Chemical Reactions and Equations

- In a chemical reaction, at least one of the following will occur:
- Change in state
- Change in colour
- Evolution of a gas
- Change in temperature
- Formation of a precipitate

A **chemical equation** is the symbolic representation of a chemical reaction in the form of chemical formulae, signs, symbols, and directions. In which the reactant entities are given on the left-hand side and the product entities on the right-hand side.

- **Balanced chemical equation**

Reactants → Products
LHS RHS
Total number of atoms on the LHS = Total number of atoms on the RHS

- How to balance an equation
- Write reactants and products
- Balance the maximum number of a particular atom on both sides
- Balance other atoms
- A complete balanced equation should look like

CO g + 2H2 g →340 atm CH3OH l
Types of reactions

- **Combination reaction**
  - Two or more reactants combine to form one single product.

Examples: CaO s + H2O l → Ca(OH)2 aqCalcium oxide Water Calcium hydroxid e (Quick lime) (Slaked lime) C s + O2 g → CO2 gCarbon Oxygen Ca rbon dioxide 2H2 g + O2 g → 2H2O lHydrogen Oxygen Water

- **Exothermic reaction** – Heat gets released in the reaction. Most combination reactions are exothermic. For example,

$$\underset{\text{Calcium oxide (Quick lime)}}{CaO(s)} + \underset{\text{Water}}{H_2O(l)} \rightarrow \underset{\text{Calcium hydroxide (Slaked lime)}}{Ca(OH)_2(aq)}$$

- **Endothermic reaction** – Heat is absorbed in the reaction. Very few combination reactions are endothermic. For example,

  12N2 g + O2 g → NO2 g

- **Decomposition reaction**
  - A single reactant breaks into several simple products.

Examples:2FeSO4 Ferrous sulphate→ Δ Fe2O3Ferric oxide+SO2 +SO3CaCO3Limestone → Δ CaOCalcium oxide+ CO22AgClSilver chloride→ Δ 2AgSilver+Cl2

  - All decomposition reactions are **endothermic [they absorb heat]**.

- **Displacement reactions:**
  - In displacement reactions, a more reactive metal can displace a less reactive metal from their compounds in aqueous solutions. (However, a less reactive metal cannot displace a more reactive metal.)

    Example: CuSO4 + Zn → ZnSO4 + Cu Copper Sulphate Zinc Zinc Sulphate Copper (Blue) (Colourless) (Red)
    Fe s + CuSO4 aq → Cu s + FeSO4 aqIron Copper sulphate Copper Iron sulphate

- **Double displacement reaction**
  - Exchange of ions occurs between two compounds.

Example; Na2SO4 aq + BaCl2 s → BaSO4 aq + 2NaCl sSodium sulphate Barium chloride Barium sulphate Sodium chloride

- When the aqueous solution of two compounds react by exchanging their respective ions, such that one of the products formed is insoluble salt and appears in the form of a precipitate, then the reaction is said to be **precipitation reaction**.
- When an acid solution reacts with a base and the two exchange their respective ions, such that only salt and water are products, then the reaction is called **neutralisation reaction**.
- When two compounds react with each other and displace their ions, in such a manner that one of the product formed either decomposes into gaseous compounds or is formed in gaseous state, then the reaction is called **gas-forming reaction**.

- **Oxidation** →When a substance gains oxygen or loses hydrogen

$$2Cu + O_2 \xrightarrow{Heat} 2CuO \qquad [Oxidation\ of\ Cu]$$

$$CuO + H_2 \xrightarrow{Heat} Cu + H_2O \qquad [Oxidation\ of\ H_2]$$

- **Oxidation in everyday life**

- **Corrosion –** When a metal is oxidised by the action of air and moisture [that's why metals are coated]
- **Rancidity** – When fats and oils are oxidised, their smell and taste change [that's why food is kept in air-tight containers]

- **Reduction** → When one substance loses oxygen or gains hydrogen

$$CuO + H_2 \xrightarrow{\text{Heat}} Cu + H_2O \qquad \text{[Reduction of CuO]}$$

- **Redox** – Oxidation–reduction reaction