CHAPTER 4
TYPES OF CHEMICAL REACTIONS

NOTES

1. Chemical reaction
2. Types of chemical reactions
3. Rancidity
4. Catalysts

Chemical Reaction:

- Whenever a chemical change occurs we say that a chemical reaction has taken place. It is the process that leads to a chemical change.

  e.g. digestion of food, burning of paper, rusting of iron, etc.
- The substances originally present and taking part in the reactions are called **reactants** and the new substances formed are called **products.**

  e.g. Paper + Air = Carbon dioxide + Water vapour

  In the above example, paper and air (oxygen) are the reactants and the ash, carbon dioxide and water vapour are the products.
- In chemical reactions the bonds present in the reactants are broken and new bonds are formed to produce new substances or products.

Chemical Equation:

The symbolic representation of an actual chemical reaction is with the help of symbols and formulae of respective reactants and products.

e.g. Magnesium is burnt into air to form Magnesium oxide can be represented as:

$$2\,Mg + O_2 \longrightarrow 2\,MgO$$

Magnesium + Oxygen → Magnesium oxide

(Reactants) (Product)

Types of Chemical Reaction:

(a) Combination Reaction: A reaction in which a single product is formed from two or more reactants.

e.g. $2Mg + O_2 \rightarrow 2MgO$

i) Combination reaction can take place between two elements. Example, Carbon or coke burns in air to produce carbon dioxide.

$C + O_2 \rightarrow CO_2$

ii) Combination reaction also takes place between two compounds to form single product.

e.g. $CaO\ (S) + H_2O\ (l) \rightarrow Ca(OH)_2\ (aq)$

Calcium oxide (quick lime) → Calcium hydroxide (slaked lime)

iii) Combination reaction also takes place between an element and a compound to form a new compound

e.g. $2NO + O_2 \rightarrow 2NO_2$

Nitric oxide (colourless) → Nitrogen dioxide (brown)

(b) Decomposition Reaction: The reaction in which a compound is broken down into two or more simpler substances.

Decomposition reaction can be of three types:

(i) Thermal decomposition – Decomposition reaction carried out by application of heat.

Example:

$CaCO_3$ (s) $\xrightarrow{heat}$ CaO (s) + CO_2(g)
Calcium carbonate → Calcium oxide (quicklime)

$2Pb(NO_3)_2$ (s) $\xrightarrow{heat}$ 2 PbO (s) +4 NO_2 (g) + O_2 (g)
Lead nitrate (colourless) → Lead monoxide (yellow), brown fumes

$FeSO_4.7H_2O$ (s) $\xrightarrow{heat}$ $FeSO_4$ (s) + 7 H_2O(g)
Ferrous sulphate (green) crystal → Ferrous sulphate (Anhydrous)

2 $FeSO_4$ (s) $\xrightarrow{heat}$ Fe_2O_3 (s) +SO_2 (g) +SO_3(g)
Iron (III) oxide (reddish brown)

(ii) Electrolytic decomposition: Decomposition reaction carried out by application of electric current.

Example:

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

2NaCl (fused) $\xrightarrow{\text{electric current}}$ 2Na (s) + Cl_2 (g)

2 Al_2O_3 (fused) $\xrightarrow{\text{electric current}}$ 4 Al + 3 O_2

During electrolysis of water:

(a) Volume of the gas collected over the cathode (hydrogen) is double that of the gas over the anode (oxygen).

(b) Hydrogen burns with a pop sound and oxygen produces flame when a glowing splinter is introduced in it.

(c) Photochemical decomposition: Decomposition reaction which takes place in presence of light.

Example:

2AgCl (s) $\xrightarrow{sunlight}$ 2 Ag (s) + Cl_2
Silver chloride

2AgBr (s) $\xrightarrow{sunlight}$ 2Ag (g) + Br_2
Silver bromide

Black and white photography is based upon such reactions. The photographic film is coated with these salts.

(d) Exothermic Reaction: Reaction which takes place with the evolution of energy in any form-heat, light, electrical, mechanical etc.

Example:

CH_4 + 2 O_2 → CO_2 + 2 H_2O + Energy
Methane

- Respiration is also an exothermic process.

(e) Endothermic Reaction: Reaction in which energy is absorbed in any form.

- Photosynthesis is also an endothermic reaction.

(f) **Displacement Reaction:** Reaction in which a more reactive element displaces another less reactive element from its solution.

$CuSO_4$ (aq) + Fe (s) → $FeSO_4$ (aq) + Cu (s)

Copper sulphate (Blue) — Fe (brown) — Ferrous sulphate (light green)

This is because iron displaces copper from copper sulphate solution.

Other examples are: Zinc displaces Copper from Copper sulphate solution.

$CuSO_4$ (aq) + Zn (s) → $ZnSO_4$ (aq) + Cu (s)

(Zinc sulphate)

- When Cl_2 water is added to Potassium iodide, KI solution, iodine is liberated.

$2\,KI\ (aq) + Cl_2\ (aq) \rightarrow 2\,KCl\ (aq) + I_2\ (aq)$

Reactivity series of metals or activity series of metals:

A more reactive metal placed higher in the activity series can displace the less reactive metal from its solution.

The arrangement of metals in the decreasing order of their reactivity is called reactivity series or activity series of metals.

Table 1 : Activity Series of Metals

Element	Symbol	Reactivity
Potassium	K	(Most Reactive)
Barium	Ba	
Calcium	Ca	
Sodium	Na	
Magnesium	Mg	
Aluminium	Al	Metals more reactive than hydrogen
Zinc	Zn	
Iron	Fe	
Cobalt	Co	
Nickel	Ni	
Tin	Sn	
Lead	Pb	
Hydrogen	H	
Copper	Cu	
	Hg	Metals less reactive than hydrogen (Least Reactive)
Silver	Ag	
Gold	Au	
Platinum	Pt	

Note: The metal which is higher in the reactivity series is more reactive than the metals below it.

(g) Double Displacement Reaction: A reaction in which two different atoms or group of atoms or ions are mutually exchanged between two reactants or compound in their solution state is known as double displacement reaction .

Such reactions which produce precipitates which are insoluble solids are called precipitation reactions.

Example: $BaCl_2(s) + H_2SO_4\,(aq) \rightarrow BaSO_4(s) + 2HCl(aq)$

(h) Neutralisation Reaction: Reaction between an acid and a base to form salt and water.

Example: Acid + Base ⟶ Salt + Water

(i) Oxidation Reaction:

- Addition of oxygen and removal of hydrogen.
- Addition of electronegative element and removal of electropositive element.
- Loss of valence electron (s), **e.g.** $Na \longrightarrow Na^+ + e^-$

(j) Reduction:

- Addition of hydrogen and removal of oxygen.
- Addition of electropositive element and removal of electronegative element.
- Gain of valence electron (s). e.g. $Cl + e \longrightarrow Cl^-$

(k) Redox Reaction: Reaction in which oxidation and reduction occur simultaneously.

Example: $CuO + H_2 \xrightarrow{\text{heat}} Cu + H_2O$

Here hydrogen undergoes oxidation to form water and copper oxide undergoes reduction to from Cu.

Other examples:

$ZnO + C \longrightarrow Zn + CO$

Here, C undergoes oxidation to CO and ZnO undergoes reduction to Zinc.

$MnO_2 + 4\,HCl \longrightarrow MnCl_2 + 2H_2O + Cl_2$

Here, HCl undergoes oxidation by losing hydrogen and MnO_2 undergoes reduction by losing oxygen.

If a substance gains oxygen or loses hydrogen during a reaction it is said to be oxidised. If a substance loses oxygen or gain hydrogen during a reaction it is said to be reduced.

e.g. $ZnO + C \rightarrow Zn + CO$

Here C is said to be oxidised and ZnO is said to be reduced.

- **Oxidising agent:** The substance which oxidises other substances but itself undergoes reduction.
- Substances which provide oxygen or help in removing hydrogen.
- **Reducing agent:** The substance which reduces other substances but itself undergoes oxidation.
- Substances which provide hydrogen or helps in removing oxygen.

In the above example,

a) ZnO undergoes reduction but oxidises carbon. So ZnO is the oxidising agent.

b) Carbon undergoes oxidation but reduces zinc oxide. So carbon is the reducing agent.

According to electronic concept:

- Oxidation is a process which involves loss of one or more electrons.
- Reduction is a process which involves gain of electrons.

Oxidation Reduction reactions in everyday life:

Rusting of iron: It is an oxidation process.

- Rust is mainly hydrated ferric oxide ($Fe_2O_3.xH_2O$)
- $4Fe(S) + 3O_2(g) + 2x\ H_2O(l) \rightarrow 2Fe_2O_3.xH_2O(s)$
- Iron (from air) moisture hydrated ferric oxide (rust).

Corrosion: Process of slowly eating away of the metal due to the action of the atmospheric gases on the surface of the metals resulting in the formation of compounds.

- **E.g.** rusting of iron, formation of black coating on silver, green coating on copper etc.
- Corrosion of metals can be prevented by providing protective layers with substances like grease, paint or metal coatings so that the metals do not come in contact with damp air.
- Rusting of iron can be effectively prevented by galvanization (coating iron with zinc).

Rancidity: Oxidation of fats or oils in food resulting into bad smell and taste.

- Antioxidants are added to prevent rancidity.
- Vacuum packing and filling with inert gases in food packing also prevents rancidity.

Slow and fast reactions:

- Reactions between covalent compounds are slow.

Example:

$$\underset{\text{Acetic acid}}{CH_3COOH\ (aq)} + \underset{\text{Ethyl alcohol}}{C_2H_5OH(l)} \xrightarrow{\text{Conc. } H_2SO_4} \underset{\text{Ethyl acetate (sweet smelling ester)}}{CH_3COOC_2H_5(aq)} + H_2O(l)$$

- Reactions between ionic compounds in solution are fast reactions.

Example: Burning of magnesium in air to form magnesium oxide.

- **Catalysts:** A substance which increases or decreases the rate of a chemical reaction without itself being consumed.
- Catalysts are specific in action.

Example: $N_2(g) + 3\ H_2\ (g) \xrightarrow{\text{Fe (catalyst)}} \underset{\text{(Ammonia)}}{2NH_3\ (g)}$

Similarly, Vanadium pentoxide (V_2O_5) is used as a catalyst in the manufacture of Sulphur trioxide.

$$2SO_2(g) + \underset{\text{catalyst}}{O_2(g)} \xrightarrow{V_2O_5} SO_3(g)$$

- Enzymes are biological catalyst.

Example: Invertase enzyme converts sucrose into glucose and fructose.

$$\underset{\text{Sucrose}}{C_{12}H_{22}O_{11}} + H_2O \xrightarrow{\text{invertase (catalyst)}} \underset{\text{Glucose}}{C_6H_{12}O_6} + \underset{\text{fructose}}{C_6H_{12}O_6}$$

- Maltose requires another enzyme called maltase.

Example:

$$\underset{\text{Maltose}}{C_{12}H_{22}O_{11}} + H_2O \xrightarrow{\text{Maltase (catalyst)}} \underset{\text{Glucose}}{2C_6H_{12}O_6}$$