacts with oxygen to form magnesium oxide. Magnesium gets oxidised by addition of oxygen.

$$\underset{\text{[magnesium]}}{2Mg} + \underset{\text{[oxygen]}}{O_2} \xrightarrow{\text{heat}} \underset{\text{[magnesium oxide]}}{2MgO}$$

(2) *Removal of hydrogen :*

Example : When hydrogen sulphide reacts with chlorine it gets oxidised to sulphur due to loss of hydrogen.

$$\underset{\text{[hydrogen sulphide]}}{H_2S} + \underset{\text{[chlorine]}}{Cl_2} \longrightarrow \underset{\text{[hydrochloric acid]}}{2HCl} + \underset{\text{[sulphur]}}{S}$$

(3) *Loss of electrons by an atom or by an ion.*

Examples

(a) The zinc atom loses two electrons to get oxidised into an electropositive zinc ion.

$$Zn - 2e^- \xrightarrow{\text{oxidation}} Zn^{2+}$$

(b) The chloride (Cl^{-1}) ion loses one electron to form the neutral chlorine atom.

$$Cl^- - e^- \xrightarrow{\text{oxidation}} Cl$$

(c) The ferrous ion (Fe^{2+}) loses one electron to form ferric ion (Fe^{3+}).

$$Fe^{2+} - e^- \rightarrow Fe^{3+}$$

***Oxidising agent* :** The atom or group of atoms that gains or accepts electron(s) during an oxidation reaction is called *oxidising agent.* Some common oxidising agents are oxygen, chlorine, potassium permanganate, carbon dioxide, conc. nitric acid, conc. sulphuric acid, *etc*. The role of an oxidising agent can be explained with the help of the following examples.

Examples

(a) We know that magnesium reacts with oxygen to form magnesium oxide.

$$2Mg + O_2 \xrightarrow{\text{eat}} 2MgO$$

In this reaction magnesium loses two electrons to get oxidised and oxygen accepts those two electrons, *i.e.* oxygen acts as an oxidising agent.

$$[Mg - 2e^- \rightarrow Mg^{2+}] \times 2$$
$$[O + 2e^- \rightarrow O^{2-}] \times 2$$
$$\mathbf{2Mg + O_2 \rightarrow 2MgO}$$

(b) Conversion of ferrous chloride into ferric chloride involves loss of electrons by the ferrous ion and gain of electrons by chlorine atom. In this reaction ferrous gets oxidised into ferric ion and chlorine acts as an oxidising agent.

$$2FeCl_2 + Cl_2 \rightarrow 2FeCl_3$$

$$\underset{\text{(ferrous ion)}}{Fe^{2+}} - e^- \rightarrow \underset{\text{(ferric ion)}}{Fe^{3+}}$$

Oxidation is defined as a chemical process that involves addition of oxygen to a substance or removal of hydrogen from a substance or loss of electrons from an atom or a group of atoms or from an ion.

(b) *Reduction* : Reduction is the reverse of oxidation. As such it is defined as a chemical process that involves :

(1) *Addition of hydrogen :*

Example : When nitrogen reacts with hydrogen under specific conditions, ammonia gas is formed. Since hydrogen is added to nitrogen, this is a reduction reaction.

$$N_2 + 3H_2 \xrightarrow[200-900\text{ atm. } 450°C]{Fe} 2NH_3$$

(2) *Removal of oxygen :*

Oxide → Element + Oxygen

Examples

(a) When mercuric oxide is heated it loses oxygen to form mercury a reduced product.

$$\underset{\text{[mercuric oxide]}}{2HgO} \xrightarrow{heat} \underset{\text{[mercury]}}{Hg} + \underset{\text{[oxygen]}}{O_2}$$

(b) When copper oxide reacts with hydrogen, copper is formed by the removal of oxygen from copper oxide, *i.e.* copper oxide is reduced.

$$\underset{\text{[copper oxide]}}{CuO} + \underset{\text{[hydrogen]}}{H_2} \xrightarrow{heat} \underset{\text{[copper]}}{Cu} + \underset{\text{[water]}}{H_2O}$$

(3) *Gain of electrons by an atom or an ion.*

Examples

(a) Cupric ion gains two electrons to get reduced to copper metal.

$$\underset{\text{[cupric ion]}}{Cu^{2+}} + 2e^- \xrightarrow{Reduction} \underset{\text{[copper metal]}}{Cu}$$

(b) When oxygen accepts two electrons it gets reduced to an electronegative oxide ion.

$$\underset{\text{[oxygen]}}{O_2} + 2e^- \xrightarrow{Reduction} \underset{\text{[oxide ion]}}{2O^{2-}}$$

Reducing agent : The atom or the group of atoms or the ion that loses electron(s) during a chemical reaction is called *reducing agent*. Some common reducing agents are hydrogen, hydrogen sulphide, potassium, carbon, sulphur dioxide, *etc*.

The action of a reducing agent can be better understood with the help of the following examples.

Examples

(a) Potassium reacts with chlorine to form potassium chloride. Potassium that loses an electron during the reaction is the reducing agent and chlorine that accepts that electron to get reduced is the oxidising agent.

$$2K + Cl_2 \rightarrow 2KCl$$

This is represented as :

$$[K - e^- \rightarrow K^+] \times 2$$
$$[Cl + e^- \rightarrow Cl^-] \times 2$$

$$\mathbf{2K + Cl_2 \rightarrow 2KCl}$$

(b) Gain of electrons by an atom is also shown by the following example.

$$\underset{\text{[ferric chloride]}}{2FeCl_3} + \underset{\text{[hydrogen sulphide]}}{H_2S} \rightarrow \underset{\text{[ferrous chloride]}}{2FeCl_2} + \underset{\text{[hydrogen chloride]}}{2HCl} + \underset{\text{[sulphur]}}{S}$$

$$\underset{\text{[ferric ion]}}{Fe^{3+}} + e^- \rightarrow \underset{\text{[ferrous ion]}}{Fe^{2+}}$$

Reduction is a chemical reaction that involves the addition of hydrogen to a substance or the removal of oxygen from a substance or the gain of electrons from an atom or a group of atoms or from an ion.

(c) *Redox reaction* : A reaction in which oxidation and reduction take place simultaneously is called *redox reaction, i.e.* one substance is oxidised and the other substance is reduced at the same time. In other words loss or gain of electrons takes place at the same time.

Examples :

(a) When chlorine reacts with hydrogen sulphide, chlorine gets reduced to hydrogen chloride and hydrogen sulphide gets oxidised to sulphur.

reduced — *addition of hydrogen* ($Cl_2 \rightarrow 2HCl$)

$$Cl_2 + H_2S \rightarrow 2HCl + S$$

oxidised — *removal of hydrogen* ($H_2S \rightarrow S$)

(b) reduced — *(removal of oxygen)* ($PbO \rightarrow Pb$)

$$PbO + H_2 \rightarrow Pb + H_2O$$

oxidised — *(addition of oxygen)* ($H_2 \rightarrow H_2O$)

In the above reaction, lead monoxide is reduced to lead and hydrogen is oxidised to water. Lead monoxide acts as an oxidising agent and hydrogen acts as a reducing agent.

Note : It has been found that all chemical reactions involving loss or gain of electrons are *redox reactions*. **[Double decomposition reactions are not redox reactions].**

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.
   (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) When an atom loses electron .... takes place.
   (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 called ....
   (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 the sulphur atom gains 2 electrons, the .... of sulphur takes place.

2. Classify the following reactions as combination, decomposition, displacement, precipitation and neutralization. Also balance the equations.
   (a) $CaCO_3(s) \xrightarrow{heat} CaO(s) + CO_2(g)$
   (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) $Ca(s) + O_2(g) \xrightarrow{heat} CaO(s)$

(h) $NaOH + HCl \rightarrow NaCl + H_2O$

(i) $KOH + H_2SO_4 \rightarrow K_2SO_4 + H_2O$

3. Classify the following as oxidation or reduction and complete them.

(a) $Fe^{2+} \rightarrow Fe^{3+}$ (b) $Cl \rightarrow Cl^-$

(c) $Na \rightarrow Na^+$ (d) $Cl^- \rightarrow Cl$

(e) $Al^{3+} \rightarrow Al$ (f) $Hg^{2+} \rightarrow Hg^+$

(g) $O \rightarrow O^{2-}$

4. Define :

(a) precipitation (b) neutralization

(c) catalyst

5. What is redox reaction ? Explain with the help of an example.

6. 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 decomposition reaction

7. Write the missing reactants and products and balance the equations :

(a) $PbO + ... \rightarrow Pb + H_2O$

(b) $NaOH + ... \rightarrow NaCl + ...$

(c) $KClO_3 \xrightarrow{heat} .... + ...$

(d) $.......... + HCl \rightarrow NaCl + H_2O + ...$

8. 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.

9. Give four differences between oxidation and reduction.

10. 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 of hydrogen chloride gas.

11. Give reason :

(a) A person suffering from acidity is advised to take an antacid.

(b) Acidic soil is treated with quick lime.

(c) Wasp sting is treated with vinegar.

C. Electrolysis

3.5 ELECTROLYSIS

Certain chemical reactions take place only under the influence of electric current. Such reactions are called electrochemical reaction or electrolysis.

Electrolysis is the process in which an electric current is passed through the aqueous solution of a compound or its molten state to bring about a chemical change.

The passage of electric current takes place because the compound in liquid state contains free ions and the compound itself dissociates into its constituents due to the chemical change.

The process involves two steps :

- In the first step the ions present in the liquid state of the compound migrate towards the oppositely charged solid conductors (electrodes).
- In the second step these ions turn into neutral atoms due to loss or gain of electrons at the electrodes.

Examples

1. When an electric current is passed through molten (fused) or aqueous solution of sodium chloride, it breaks up into sodium ion (Na^+) and chloride (Cl^-) ion.

$$\underset{\text{(molten or aqueous solution)}}{NaCl} \xrightleftharpoons[\text{current}]{\text{electric}} \underset{\text{(sodium ion)}}{Na^+} + \underset{\text{(chloride ion)}}{Cl^-}$$

The positively charged sodium ion migrates towards the cathode end. At the cathode it gets discharged by accepting an electron and forms the neutral sodium atom.

At cathode :

$$Na^+ + e^- \rightarrow Na$$

On the other hand, the negatively charged chloride ion migrates towards the anode and loses an electron to get discharged and forms a neutral chlorine atom. Two such atoms then combine to form a chlorine molecule, since a chlorine atom cannot exist independently.

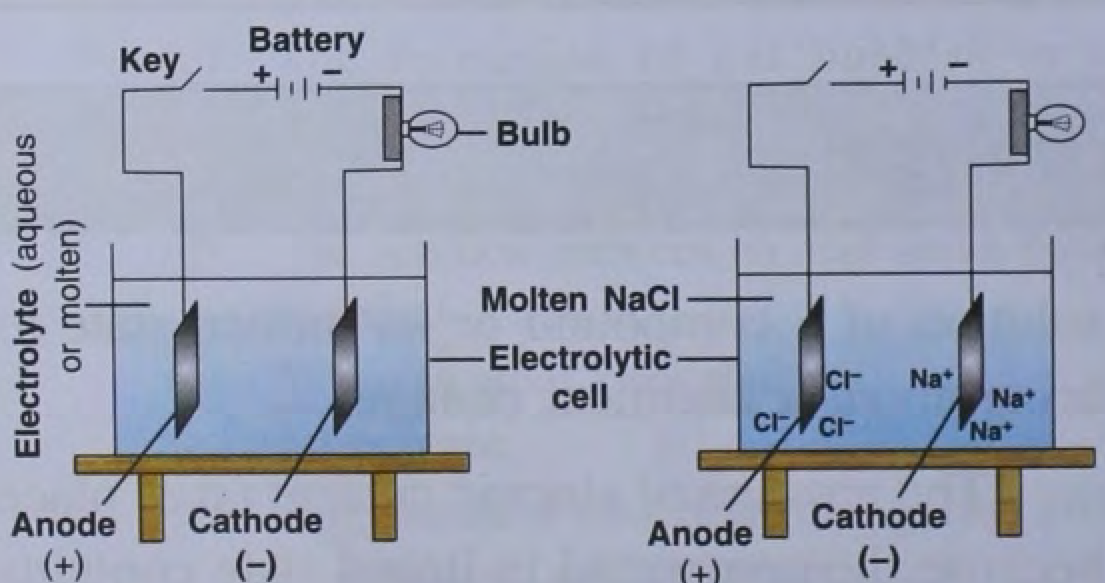


Fig. 3.5 (a) Electrolytic cell
(b) The diagram show movement of cation (Na^+) towards cathode and that of anions (Cl^-) towards anode

At anode :

$$Cl^- - e^- \rightarrow Cl$$
$$Cl + Cl \rightarrow Cl_2$$

Thus sodium is deposited at the cathode and chlorine is evolved at the anode.

2. When an electric current is passed through acidulated water, it decomposes into hydrogen and oxygen gases in the ratio of 2 : 1 by volume. Hydrogen gas is collected at the cathode and oxygen gas is collected at the anode.

The reaction takes place as follows :

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

Water first dissociates into a hydrogen ion (H^+) and a hydroxyl ion (OH^-).

$$\underset{\text{[water]}}{H_2O} \rightleftharpoons \underset{\text{[hydrogen ion]}}{H^+} + \underset{\text{[hydroxyl ion]}}{OH^-}$$

Hydrogen ions (H^+) migrate towards the cathode and accept an electron to form a neutral hydrogen atom. Since a hydrogen atom cannot exist independently two of them combine to form a hydrogen molecule.

At anode :

$$H^+ + e^- \rightarrow H$$
$$H + H \rightarrow H_2$$

Hydroxyl ions (OH^-) being negatively charged, migrate towards the anode. They lose electrons and ultimately form water and oxygen gas.

At anode :

$$OH^- - e^- \rightarrow OH$$
$$4OH \rightarrow 2H_2O + O_2.$$

The volume of hydrogen evolved is double of the volume of oxygen evolved.

3.5.1 Terms associated with electrolysis

The terms associated with electrolytic reactions are :

1. **Electrolytes :** Electrolytes are the aqueous* or the molten compounds that conduct electric current and get decomposed into ions *i.e.*, the compound in liquid form.

Note* : Aqueous compound is the solution of that compound in water.

Examples : Acids, bases & salts are electrolytes.

Electrolytes are of *two* types :

(a) Strong electrolytes : These substances dissociate completely into ions.

Example : Sodium chloride, potassium chloride, lead bromide, sodium hydroxide, potassium hydroxide, hydrochloric acid, nitric acid, sulphuric acid, etc.

(b) Weak electrolytes : These substances dissociate into ions to a lesser extent. Hence their solutions contain both ions and molecules.

Example : Ammonium acetate, ammonium hydroxide, acetic acid, etc.

2. Non-electrolytes : Non-electrolytes are the aqueous or molten compounds that do not decompose into ions and thus do not conduct electric current.

Examples : Sugar, pure water, alcohol, and organic compounds like glycerine, benzene, etc., are non-electrolytes.

3. Electrodes : The two conducting poles through which an electric current enters or leaves an electrolyte are called electrodes. Electrodes are made up of metals or graphite. The two electrodes are *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 : The electrically charged atoms or groups of atoms formed due to the decomposition of an electrolyte during electrolysis are called *ions*. Ions are formed due to the loss or gain of electron(s) by the atoms. The electrical charge on an ion is equal to the valency of the atom from which it is formed. There are two types of ions, *cation* and *anion*.

Anion	*Cation*
1. Anions are negatively charged atoms.	1. Cations are positively charged 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.

Electrolytic cell : It is the container in which electrolysis takes place.

1. How do we detect that the given substance allows passage of electric current ?

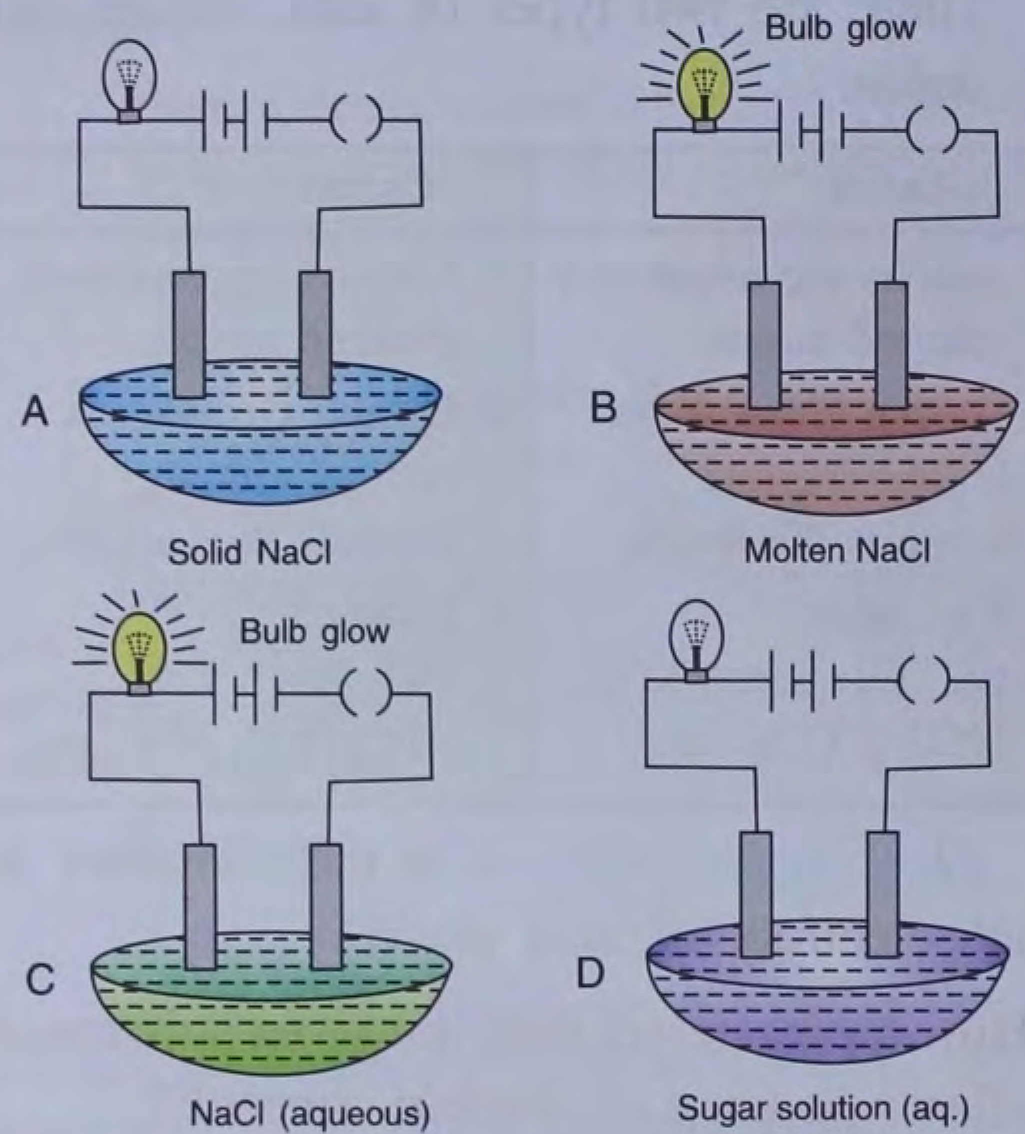
For this the following experiment is done:

Four crucibles A, B, C and D are taken. Crucible A contains solid sodium chloride, B has molten sodium chloride and C contains an aqueous solution of sodium chloride. Crucible D has sugar solution. To each crucible two graphite rods are dipped which act as electrodes. These rods are connected to a battery and a bulb with wires. The circuit is closed.

Observation : The bulb does not glow in crucible A and D but it glows in B and C.

Conclusion : This shows that solid NaCl and sugar solutions are non electrolytes and they do not conduct electricity while sodium chloride in molten and aqueous solution states conducts electricity and acts as electrolyte.

> **Note :** If the electrolyte is strong, the bulb glows brightly and if the electrolyte is weak, e.g. vinegar, the bulb glows dimly.



3.5.2 Applications of electrolysis

The main applications of electrolysis are :

1. ***Electroplating* :** It is the process of plating a metal or an article with another metal.
2. ***Electrorefining* :** It is the process of refining an impure metal to obtain its pure form.
3. ***Electrometallurgy (extraction of metals)* :** It is the process of extracting a metal in its pure form from compounds containing that metal.

> **Do You Know ?**
>
> The English chemist Humphry Davy was one of the first to use electrolysis. He separated sodium, potassium, calcium from their compounds by electrolysis.

3.6 ELECTROPLATING

Electroplating is an electrolytic process in which the deposition of a superior metal is done on the surface of an inferior metal or an article. Thus iron or copper articles can be electroplated with a layer of gold or silver or nickel or chromium.

A metal or an article is electroplated :

(i) to preserve it and to protect it from rust.

(ii) to modify its appearance so as to make it appear more attractive and expensive.

Process of electroplating : For electroplating :

(i) *cathode* is made up of an article or a metal that has to be plated.

(ii) *anode* is made up of a superior metal that is to be plated over the inferior one.

(iii) *electrolyte* contains the metallic ions of a superior metal that has to be electroplated.

On passage of an electric current through the electrolyte, the anode metal ionizes to give its metal ions. These ions enter the electrolyte. Being positively charged, these metal ions are discharged at the cathode, *i.e.* on the article to be plated. Thus a thin coating of the superior metal is formed on the metal article.

3.6.1 Conditions for firm and smooth electroplating

1. The metal surface should be smooth, thoroughly cleaned, and free of grease and oil.
2. The electrolyte should contain appropriate concentration of metal ions to be plated.
3. The temperature of the electrolyte should remain constant throughout the process of electroplating.
4. A low voltage direct current should be passed for a long time.

3.6.2 Electroplating an article with nickel

For electroplating an article with nickel the following implements are required :

Cathode : Made up of an article to be plated.

Anode : Metal plate made up of pure nickel.

Electrolyte : Nickel sulphate solution (containing nickel ions Ni^{2+}).

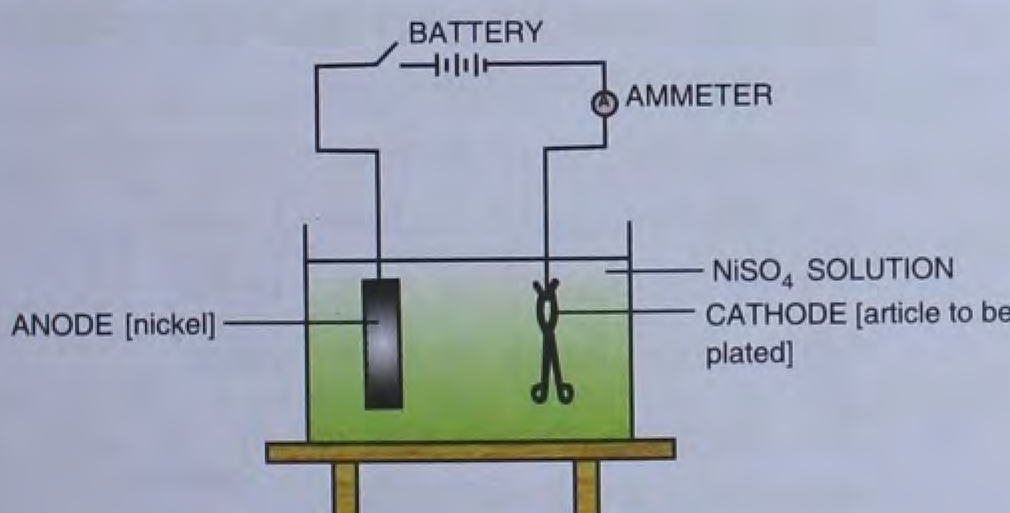


Fig. 3.6 Electroplating of nickel on an article.

When an electric current is passed through the electrolyte, nickel sulphate dissociates into ions, as shown below :

$$NiSO_4 \xrightleftharpoons{\text{electrolysis}} Ni^{2+} + SO_4^{2-}$$

Reaction at cathode : Ni^{2+} being positively charged, it migrates towards the cathode. Ni^{2+} ion takes up two electrons, thus forming a neutral nickel atom. This gets deposited on the cathode and gradually forms a coating over the article to be plated.

$$\underset{\text{(ion)}}{Ni^{2+}} + 2e^- \rightarrow \underset{\text{(atom)}}{Ni}$$

Reaction at anode : At the anode, the nickel atom ionises by losing two electrons and forms an Ni^{2+} ion, which enters the electrolyte.

$$\underset{\text{(atom)}}{Ni} - 2e^- \rightarrow \underset{\text{(ion)}}{Ni^{2+}}$$

The concentration of nickel ions is maintained in the electrolyte because the number of nickel atoms formed at the cathode is equal to the number of nickel ions formed at the anode. The sulphate ions remain in the solution.

[Electrorefining and electrometallurgy are taught in detail in higher classes].

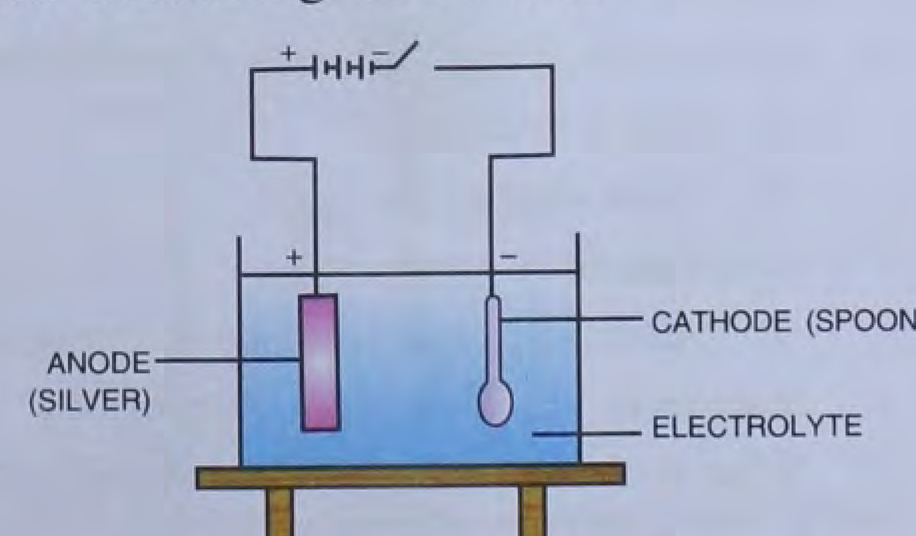
3.6.3 Electroplating a spoon with silver

Following implements are required.

Cathode : Made up of spoon to be plated.

Anode : Metal plate made up of pure silver.

Electrolyte : Potassium argentocyanide solution containing silver ions.



When an electric current is passed through the electrolyte, it dissociates into ions :

$$K[Ag(CN)_2] \rightleftharpoons K^+ + Ag^+ + 2CN^-$$

At cathode : Silver (Ag^+) ions being positively charged migrates towards cathode. It takes up one electron, thus forming neutral silver atom. This gets deposited on the spoon and forms a coating on it.

$$\underset{\text{(ion)}}{Ag^+} + e^- \rightarrow \underset{\text{(atom)}}{Ag}$$

At anode : At the anode the silver atom ionises by losing an electron and forms Ag^+ ions which enters the electrolyte.

$$\underset{\text{(atom)}}{Ag} - e^- \rightarrow \underset{\text{(ion)}}{Ag^+}$$

The concentration of silver ion is thus maintained in the electrolyte. The potassium ions and cyanide ions remain in the solution (electrolyte).

USES OF ELECTROPLATING

1. Cheaper metals are coated with gold or silver to make them appear more attractive as real jewellery.

   Stainless steel cutlery are electroplated with silver, gas stoves commonly used in households are electroplated with chromium.

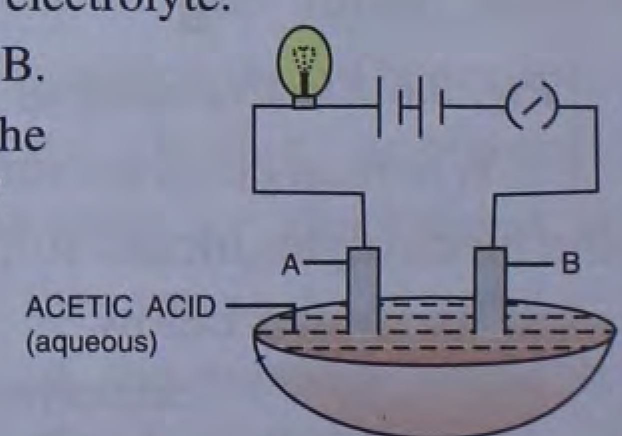
2. Electroplating is also used in the manufacture of printed circuit boards used in radios, T.V., computers, etc.

DO YOU KNOW HOW COLOURED FOILS ARE MADE ?

Coloured foils are made by *anodizing*. In this process the electricity is passed through acid with aluminium as anode, oxygen is formed at anode that reacts with aluminium to form a protective layer of aluminium oxide. These oxide layers are then subjected to dyeing to form coloured foils.

EXERCISE – III

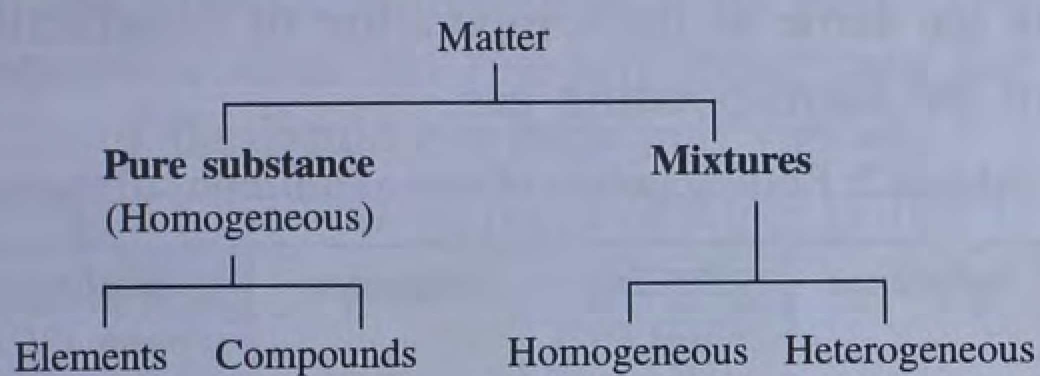
1. Define the following terms :
   (a) Electrolysis. (b) Electrolyte
   (c) Cathode (d) Anode
   (e) Cation (f) Anion
2. Differentiate between :
   (a) Cation and anion
   (b) Cathode and anode
   (c) Electrolyte and non-electrolyte
3. Give two examples for each of the following :
   (a) Strong electrolyte
   (b) Weak electrolyte
   (c) Non-electrolyte
   (d) Substances used to make electrodes.
4. State three applications of electrolysis.
5. What is electroplating ? In the electroplating of an article by nickel name the cathode, anode and metal ions present in the electrolyte.
6. Give equations for reactions at the cathode and the anode for the following :
   (a) electrolysis of molten sodium chloride
   (b) electrolysis of acidulated water
   (c) electroplating by nickel metal
   (d) electroplating by silver metal.
7. The diagram shows electrolysis of acetic acid which is a weak electrolyte.
   (a) Label A and B.
   (b) How will the bulb glow ?



D. Pure Substances — Purification

INTRODUCTION

There are different kinds of substances in this world. They are broadly classified into two types which are again subdivided.



Pure substances are always homogeneous. They are represented by a definite symbol or formula while mixtures can be homogeneous like solutions or heterogeneous like suspensions. They do not have any definite symbol or formula.

3.7 CHARACTERISTICS OF PURE SUBSTANCES

All pure substances, whether elements or compounds, have certain characteristic physical properties. The melting point and the boiling point are two such properties. *For example*, the **melting** point of pure ice is 0°C at normal atmospheric pressure (76 cm of mercury) and the **boiling** point of pure water is 100°C. Interestingly, the melting point of a solid is the same as the freezing point of the corresponding liquid.

The melting point and the boiling point of a pure substance are fixed at a particular pressure.

3.7.1 Melting point

The melting point of a solid is the temperature at which it starts changing into a liquid at unit atmospheric pressure. The temperature remains constant till all the solid has melted.

Table 3.1 : Melting points of some common substances

Substance	*Melting point* (°C)	*Substance*	*Melting point* (°C)
Gold	1064	Calcium	839
Carbon	3550	Aluminium	660
Iron	1535	Sodium hydroxide	318
Silicon	1410	Ice	0
Silver	961	Nitric acid	– 42

Experiment to determine the melting point of ice

Put a few ice cubes in a glass funnel. Place a thermometer in the middle of the funnel with the help of a stand, its bulb resting inside ice as shown in figure 3.7. Keep a beaker under the funnel. Note down the fall in the temperature of the thermometer. When it reaches 0°C drops of water begin to fall into

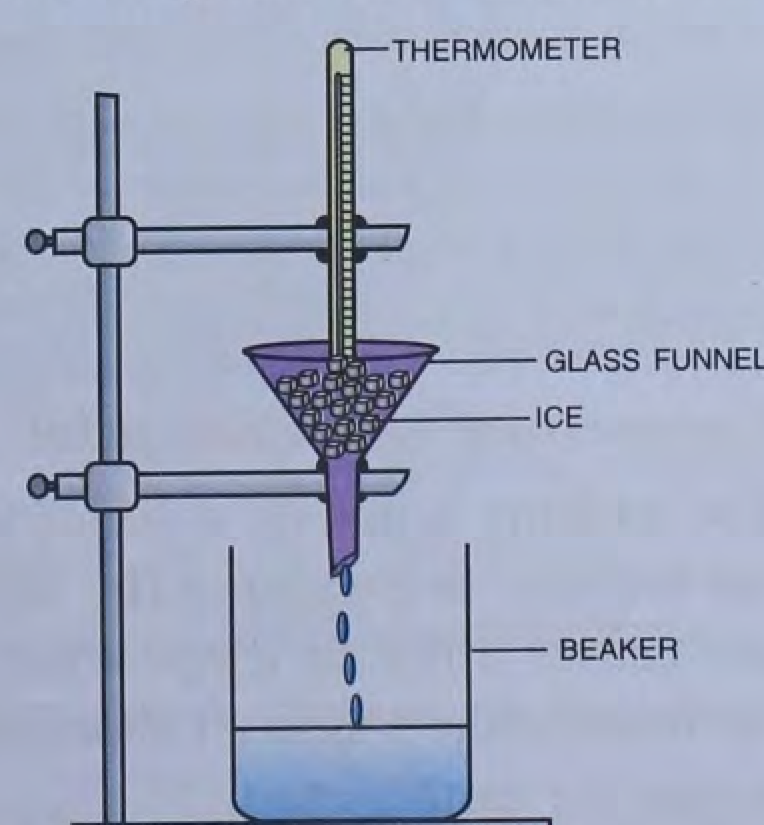


Fig. 3.7 Determining the melting point of ice.

the beaker. This temperature remains constant until all the ice melts. Hence **0°C is the melting point** of ice.

Effect of impurity on melting point

The presence of impurity causes a decrease in the melting point of a substance. This is why the melting point of freezing mixture* is about –15°C, which is lower in comparison to that of pure ice (0°C). Note that the melting point of ice formed from tap water is not 0°C, since tap water contains dissolved minerals as impurity.

Measuring its melting point is a good way of checking the degree of the purity of a substance.

Activity 1

To check the degree of the purity of a substance with the help of its melting point.

Take four cups. Pour sugar solution into the first cup and salt solution into the second one. Fill the third cup with tap water and the fourth with distilled water. Place all the four cups in a freezer for some time. Then take the cups out of the freezer, crush the frozen material and note the temperature at which the material in each of the cups melts.

Are the melting points same in each case? No, the ice formed from distilled water melts at 0°C, while the other samples have respectively lower melting points.

Effect of pressure on melting point

The melting point of a solid decreases with an increase in pressure. The increase in pressure causes a rise in temperature. This is why ice melts under a skater's shoes and rocks melt under the earth's crust.

* *A mixture of ice and salt is known as freezing mixture.*

3.7.2 Boiling point

The temperature at which a liquid begins to boil and change into vapour state under normal atmospheric pressure is called its *boiling point*. The temperature remains constant till the whole of the liquid changes into vapour. The boiling point of a pure liquid is the same as the temperature of liquefaction of the corresponding gas.

Table 3.2: Boiling points of some common substances

Substance	*Boiling point (°C)*	*Substance*	*Boiling point (°C)*
Nitric acid	83	Oxygen	– 183
Sulphuric acid	280	Mercury	356
Hydrogen	– 252	Iodine	184
Nitrogen	– 195	Chlorine	– 34

Experiment to determine the boiling point of pure water

Take distilled water in a round bottom flask. Fit a thermometer in the mouth of the flask with the help of a cork in such a way that its bulb remains just above the surface of

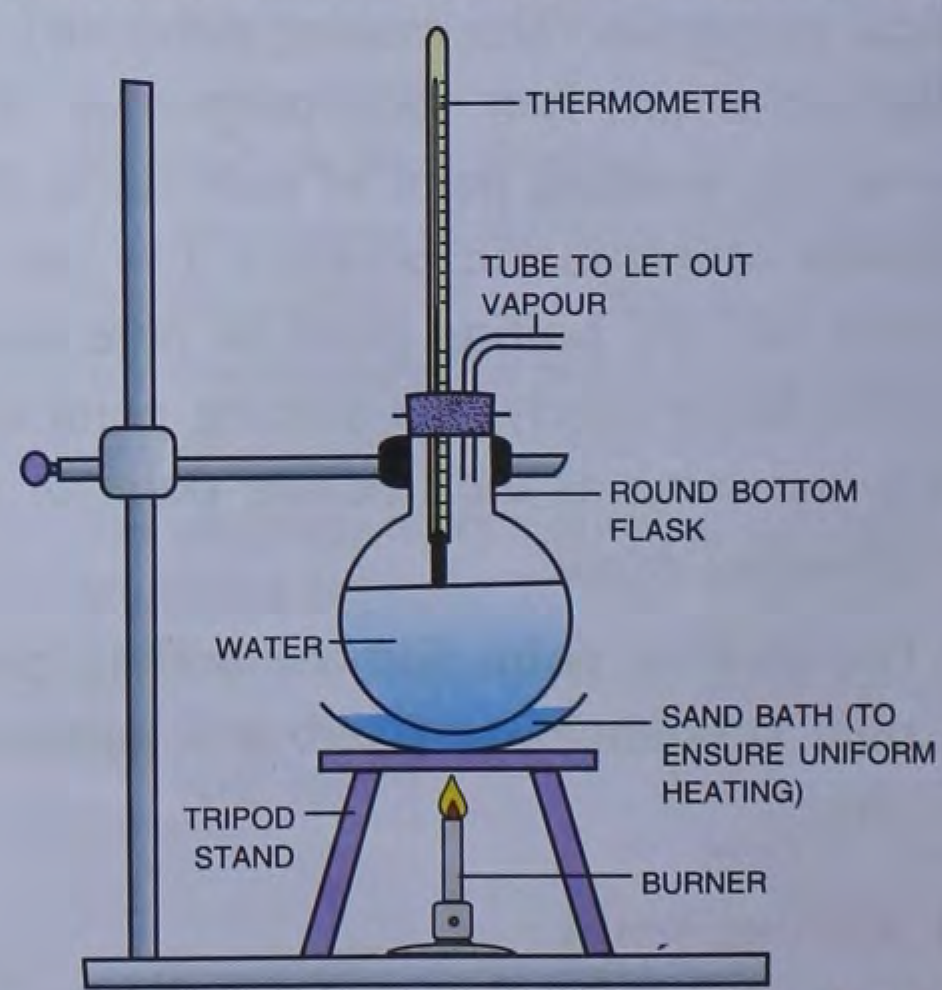


Fig. 3.8 Determining the boiling point of a liquid.

the liquid. Fit a tube in the mouth of the flask with the help of a cork. Place the flask on a tripod stand over the sand bath (to ensure uniform heating) with the help of a stand.

Now heat the liquid. As the liquid is slowly heated you will observe a rise in the temperature of the thermometer. But when it reaches 100°C, water begins to boil. The temperature does not rise any more until the whole of the liquid converts into vapour.

Hence the boiling point of water is 100°C under normal pressure.

Effect of impurity on boiling point

The presence of impurity increases the boiling point of a liquid. Since tap water contains dissolved minerals, it boils at more than 100°C.

Thus we can say that the boiling point too helps to check the degree of the purity of a substance.

Activity 2

To check the degree of the purity of a substance with the help of its boiling point.

Heat tap water, distilled water, sugar solution and salt solution separately and check the temperature at which each of them boils.

Distilled water boils at 100°C, but the other samples boil each at a higher temperature under the same pressure.

Effect of pressure on boiling point

The boiling point of a liquid decreases with a decrease in pressure and increases with an increase in pressure. This is why pure water does not always boil at 100°C, given that atmospheric pressure does not always remain the same at all places.

In the hilly regions the density of air is somewhat less than normal. Therefore atmospheric pressure is also low. This decreases the boiling point of water at higher altitudes.

3.8 PURIFICATION OF SUBSTANCES

The method employed to purify a substance depends on the properties and the uses of that substance. One of the methods used for obtaining a very pure substance is *distillation*.

Distillation

Distillation is a process to convert a liquid into its vapour state by heating it and then condensation of the vapour back into the liquid state by cooling it. Distillation is used to purify a liquid that contains soluble impurities. The liquid is first heated and then vapours from the boiling liquid are collected as shown in Fig 3.9. The vapours are then condensed by cooling it. The liquid thus obtained is highly pure. By this method impure water is distilled to obtain pure water in the laboratory.

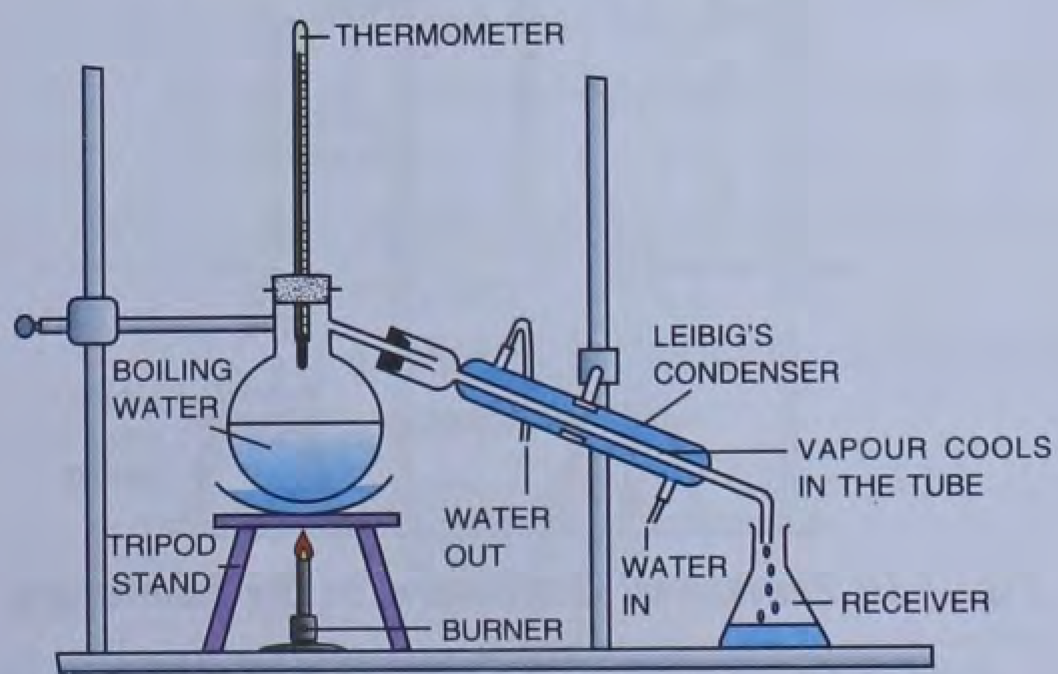


Fig. 3.9 Preparing distilled water in the laboratory.

Note : Leibig's condenser is used to condense the vapour into liquid.

Fractional distillation

Fractional distillation is a process employed to separate the different components (fractions) of a miscible liquid, which is usually a homogenous mixture. The separation is possible due to the difference in the boiling points of the respective fractions of the liquid mixture. However, for fractional distillation to be possible, certain conditions must be met.

1. Any desired component of the liquid mixture must not itself break up into some other substance before it boils.
2. There should be enough difference (about 30°C) between the boiling points of the components, otherwise the vapours of two or more liquids could get mixed up.*

When this mixture of liquids is heated the more volatile (with a lower boiling point) component boils first. The vapour rises up the fractionating column and passes into the tube that is cooled by the condenser. The condensed vapour is collected in the collecting flask.

Since liquids evaporate at all temperatures, even the vapours of the less volatile liquid rise into the fractionating column. But they condense in the column itself and fall back into the flask. The method is widely used in industries to purify and separate miscible liquids.

Examples :

(a) To separate petrol, diesel, kerosene, *etc*., from crude oil.

(b) To separate nitrogen and oxygen from liquefied air.

(c) To separate alcohol and water from their mixture.

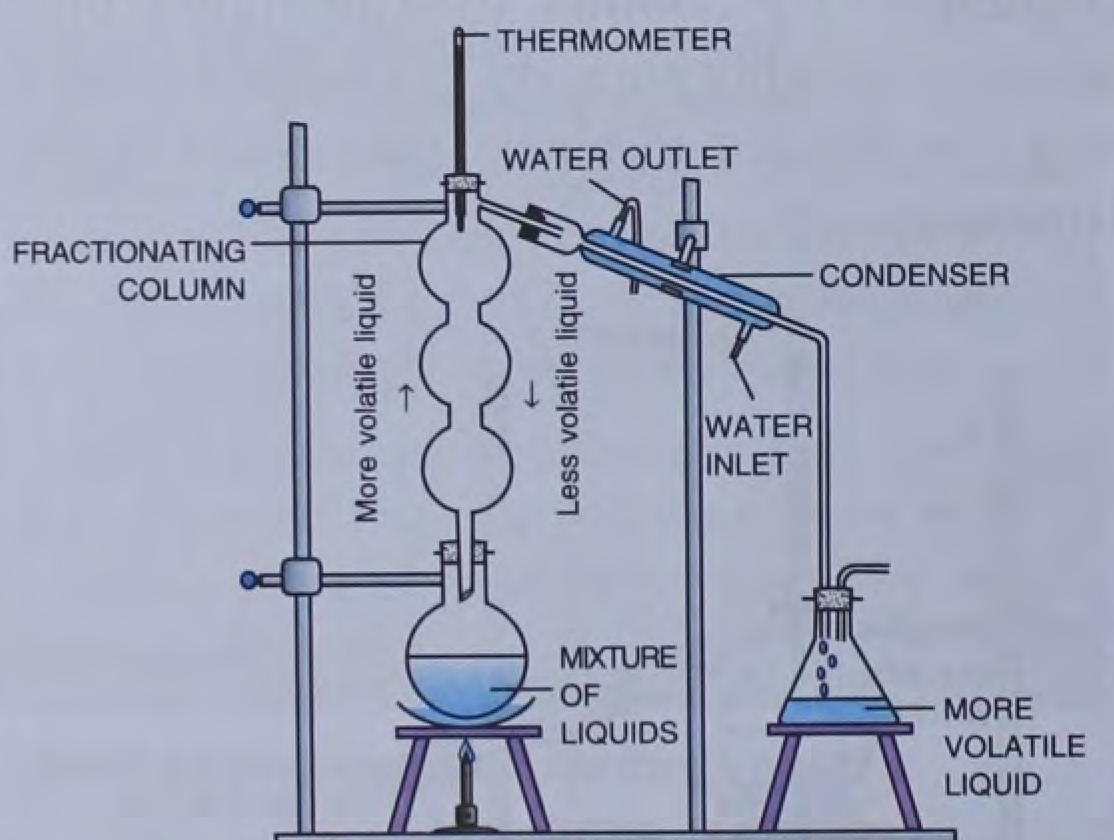


Fig. 3.10 Fractional distillation in the laboratory.

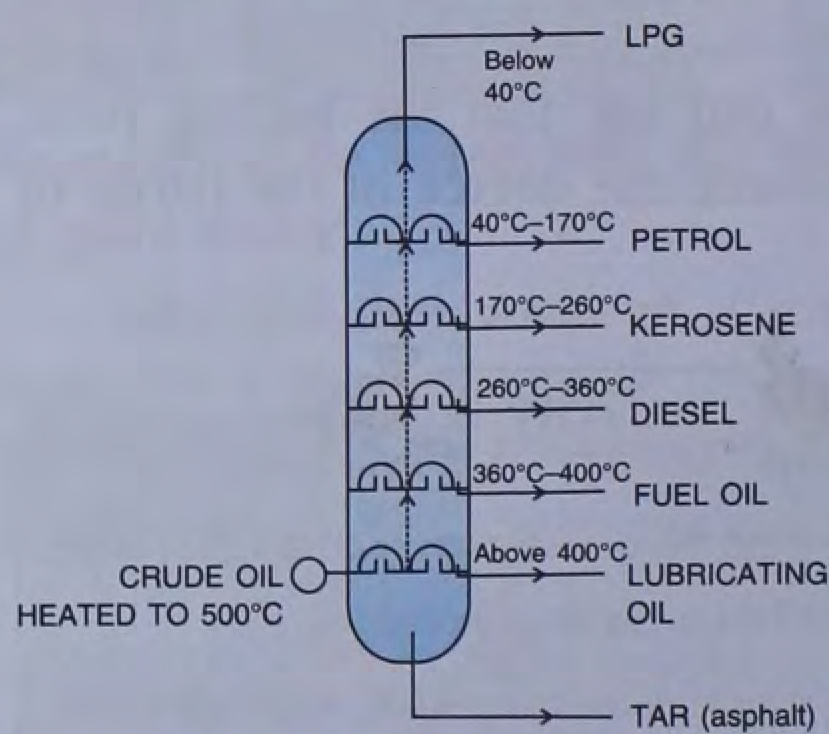


Fig. 3.11 Fractional distillation in the petroleum industry.

	Mixture of miscible liquids			Component that distils over
1.	Ethyl alcohol (b.p. 78°C)	+	water (b.p. 100°C)	Ethyl alcohol
2.	Methyl alcohol (b.p. 64·5°C)	+	ethyl alcohol (b.p. 78°C)	Methyl alcohol
3.	Ethyl alcohol (b.p. 78°C)	+	chloroform (b.p. 61°C)	Chloroform
4.	Acetone (b.p. 56·5°C)	+	water (b.p. 100°C)	Acetone
5.	Acetone (b.p. 56·5°C)	+	ethyl alcohol (b.p. 78°C)	Acetone

* *If the difference in the boiling points of the desired liquid fractions is not enough to prevent the mixing of their vapours, the fractionating column is fitted with the distilling flask. The column then allows the vapours of the less volatile liquid to condense and fall back into the flask.*

EXERCISE – IV

1. Fill in the blanks :
   (a) The boiling point of a pure liquid .... with the decrease in pressure.
   (b) When a solid changes to liquid on melting .... remains constant.
   (c) A mixture of ice and salt is known as ....
   (d) The melting point of pure ice is ....
   (e) An increase in pressure or impurity causes a .... in the melting point of a substance.
   (f) The boiling point of a pure liquid at a particular pressure remains ....
   (g) Impure water is .... to obtain pure water in the laboratory.
2. Define the following :
   (a) Melting point (b) Boiling point
   (c) Distillation
3. Give reasons :
   (a) Why does ice under a skater's shoes melt ?
   (b) Why is the melting point of freezing mixture about –15°C ?
   (c) Why does tap water boil at more than 100°C ?
   (d) Why is a mixture of alcohol and water separated by fractional distillation ?
4. Which is a good way of checking the degree of impurity of a substance ?
5. Define fractional distillation. How is it useful in industries ?
6. What are the effects of pressure and impurity on (a) melting point (b) boiling point of a substance ?
7. Draw a neat labelled diagram to show how to determine the melting point of a solid other than ice.
8. **Multiple Choice Questions**
   (a) The substance formed as a result of chemical reaction is called
      (i) catalyst (ii) reactant
      (iii) product (iv) intermediate
   (b) A chemical equation needs to be balanced because
      (i) atoms can neither be created nor be destroyed
      (ii) energy can neither be created nor be destroyed.
      (iii) electrons are neither lost nor gained during electrolysis.
      (iv) none of the above.
   (c) During electroplating the article to be plated is made
      (i) cathode (ii) anode
      (iii) electrolyte (iv) anode mud
   (d) For silver plating the electrolyte taken is
      (i) silver nitrate
      (ii) sodium chloride
      (iii) silver oxide
      (iv) potassium argento cyanide
   (e) A strong electrolyte is
      (i) calcium hydroxide
      (ii) sugar solution
      (iii) copper metal
      (iv) molten sodium chloride
   (f) Under normal conditions 0°C and 100°C are
      (i) melting point of ice and boiling point of water
      (ii) boiling point of water and freezing point of ice
      (iii) melting point of water and boiling point of water
      (iv) all of the above.
   (g) If the electrolyte is strong the bulb glows
      (i) dimly (ii) brightly
      (iii) does not glow

(h) Red litmus changes into blue in
(i) a neutral solution
(ii) an acidic solution
(iii) an alkaline solution
(iv) none of the above

(i) A person suffering from acidity is adviced to take
(i) an acid (ii) water
(iii) salt solution (iv) a basic solution

(j) When an iron nail is kept in copper sulphate solution, the colour of the solution changes from
(i) green to blue
(ii) blue to green
(iii) blue to pink
(iv) green to colourless

(k) Oxidation is
(i) loss of electron (ii) gain of electron
(iii) both (iv) none

(l) Action of heat on potassium nitrate produces
(i) potassium nitrite and oxygen
(ii) potassium oxide and nitrogen dioxide
(iii) potassium and nitrogen dioxide
(iv) potassium, nitrogen and oxygen

E. Chemical Equations and Calculation Based on Them

3.9 CHEMICAL EQUATIONS

'A chemical equation is the symbolic representation of a chemical reaction with the help of the symbols and the formulae of the substances involved in that 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***.

$$\text{Reactants} \xrightarrow[\text{Reaction}]{} \text{Products}$$

Examples

(a) Burning of coal in air is a chemical reaction in which a new substance, carbon dioxide, is formed*. This reaction can be represented by a chemical equation, using formulae and symbols, as shown below.

$$\underset{\text{[carbon]}}{C} + \underset{\text{[oxygen]}}{O_2} \xrightarrow{\text{heat}} \underset{\text{[carbon dioxide]}}{CO_2}$$

* *Since coal is impure carbon, the burnt impurities are left behind as ash.*

Steps involved in writing a chemical equation

1. The symbols or the formulae of the reactants are written on the left hand side and those of the products are written on the right hand side of the arrow.
2. The reactants and the products are represented in their **molecular forms**.
   [**Note :** atoms are usually neither stable nor capable of separate existence].
3. The number of atoms of each element must be equal on the reactant and the product sides.

(b) Now consider the reaction of hydrogen with chlorine, the reactants, to form hydrogen chloride, the product.

$$\underset{\text{[hydrogen]}}{H_2} + \underset{\text{[chlorine]}}{Cl_2} \rightarrow \underset{\text{[hydrogen chloride]}}{HCl}$$

In the above reaction the number of hydrogen (or of chlorine) atoms on the left hand side does not equal the number of their atoms on the right hand side. Such an equation is an unbalanced one and it is called *skeletal equation*.

Need to balance a chemical equation

A chemical equation needs to be balanced so as to make the number of atoms on the reactant side equal to the number of atoms on the product side. This is necessary because a chemical reaction is just a rearrangement of atoms. The atoms themselves are not created or destroyed during a chemical reaction (law of conservation of matter). Therefore, the balanced chemical equation for the reaction of hydrogen with chlorine is written as :

$$H_2 + Cl_2 \rightarrow 2HCl$$

> A balanced chemical equation is 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.

3.9.1 How to balance a chemical equation?

A chemical equation is balanced by taking the following steps :

1. Count the number of atoms of each element on either side.
2. Multiply the reactants and the products with appropriate numerals to make the number of each kind of atoms the same on both sides.

Example : Reaction of zinc with hydrochloric acid gives zinc chloride and hydrogen.

The skeletal equation is :

Symbol		*Formula*		*Formula*		*Formula*
Zn	+	HCl	$\rightarrow$	$ZnCl_2$	+	H_2
element		compound		compound		element

Steps for balancing the equation :

Step 1 : Count the number of atoms of each element on either side.

Reactant side Zinc (Zn) = 1
Hydrogen (H) = 1
Chlorine (Cl) = 1

Product side Zinc (Zn) = 1
Chlorine (Cl) = 2
Hydrogen (H) = 2

Both hydrogen and chlorine have an extra atom each present on the product side.

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

$$Zn + \boxed{2}\,HCl \rightarrow ZnCl_2 + H_2$$

Now we have :

Zn = 1, H = 2, Cl = 2 → Zn = 1, H = 2, Cl = 2.

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

3.9.2 Information provided by a balanced chemical equation

A balanced chemical equation provides a lot of information :

1. It tells us about the substances taking part in a chemical reaction (reactants) and about the substances that are formed (products) as a result of it.
2. It shows both the number of molecules and the number of atoms involved in the reaction.
3. It shows the chemical composition of the respective molecules.
4. It tells about the relative molecular masses of the different substances, *i.e.* the actual amounts of the reactants and the products taking part in the chemical reaction.
5. It also proves the law of the conservation of mass, *i.e.* the total mass of the substances on the two sides of the equation is the same.
6. It makes the study of chemistry universally standardized.

Information provided by a chemical equation.

$$\underset{48}{2Mg} + \underset{32}{O_2} \xrightarrow{heat} \underset{80}{2MgO}$$

This chemical equation indicates that :

1. Magnesium and oxygen react to produce magnesium oxide.
2. Two monoatomic molecules of magnesium (element) react with one diatomic molecule of oxygen (element) to give two molecules of magnesium oxide (compound) containing one atom of magnesium and one atom of oxygen per molecule (MgO).
3. 48 g of magnesium combines with 32 g of oxygen to form 80 g of magnesium oxide.

3.10 SIMPLE CALCULATIONS BASED ON CHEMICAL EQUATIONS

To carry out calculations based on chemical equations, proceed in the following manner.

1. Balance the equation.
2. Find the molecular weight of each substance involved by adding up the weights of the atoms in the molecule.
3. The product of the number of molecules and the molecular weight of each substance is written below its formula. Molecular weight is expressed in a.m.u. as a molecule is very small. But for a collection of molecules it can be expressed in grams, kilograms or tonnes.

Relative atomic mass : The number of times an atom of an element is heavier than 1/12th of the mass of a carbon-12 atom is called relative atomic mass.

Relative molecular mass : The number of times one molecule of an element or a compound is heavier than 1/12th of the mass a carbon-12 atom is called relative molecular mass.

Calculation of molecular mass

1. **Calcium oxide (CaO) :**

   At. mass of Ca = 40, O = 16

   ∴ Mol. mass = 40 + 16 = 56 amu

2. **Sodium chloride (NaCl) :**

   At. mass of Na = 23, Cl = 35·5

   ∴ Mol. mass = 23 + 35·5 = 58·5 amu

3. **Sulphuric acid (H_2SO_4) :**

   At. mass of H = 1, S = 32, O = 16

   ∴ Mol. mass = 2 × 1 + 32 + 4 × 16
   = 2 + 32 + 64 = 98 amu

Solved problems :

Example 1 : Calculate the amount of magnesium oxide formed when 6 g of magnesium burns completely with oxygen [Atomic mass of Mg = 24 and of O = 16].

Solution : The balanced chemical equation for this reaction is :

$2Mg$ +	O_2	→ $2MgO$
2 × 24	2 × 16	2(24 + 16)
48 amu	32 amu	80 amu
or	or	or
48 g	32 g	80 g

∵ 48 g of Mg forms 80 g of MgO,

∴ 6 g of Mg forms $\frac{80}{48} \times 6 =$ **10 g** of magnesium oxide

Example 2 : Calculate the respective weights of calcium oxide and carbon dioxide formed when 40 g of limestone ($CaCO_3$) is heated strongly.

Solution : The balanced chemical equation for the reaction is written as :

$CaCO_3$ $\xrightarrow{heat}$	CaO +	CO_2
[40 + 12 + (3 × 16)]	[40 +16]	[12 + (2 × 16)]
100 amu	56 amu	44 amu
or	or	or
100 g	56 g	44 g

∵ 100 g of $CaCO_3$ produces 56 g of CaO,

∴ 40 g of $CaCO_3$ produces $\frac{56 \times 40}{100}$ g of CaO

= **22·4 g of CaO**

$\because$ 100 g of $CaCO_3$ produces 44 g of CO_2,

$\therefore$ 40 g of $CaCO_3$ produces $\frac{44 \times 40}{100}$ g of CO_2

= **17·6 g of CO_2**

Example 3 : Calculate the weight of potassium nitrite formed due to the thermal decomposition of 15·15 g of potassium nitrate [Atomic mass of K = 39, N = 14 and O = 16].

Solution : The balanced equation for the reaction is :

$2KNO_3$	$\rightarrow$	$2KNO_2$	+	O_2
2[39 + 14 + (3 × 16)]		2[39 + 14 + (2 × 16)]		[2 × 16]
202 amu		170 amu		32 amu
or		or		or
202 g		170 g		32 g

$\because$ 202 g of KNO_3 produces 170 g of KNO_2

$\therefore$ 15·15 g of KNO_3 produces $\frac{170 \times 15{\cdot}15}{202}$ g of KNO_2

= **12·75 g of KNO_2**

Example 4 : 100 kg of limestone is so strongly heated that all the carbon dioxide formed is driven off. What is the mass of calcium oxide left as the residue ? (Relative atomic masses : Ca = 40, C = 12, O = 16).

Solution : The balanced chemical equation for the reaction can be expressed as follows.

$CaCO_3$	$\rightarrow$	CaO	+	CO_2	
40 + 12 + 3 × 16		40 + 16		12 + 2 × 16	
= 100		= 56		= 44	parts by mass
100 kg		56 kg		44 kg	

$\therefore$ the mass of calcium oxide left = **56 kg**

Example 5 : What will be the increase in mass if 48 g of magnesium is heated in oxygen ? (Rel. At. masses : Mg = 24·0 = 16)

Solution : The balanced chemical equation for the reaction is

$2Mg$	+	O_2	$\rightarrow$	$2MgO$
2 × 24		2 × 16		2(24 + 16)
= 48		= 32		= 80

From the equation, it is evident that on reaction with oxygen, 48 g of Mg will form 80 g of MgO.

Thus, the increase in the mass for 48 g of Mg is

80 g – 48 g = 32 g.

Example 6 : What will be the loss in the mass when 42 g of magnesium carbonate is strongly heated to give magnesium oxide ? (Relative atomic masses : Mg = 24, C = 12, O = 16).

Solution : The balanced chemical equation for the reaction can be expressed as follows.

$MgCO_3$	$\rightarrow$	MgO	+	CO_2
24 + 12 + (3 × 16)		24 + 16		12 + (2 × 16)
= 84		= 40		= 44

When strongly heated, 84 g of $MgCO_3$ will lose 44 g of CO_2.

$\therefore$ 42 g of $MgCO_3$ will lose

$\left(\frac{\cancel{44}^{22}}{\cancel{84}_{2}} \times \cancel{42}\right)$ g = 22 g of CO_2.

So, the loss in the mass = **22 g**

EXERCISE – V

1. Fill in the blanks :
   (a) The substances that are formed as a result of the reaction are called ....
   (b) The symbols or the formulae of the reactants are written on the .... hand side of the arrow.
   (c) Molecular weight has .... unit.
   (d) An unbalanced equation is called .... equation.
   (e) Magnesium oxide is a .... of magnesium and oxygen.
2. Define the following :
   (a) Reactants
   (b) Products
   (c) Skeletal equation
   (d) Chemical equation
3. What is a balanced chemical equation ? How is a chemical equation balanced ? Explain with an example.

4. Write the steps involved in writing a chemical equation.
5. Why does a chemical equation need to be balanced ?
6. What information is conveyed from the following chemical equation ?

$$2Mg + O_2 \rightarrow 2MgO$$

7. Calculate the molecular mass of
   (a) $CaCO_3$ (b) $Mg(OH)_2$
   (c) HNO_3 (d) $AlCl_3$
   [At mass H = 1, N = 14, O = 16, Cl = 35·5, Mg = 24, At = 27, Ca = 40].
8. Solve the following problems :
   (a) What will be the mass of ferrous sulphide formed when 20 g of pure sulphur reacts with iron ?
   [Atomic mass of Fe = 56 and S = 32]
   (b) Calculate the mass of oxygen used in the formation of water from 10·2 g of hydrogen.
   [Atomic mass of H = 1 and O = 16]
   (c) Calculate the mass of calcium carbonate which decomposes to produce :
   (1) 50 g of calcium oxide and
   (2) 5 g of carbon dioxide.
   [At. mass of Ca = 40, C = 12, O = 16]
   (d) Calculate the mass of magnesium oxide when 24 g of magnesium reacts with 32 g of oxygen. Also find the weight of oxygen left unused after the reaction.
   [At wt of Mg = 24, Ca = 40]

RECAPITULATION

- A physical change is one in which a substance changes temporarily in some of its physical properties such as state, shape, size and appearance without any change in its chemical composition.
- A chemical change is a permanent change in which the original substance gives rise to one or more new substances with entirely different compositions and properties.
- 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 exothermic reaction.
- The chemical reaction, in which heat is absorbed, is called endothermic reaction.
- A catalyst is a substance which alters the rate of a reaction but itself it does not change.
- 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 more active element.
- A chemical reaction, in which two compounds react in their aqueous solutions by exchanging their radicals, is called double decomposition reaction.
- Precipitation is a double decomposition reaction in which one of the products is an insoluble substance called precipitate.
- The reaction of an alkali with an acid or vice versa is called neutralisation reaction.

- Oxidation is a process which involves loss of electrons from an atom or a group of atoms.
- Reduction is a process which involves gain of electrons by an atom or a group of atoms.
- A redox reaction is one in which oxidation and reduction take place simultaneously.

- The process due to which a compound in fused state or in aqueous solution state breaks up into ions, by passage of electric current is known as electrolysis.
- Cations and anions are the two types of ions. Cations are positively charged and anions are negatively charged.
- The compound which allows passage of electric current in liquid state is called an electrolyte. There are strong and weak electrolytes.
- The compounds which do not allow passage of electric current in fused state or solution form are called non-electrolytes.
- Electrodes are the two poles by which the electrolytic cell is attached to an electric circuit. The current enters the electrolyte through the positively charged electrode called anode and leaves the electrolyte through the negatively charged electrode called cathode.
- The applications of electrolysis are electroplating, electrorefining and electrometallurgy.

- Substances are of two types *pure substances* and *mixtures.*
- The melting points and boiling points are the two characteristic properties, which help to recognise substances in their pure form.
- The process of separation of two miscible liquids, by the process of distillation making use of the difference in their boiling points, is called fractional distillation.

- A chemical equation is the symbolic representation of a chemical reaction with the help of the symbols and formulae.
- A chemical equation needs to be balanced because atoms can neither be created nor be destroyed.
- A balanced chemical equation provides many information about a chemical reaction.