Chapter-2

Atoms, Molecules, and Stoichiometry

Relative Masses of Atoms and Molecules

Because individual atoms have such small masses (e.g., a hydrogen atom weighs 1.67×10^{-27} kg), chemists use a relative scale based on the unified atomic mass unit (u).

- **Unified Atomic Mass Unit (u):**one-twelfth of the mass of a carbon-12 atom.
- **Relative Atomic Mass (Ar):** The ratio of the weighted average mass of an atom of an element to the unified atomic mass unit. It has no units and accounts for the natural abundances of an element's isotopes.
- **Relative Isotopic Mass:** The mass an atom of a specific isotope of an element relative unified atomic mass.
- **Relative Molecular Mass (Mr):** The sum of the relative atomic masses of all atoms in one molecule of a compound.
- **Relative Formula Mass (Mr):** Used for ionic compounds or giant structures (like SiO_2), calculated in the same way as Mr by adding Ar values in the simplest formula unit.

The Mole Concept

Amount of substance of a system which contain as many elementary entities as there are number of atoms in 0.012 kilogram of carbon-12.

- **Avogadro Constant (L or N_A):** The number of particles (6.02×10^{23}) in one mole of any substance.

Moles and Mass

Molar Mass (M): The mass of one mole of a substance in grams (g mol^{-1}). Molar mass is simply the molecular mass expressed in gram per mole. For example molar mass of CO_2 is 44 g/mol.

Number of moles (mol) = $\frac{mass\ of\ substance\ (g)}{molar\ mass\ (g/mol)}$

Moles and Volume

Molar volume: the volume (dm^3) of one mole substance. Since volume depends on pressure and temperature as well, condition of pressure and temperature are necessary to mention.

Molar volume of any gas/vapour =22.4 dm^3 at STP (273K and 101 kPascal)
=24.0 dm^3 at RTP (298K and 101 kPascal)

Number of moles (mol) = $\frac{volume\ of\ substance\ at\ RTP\ (dm3)}{24\ (dm3)}$

Note:
1 dm^3= 1 liter
1m^3 = 1000 dm^3
1dm^3 = 1000 cm^3

STP: Standard Temperature and Pressure
RTP: Room Temperature and Pressure
dm=decimeter
cm=centimeter

Moles and Concentration
Concentration of a solution is defined as amount of substance (mole) present in unit volume (dm^3) of solution.
Concentration = $\frac{moles}{volume\ (dm3)}$

Chemical Formulae

- **Empirical Formula:** The simplest whole-number ratio of elements in a compound. For ionic compounds, the formula is always the empirical formula. For example empirical formula of glucose is CH_2O, empirical formula of ethane is CH_3
  Calculation of empirical formula
  Step:1- divide the mass or percentage by mass of each elements by its relative atomic mass.
  Step:2- divide the results of step 1 by the smallest result
  Step:3- change the results of step 2 in to whole number ratio.

- **Molecular Formula:** The actual number of each atom in a molecule. It is always a multiple of the empirical formula. For examples, molecular formula of glucose is $C_6H_{12}O_6$, molecular formula of ethane is C_2H_6.

  Molecular formula = (empirical formula)$_n$

$$n = \frac{molecular\ mass}{empirical\ formula\ mass}$$

- **Hydrated Compounds:** Compounds containing **water of crystallisation**, represented with a dot (e.g., $CuSO_4.5H_2O$). Anhydrous compounds lack this water.

Stoichiometry and Equations

Balanced chemical equations show the mole ratios (stoichiometry) of reactants and products.

- **Ionic Equations:** Omit **spectator ions** (ions that do not change during the reaction) to focus on the actual chemical change.

- **State Symbols:** (s) solid, (l) liquid, (g) gas, and (aq) aqueous.

- **Reacting Masses:** Use molar masses and stoichiometric ratios to calculate the mass of products formed from a given mass of reactants.

- **Limiting and Excess Reagents:** The **limiting reagent** is the reactant that is completely consumed, while the **excess reagent** remains after the reaction.

- **Percentage Yield:**

  Percentage yield} = $\frac{Actual\ yield}{Predicted\ yield} \times 100$.

- **Gas Stoichiometry:** For reactions involving only gases, the mole ratio is equivalent to the volume ratio. For example, in $2H_2 + O_2 \text{---->} 2H_2O$, 20 cm^3 of H_2 reacts with 10 cm^3 of O_2.

Titration

A titration is a common laboratory procedure used to determine the exact amount of a substance (or its concentration) present in a solution, Also known as volumetric analysis, it involves reacting a solution of unknown concentration with a solution of known concentration, called a standard solution.

Types of Titrations

While acid-base neutralisation is the most frequent type, there are other variations:

- **Acid–Base Titrations:** Used to find the concentration of an acid or alkali.

- **Redox Titrations:** Used to calculate the amount of a substance that can be oxidised or reduced. For example, **potassium manganate(VII)** can be titrated against iron(II) ions to determine the iron content in a sample.

- **Deducing Stoichiometry:** Titration results can also be used to find the simplest mole ratio (stoichiometry) of a reaction if the concentrations and volumes of both reactants are already known.

Essential Equipment

Titrations require specialized glassware designed for high accuracy:

- **Burette:** A long, graduated tube used to deliver and measure variable volumes of liquid, usually to an accuracy of 0.05 cm^3.
- **Volumetric Pipette:** Used to transfer a specific, fixed volume of liquid (often 10 cm^3or 25 cm^3).
- **Conical Flask:** Used to hold the reaction mixture and allow for easy swirling without spilling.