Mole Concept and Stoichiometry

Gas Laws

Boyle's Law
The volume of a given mass of a dry gas is inversely proportional to its pressure at a constant temperature.

$P_1V_I = P_2V_2 = k$ at constant temperature

Charles's Law
The volume of a given mass of a dry gas is directly proportional to its absolute temperature if the pressure is kept constant.

$\frac{V1}{T1} = \frac{V2}{T2} = k$ at constant pressure

Gas Equation
The volume of a given mass of a dry gas is inversely proportional to the pressure and directly proportional to the absolute temperature.

$V \propto \frac{1}{P} \times T$ or $\frac{PV}{T} = k$

Standard or Normal Temperature and Pressure
For temperature: 0°C or 273 K

For pressure: 760 mm or 76 cm of Hg

Gay-Lussac's law of combining volumes
At the same temperature and pressure, the volume of gases taking part in a chemical reaction as either reactants or products bears a whole number ratio to one another.

Example:	$H_{2(g)}$	+	$Cl_{2(g)}$	$\longrightarrow$	$2HCl_{(g)}$
	1 vol.		1 vol.		2 vols.

The ratio of reacting gases and products is 1:1:2, which is a simple ratio.

Avogadro's Law
Under the same conditions of temperature and pressure, equal volumes of all the gases contain the same number of molecules.

Example: A molecule of NH_3 is made of one atom of nitrogen and three atoms of hydrogen.

$N_{2(g)}$	+	$3H_{2(g)}$	$\longrightarrow$	$2NH_{3(g)}$
1 vol.		3 vols.		2 vols.
1 molecule		3 molecules		2 molecules
Nitrogen		Hydrogen		Ammonia

Atomicity

The number of atoms in a molecule of an element is called its atomicity.

a. **Monatomic**: It is composed of only one atom.
   Examples: Inert gases such as Helium, Neon etc.
b. **Diatomic**: It is composed of two identical atoms.
   Examples: H_2, O_2, Cl_2 etc.
c. **Triatomic**: It is composed of three identical atoms.
   Example: Ozone (O_3)
d. **Tetratomic**: It is composed of four identical atoms.
   Example: Phosphorus (P_4)
e. **Octatomic**: It is composed of eight identical atoms.
   Example: Sulphur (S_8)

Atomic Mass or Relative Atomic Mass

It is the number which represents how many times one atom of an element is heavier than $1/12^{th}$ the mass of an atom of carbon-12 (^{12}C).

$$\text{Relative atomic mass} = \frac{\text{Mass of an atom of an element}}{1/12^{th} \text{ the mass of one C-12 atom}}$$

Molecular Mass or Relative Molecular Mass

It is the number which represents how many times one molecule of an element is heavier than $1/12^{th}$ the mass of an atom of carbon-12 (^{12}C).

$$\text{Relative molecular mass} = \frac{\text{Mass of one molecule of an element}}{1/12^{th} \text{ the mass of one C-12 atom}}$$

Gram Atomic Mass

The atomic mass of an element expressed in gram is called gram atomic mass.
Example: Gram atomic mass of oxygen is 16 gram.

Gram Molecular Mass

The molecular mass of a substance expressed in gram is called gram molecular mass or molar mass.
Example: Gram molecular mass of water is 18 gram.

Mole Concept

A mole is a collection of 6.022×10^{23} particles.
A mole is defined as the amount of a substance containing elementary particles such as atoms, molecules or ions in 12 gram of carbon-12 (^{12}C).

Avogadro's Number

It is defined as the number of atoms present in 12 gram of C-12 isotope, i.e. $6{\cdot}023 \times 10^{23}$ atoms. It is denoted by N_A or L.

N_A = $6{\cdot}023 \times 10^{23}$

1 mole of atoms = $6{\cdot}023 \times 10^{23}$ atoms

1 mole of molecules = $6{\cdot}023 \times 10^{23}$ molecules

1 mole of electrons = $6{\cdot}023 \times 10^{23}$ electrons

1 mole of a gas = 22·4 litre at STP

Applications of Avogadro's Law

i. It explains Gay-Lussac's law.
ii. It determines atomicity of the gases.
iii. It determines the molecular formula of a gas.
iv. It determines the relation between molecular mass and vapour density.
v. It gives the relationship between gram molecular mass and gram molar volume.

Relative Vapour Density (VD)

Relative vapour density is the ratio between the masses of equal volumes of a gas (or vapour) and hydrogen under the same conditions of temperature and pressure.

$$\text{Relative VD} = \frac{\text{Mass of volume 'v' of the gas under similar conditions}}{\text{Mass of volume 'v' of hydrogen gas under similar conditions}}$$

Relative molecular mass of a gas or vapour = $2 \times \text{VD}$

Important Formulae

Mole and Gram Atomic Mass: One mole of atoms $= 6.022 \times 10^{23}$ atoms = Gram atomic mass of an element = 1 gram atom of the element	
Mole and Gram Molecular Mass: One mole of molecules $= 6.022 \times 10^{23}$ molecules = Gram molecular mass = 1 gram molecule of the compound	
Mole in terms of volume: One mole of a gas = 22.4 litres at STP	
Moles of an element $= \frac{\text{Mass of the element}}{\text{Atomic mass or GAW}}$	Moles of a compound $= \frac{\text{Mass of the compound}}{\text{Molecular mass or GMW}}$
Mass of one atom $= \frac{\text{Atomic Mass or GAW}}{6.022 \times 10^{23}}$	Mass of one molecule $= \frac{\text{Molecular Mass or GAW}}{6.022 \times 10^{23}}$
Number of molecules = Moles $\times 6.022 \times 10^{23}$	Number of atoms = Moles $\times 6.022 \times 10^{23}$

Percentage Composition

The percentage by weight of each element present in a compound is called percentage composition of the compound.

$$\text{Percentage} = \frac{\text{Weight of an element in a molecule of a compound}}{\text{Gram molecular weight of compound}} \times 100$$

Empirical Formula

It is the chemical formula which gives the simplest ratio in whole numbers of atoms of different elements present in one molecule of the compound.

Empirical Formula Mass

It is the sum of atomic masses of various elements present in the empirical formula.

Empirical Formula Weight (EFW)

The empirical formula weight is the atomic masses of the elements present in the empirical formula.

$$\text{EFW of } H_2O_2 = 2 \times (H) + 2 \times (0)$$
$$= 2 \times 1 + 2 \times 16$$
$$= 34 \text{ amu}$$

Molecular Formula

It denotes the actual number of atoms of different elements present in one molecule of the compound.

Molecular formula = Empirical formula × n

Where $n = \frac{\text{Molecular weight}}{\text{Empirical formula weight}}$

Relationship between Empirical Formula and Molecular Formula

Molecular formula = Empirical formula × *n*

Where 'n' is a positive whole number

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

Chemical Equation

A shorthand notation of describing an actual chemical reaction in terms of symbols and formula along with the number of atoms and molecules of the reactants and products is called a chemical equation.

A chemical equation is a balanced account of a chemical transaction.

$$2KClO_3(s) \xrightarrow{MnO_2} 2KCl(s) + 3O_2(g)$$