CHAPTER

1 Some Basic Concepts of Chemistry

Some Useful Conversion Factors

$1\ \text{Å} = 10^{-10}\text{m},\ 1\text{nm} = 10^{-9}\ \text{m}$

$1\ \text{pm} = 10^{-12}\text{m}$

$1\ \text{litre} = 10^{-3}\ \text{m}^3 = 1\ \text{dm}^3$

$1\ \text{atm} = 760\ \text{mm or torr}$

$= 101325\ \text{Pa or Nm}^{-2}$

$1\ \text{bar} = 10^5\ \text{Nm}^{-2} = 10^5\ \text{Pa}$

$1\ \text{calorie} = 4.184\ \text{J}$

$1\ \text{electron volt(eV)} = 1.6022 \times 10^{-19}\ \text{J}$

$(1\ \text{J} = 10^7\ \text{ergs})$

$(1\ \text{cal} > 1\ \text{J} > 1\ \text{erg} > 1\ \text{eV})$

Atomic Mass or Molecular Mass

Mass of one atom or molecule in a.m.u.

$C \rightarrow 12$ amu

$NH_3 \rightarrow 17$ amu

Actual Mass

Mass of one atom or molecule in grams:

$C \rightarrow 12 \times 1.6 \times 10^{-24}$ g

$CH_4 \rightarrow 16 \times 1.6 \times 10^{-24}$ g

Relative Atomic Mass or Relative Molecular Mass

Mass of one atom or molecule w.r.t. $1/12^{th}$ or ^{12}C atom:

$C \rightarrow 12$

$CH_4 \rightarrow 16$

It is unitless.

Gram Atomic Mass or Gram Molecular Mass

Mass of one mole of atom or molecule:

$C \rightarrow 12$ g

$CO_2 \rightarrow 44$ g

It is also called as molar mass.

Definition of Mole

One mole is a collection of that many entities as there are number of atoms exactly in 12 gm of C-12 isotope.

The number of atoms present in exactly 12 gm of C-12 isotope is called Avogadro's number [$N_A = 6.022 \times 10^{23}$]

$$1\text{u} = 1\text{amu} = (1/12)^{th} \text{ of mass of 1 atom of } C^{12} = \frac{1g}{N_A}$$

$$= 1.66 \times 10^{-24}\ \text{g}$$

For Elements

- 1 g atom = 1 mole of atoms = N_A atoms
- g atomic mass (GAM) = mass of N_A atoms in g
- Mole of atoms = $\dfrac{\text{Mass }(g)}{\text{GAM or Molar mass}}$

For Molecule

- 1g molecule = 1 mole of molecule = N_A molecule
- g molecular mass (GMM) = mass of N_A molecule in g.
- Mole of molecule = $\dfrac{\text{Mass }(g)}{\text{GMM or Molar mass}}$

1 Mole of Substance

- Contains 6.022×10^{23} particles.
- Weighs as much as molecular mass/ atomic mass/ionic mass in grams.
- If it is a gas, one mole occupies a volume of 22.4 L at 1 atm & at 273 K or 22.7 L at STP.

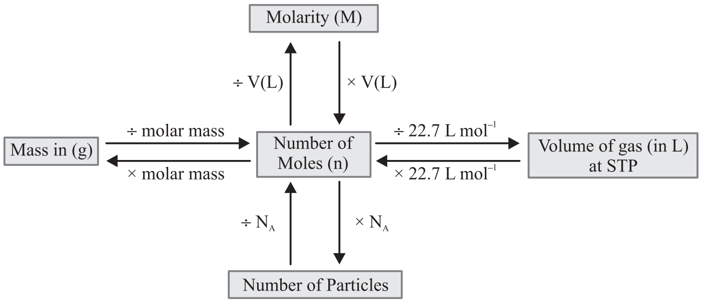
For Ionic Compounds

- 1 g formula unit = 1 mole of formula unit = N_A formula unit.
- g formula mass (GFM) = mass of N_A formula unit in g.
- Mole of formula unit = $\dfrac{\text{Mass }(g)}{\text{GMM or Molar mass}}$

Vapour density

Ratio of density of vapour to the density of hydrogen at similar pressure and temperature.

$$\text{Vapour density} = \frac{\text{Molar mass}}{2}$$



Stoichiometry Based Concept

$$aA + bB \rightarrow cC + dD$$

- a,b,c,d, represents the ratios of moles, volumes [for gaseous] molecules in which the reactants react or products formed.
- a,b,c,d does not represent the ratio of masses.
- The stoichiometic amount of components may be related as:

$$\frac{\text{Moles of A reacted}}{a} = \frac{\text{Moles of B reacted}}{b} = \frac{\text{Moles of C reacted}}{c} = \frac{\text{Moles of D reacted}}{d}$$

Concept of Limiting Reagent

If data of more than one reactant is given then first convert all the data into moles then divide the moles of reactants with their respective stoichiometric coefficient. The reactant having minimum ratio will be L.R. then find the moles of product formed or excess reagent left by comparing it with L.R. through stoichiometric concept.

Percentage Purity

The percentage of a specified compound or element in an impure sample may be given as:

$$\%\text{purity} = \frac{\text{Actual mass of compound}}{\text{Total mass of sample}} \times 100$$

If impurity is unknown, it is always considered as inert (unreactive) material.

Empirical and Molecular Formula

- **Empirical formula:** Formula depicting constituent atoms in their simplest ratio.
- **Molecular formula:** Formula depicting actual number of atoms in one molecule of the compound.
- The molecular formula is generally an integral multiple of the empirical formula.

  i.e. molecular formula = empirical formula × n

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

Concentration Terms

Concentration Type	Mathematical Formula	Concept
Percentage by mass	$\%\left(\frac{w}{w}\right) = \frac{\text{Mass of solute} \times 100}{\text{Mass of solution}}$	Mass of solute (in gm) present in 100 gm of solution.
Volume percentage	$\%\left(\frac{v}{v}\right) = \frac{\text{Volume of solute} \times 100}{\text{Volume of solution}}$	Volume of solute (in cm^3) present in 100 cm^3 of solution.
Mass-volume percentage	$\%\left(\frac{w}{v}\right) = \frac{\text{Mass of solute} \times 100}{\text{Volume of solution}}$	Mass of solute (in gm) present in 100 cm^3 of solution.
Parts per million	$\text{ppm} = \frac{\text{Mass of solute} \times 10^6}{\text{Mass of solution}}$	Parts by mass of solute per million parts by mass of the solution.
Mole fraction	$X_A = \frac{\text{Mole of A}}{\text{Mole of A} + \text{Mole of B} + \text{Mole of C} + \ldots}$ $X_B = \frac{\text{Mole of B}}{\text{Mole of A} + \text{Mole of B} + \text{Mole of C} + \ldots}$	Ratio of number of moles of one component to the total number of moles.
Molarity	$M = \frac{\text{Mole of solute}}{\text{Volume of solution (in L)}}$	Moles of solute in one liter of solution.
Molality	$m = \frac{\text{Mole of solute}}{\text{Mass of solvent (Kg)}}$	Moles of solute in one kg of solvent.

Mixing of Solutions

It is based on law of conservation of moles.

(i) Two solutions having same solute:

$$\text{Final molarity} = \frac{\text{Total moles}}{\text{Total volume}} = \frac{M_1V_1 + M_2V_2}{V_1 + V_2}$$

(ii) Dilution Effect: Final molarity, $M_2 = \dfrac{M_1V_1}{V_1 + V_2}$