INTRODUCTION TO CHEMISTRY

Atomic hypothesis:

Keeping in view various laws of chemical combination, a theoretical proof for the validity of different laws was given by John Dalton in the form of hypothesis called Dalton's atomic hypothesis. Postulates of Dalton's hypothesis are as follows:

(i) Each element is composed of extremely small particles called atoms which can take part in chemical combination.

(ii) All atoms of a given element are identical i.e., atoms of a particular element are all alike but differ from atoms of other elements.

(iii) Atoms of different elements possess different properties (including different masses).

(iv) Atoms are indestructible i.e., atoms are neither created nor destroyed in chemical reactions.

(v) Atoms of elements combine to form molecules and compounds are formed when atoms of more than one element combine.

(vi) In a given compound, the relative number and kind of atoms is constant.

Modern atomic hypothesis: The main modifications made in Dalton's hypothesis as a result of new discoveries about atoms are :

(i) Atom is no longer considered to be indivisible.

(ii) Atoms of the same element may have different atomic weights. E.g. isotopes of oxygen O^{16}, O^{17} and O^{18}.

(iii) Atoms of different element may have same atomic weights. E.g. isobars Ca^{40} and Ar^{40}.

(iv) Atom is no longer indestructible. In many nuclear reactions, a certain mass of the nucleus is converted into energy along with α, β and γ rays.

(v) Atoms may not always combine in simple whole number ratios. E.g. in sucrose ($C_{12}H_{22}O_{11}$), the elements carbon, hydrogen and oxygen are present in the ratio of 12 : 22 : 11 and the ratio is not a simple whole number ratio.

Atomic & Molecular masses:

Atomic mass: It is the average relative mass of atom of element as compared with $\frac{1}{12}$ times the mass of an atom of carbon-12 isotope.

$$\text{Atomic mass } = \frac{\text{Average mass of an atom}}{1/12 \times \text{Mass of an atom of } C^{12}}$$

Average atomic mass: If an element exists in two isotopes having atomic masses 'a' and 'b' in the ratio m : n, then average atomic mass = $\frac{(m \times a) + (n \times b)}{m + n}$. Atomic mass is expressed in amu. 1 amu = 1.66 × 10^{-24} g. One atomic mass unit (amu) is equal to $\frac{1}{12}$th of the mass of an atom of carbon-12 isotope.

Gram atomic mass (GAM): Atomic mass of an element expressed in grams is called Gram atomic mass or gram atom or mole atom.

(i) Number of gram atoms = $\frac{\text{Mass of an element}}{\text{GAM}}$

(ii) Mass of an element in g = No. of gram atoms × GAM

(iii) Number of atoms in 1 GAM = 6.02×10^{23}

Number of atoms in a given substance = No. of gram atoms × 6.02×10^{23} = $\frac{\text{Mass}}{\text{GAM}} \times 6.02 \times 10^{23}$

(iv) Number of atoms in 1 g of element = $\frac{6.02 \times 10^{23}}{\text{GAM}}$

(v) Mass of one atom of the element (in g) = $\frac{\text{GAM}}{6.02 \times 10^{23}}$

Molecular mass: Molecular mass of a molecule, of an element or a compound may be defined as a number which indicates how many times heavier is a molecule of that element or compound as compared with $\frac{1}{12}$ of the mass of an atom of carbon-12. Molecular mass is also expressed in amu.

$$\text{Molecular mass} = \frac{\text{Mass of one molecule of the substance}}{1/12 \times \text{Mass of one atom of C-12}}$$

Actual mass of one molecule = Mol. mass (in amu) × 1.66×10^{-24} g

Molecular mass of a substance is the additive property and can be calculated by adding the atomic masses of atoms present in one molecule.

Gram molecular mass (GMM): Molecular mass of an element or compound when expressed in g is called its gram molecular mass, gram molecule or mole molecule.

Number of gram molecules = $\frac{\text{Mass of substance}}{\text{GMM}}$

Mass of substance in g = No. of gram molecules × GMM

Average atomic mass and molecular mass

$\bar{A}$ (Average atomic mass) = $\frac{\sum A_i X_i}{\sum X_{total}}$

(Average molecular mass) = $\frac{\sum M_i X_i}{\sum X_{total}}$

Where A_1, A_2, A_3 …… are atomic mass of species 1, 2, 3,…. etc. with % as X_1, X_2, X_3 …… etc. Similar terms are for molecular masses.

The Mole Concept

One mole of any substance contains a fixed number (6.022×10^{23}) of any type of particles (atoms or molecules or ions) and has a mass equal to the atomic or molecular weight, in grams. Thus it is correct to refer to a mole of helium, a mole of electrons or a mole of any ion, meaning respectively Avogadro's number of atoms, electrons or ions.

$$\text{Number of moles} = \frac{\text{Weight (grams)}}{\text{Weight of one mole (g/mole)}} = \frac{\text{Weight}}{\text{GAM or GMM}}$$

Note : 1 mole = 1 g-atom = 1 g-molecule = 1 g-ion.

Properties of Gases

The state of matter in which the molecular forces of attraction between the particles of matter are minimum, is known as gaseous state. It is the simplest state and shows great uniformity in behaviour.

Characteristics of gases

(1) Gases or their mixtures are homogeneous in composition.

(2) Gases have very low density due to negligible intermolecular forces.
(3) Gases have infinite expansibility and high compressibility.
(4) Gases exert pressure.
(5) Gases possess high diffusibility.
(6) Gases do not have definite shape and volume like liquids.
(7) Gaseous molecules move very rapidly in all directions in a random manner i.e., gases have highest kinetic energy.
(8) Gaseous molecules collide with one another and also with the walls of container.
(9) Gases can be liquefied, if subjected to low temperatures & high pressures.
(10) Thermal energy of gases >> molecular attraction.
(11) Gases undergo similar change with the change of temperature and pressure. In other words, gases obey certain laws known as gas laws.

Measurable properties of gases

The characteristics of gases are described fully in terms of four parameters or measurable properties:
(i) The volume, V, of the gas. (ii) Its pressure, P
(iii) Its temperature, T (iv) The amount of the gas (i.e., mass or number of moles).

(1) Volume : (i) Since gases occupy the entire space available to them, the measurement of volume of a gas only requires a measurement of the container confining the gas.
(ii) Volume is expressed in litres (L), millilitres (mL) or cubic centimetres (cm^3), cubic metres (m^3).
(iii) $1\ L = 1000\ mL$; $1\ mL = 10^{-3}\ L$; $1\ L = 1\ dm^3 = 10^{-3}\ m^3$
$1\ m^3 = 10^3\ dm^3 = 10^6\ cm^3 = 10^6\ mL = 10^3\ L$

(2) Mass : (i) The mass of a gas can be determined by weighing the container in which the gas is enclosed and again weighing the container after removing the gas. The difference between the two weights gives the mass of the gas.
(ii) The mass of the gas is related to the number of moles of the gas i.e.

$$\text{moles of gas (n)} = \frac{\text{Mass in grams}}{\text{Molar mass}} = \frac{m}{M}$$

(3) Temperature : (i) Gases expand on increasing the temperature. If temperature is increased twice, the square of the velocity (v^2) also increases two times.
(ii) Temperature is measured in centigrade degree (ºC) or celsius degree with the help of thermometers. Temperature is also measured in Fahrenheit (ºF).
(iii) S.I. unit of temperature is kelvin (K) or absolute degree.

$$K = {}^oC + 273$$

(iv) Relation between ºF and ºC is $\frac{{}^oC}{5} = \frac{{}^oF - 32}{9}$

(4) Pressure : (i) Pressure of the gas is the force exerted by the gas per unit area of the walls of the container in all directions. Thus, Pressure (P) $= \frac{\text{Force(F)}}{\text{Area(A)}} = \frac{\text{Mass(m)} \times \text{Acceleration(a)}}{\text{Area(A)}}$

(ii) Pressure exerted by a gas is due to kinetic energy $(KE = \frac{1}{2}mv^2)$ of the molecules. Kinetic energy of the gas molecules increases, as the temperature is increased.
(iii) Pressure of a gas is measured by manometer or barometer.
(iv) Commonly two types of manometers are used:
(a) Open end manometer; (b) Closed end manometer
(v) The S.I. unit of pressure, the pascal (Pa), is defined as 1 newton per metre square. It is very small unit.

$1 Pa = 1 Nm^{-2} = 1 kgm^{-1}s^{-2}$

(vi) C.G.S. unit of pressure is dynes cm^{-2}.

(vii) M.K.S. unit of pressure is Newton m^{-2}. The unit Newton m^{-2} is sometimes called pascal (Pa).

(viii) Higher unit of pressure is bar, kPa or MPa.

$1 bar = 10^5 Pa = 10^5 Nm^{-2} = 100 KNm^{-2} = 100 KPa$

(ix) Several other units used for pressure are,

Name	Symbol	Value
bar	*bar*	1 bar = 10^5 Pa
atmosphere	*atm*	1 atm = 1.01325×10^5 Pa
Torr	Torr	1 Torr = $\frac{101325}{760}$ Pa = 133.322 Pa
millimetre of mercury	mm *Hg*	1 mm Hg = 133.322 Pa

Ideal Gas Equation

$PV = nRT$

where, P : Pressure of gas ; V : Volume of gas ; n = Number of moles of gas

T : Temperature of gas ; R : Universal gas constant.

Values of R : $0.082 LatmK^{-1}mol^{-1}$; $8.314 JK^{-1}mol^{-1}$; $1.987 CalK^{-1}mol^{-1}$

Prefixes used in the SI System

Multiple	Prefix	Symbol
10^{-24}	yocto	y
10^{-21}	zepto	z
10^{-18}	atto	a
10^{-15}	femto	f
10^{-12}	pico	p
10^{-9}	nano	n
10^{-6}	micro	μ
10^{-3}	milli	m
10^{-2}	centi	c
10^{-1}	deci	d
10	deca	da
10^{2}	hecto	h
10^{3}	kilo	k
10^{6}	mega	M
10^{9}	giga	G
10^{12}	tera	T
10^{15}	peta	P
10^{18}	exa	E
10^{21}	zeta	Z
10^{24}	yotta	Y