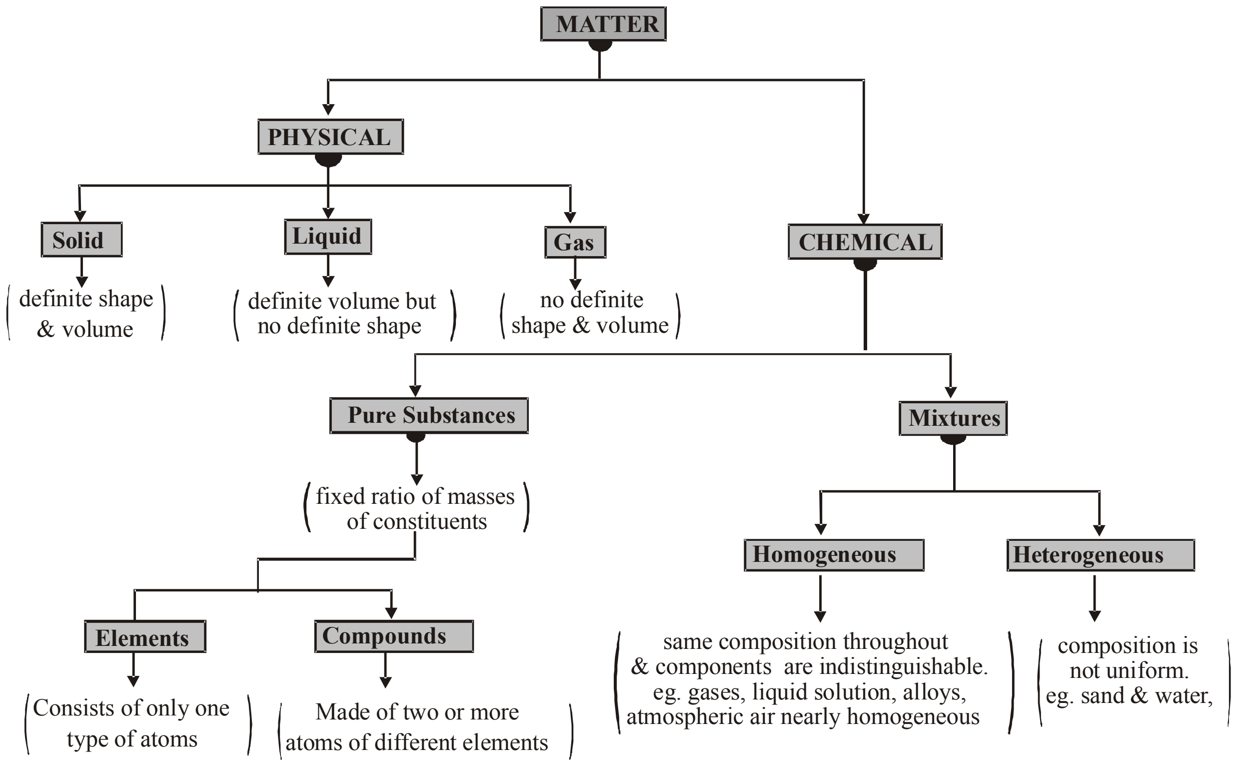
SOME BASIC CONCEPTS OF CHEMISTRY

INTRODUCTION

Chemistry deals with the composition, structure and properties of matter. These aspects can be best described and understood in terms of basic constituents of matter: **atoms** and **molecules**. That is why chemistry is called the science of atoms and molecules. Can we see, weight and perceive these entities? Is it possible to count the number of atoms and molecules in a given mass of matter and have a quantitative relationship between the mass and number of these particles (atoms and molecules)? We will like to answer some of these questions in this Unit. We would further describe how physical properties of matter can be quantitatively described using numerical values with suitable units.



Classification of universe

Universe is classified into two types i.e. matter and energy.

(A) **MATTER :** The thing which occupy space and having mass which can be felt by our five senses is called matter. Matter is further classified into two categories :

(I) Physical classification (II) Chemical classification

PHYSICAL CLASSIFICATION

It is based on physical state under ordinary conditions of temperature and pressure, **so on the basis of two nature of forces matter** can be classified into the following three ways :

(a) Solid (b) Liquid (c) Gas

(a) **Solid :** A substance is said to be solid if it possesses a definite volume and a definite shape.

e.g. Sugar, Iron, Gold, Wood etc.

(b) **Liquid :** A substance is said to be liquid if it possesses a definite volume but not definite shape. They take the shape of the vessel in which they are palced.

e.g. Water, Milk, Oil, Mercury, Alcohol etc.

(c) **Gas :** A substance is said to be gas if it neither possesses a definite volume nor a definite shape. This is because they completely occupy the whole vessel in which they are placed.

e.g. Hydrogen(H_2), Oxygen(O_2), Carbon dioxide(CO_2) etc.

Chemical Classification

It may be classified into two types :

(a) Pure Substances (b) Mixtures

(a) Pure Substance : A material containing only one type of substance. Pure Substance can not be separated into simpler substance by physical method.

e.g. : Elements = Na, Mg, Ca .... etc.

Compounds = HCl, H_2O, CO_2, HNO_3 .... etc.

Pure substances are classified into two types :

(a) Elements (b) Compounds

(i) Elements : The pure substances containing only one kind of atoms.

It is classified into 3 types (depend on physical and chemical property)

(i) Metal $\rightarrow$ Zn, Cu, Hg, Ac, Sn, Pb etc.

(ii) Non-metal $\rightarrow$ N_2, O_2, Cl_2, Br_2, F_2, P_4, S_8 etc.

(iii) Metalloids $\rightarrow$ B, Si, As, Te etc.

(ii) Compounds : It is defined as pure substances containing more than one kind of elements or atoms which are combined together in a fixed proportion by weight and which can be decomposed into simpler substances by the suitable chemical methods. The properties of a compound are completely different from those of its constituent elements.

e.g. HCl, H_2O , H_2SO_4, $HClO_4$, HNO_3 etc.

(b) Mixtures : A material which contains more than one type of substance and which are mixed in any ratio by weight are known as mixtures. The properties of a mixture are same as the property individual components. The components of a mixture can be separated by simple physical methods.

Mixtures are classified into two types :

(i) Homogeneous mixtures : The mixtures in which all the components are present **uniformly** are called as homogeneous mixtures. Components of a mixture are present in single phase.

e.g. Water + Salt, Water + Sugar, Water + alcohol,

(ii) Heterogenous mixtures : The mixtures in which all the components are present **non-uniformly** are called as Heterogenous mixture.

e.g. Water + Sand, Water + Oil, blood, petrol etc.

S.I. UNITS (INTERNATIONAL SYSTEM OF UNITS)

Different types of units of measurements have been in use in different parts of the world e.g. kilograms, pounds etc. for mass ; miles, furlongs, yards etc. for distance.

To have a common system of units throughout the world. French Academy of Science, in 1791, introduced a new system of measurements called metric system in which the different units of a physical quantity are related to each other as multiples of powers of 10, e.g. 1 km = 10^3 m, 1cm = 10^{-2} m etc. This system of units was found to be so convenient that scientists all over the world adopted this system for scientific data.

(A) Seven Basic Units

The seven basic physical quantities in the International System of Units, their symbols, the names of their units (called the base units) and the symbols of these units are given in Table.

TABLE : SEVEN BASIC PHYSICAL QUANTITIES AND THEIR S.I. UNITS

Physical Quantity	Symbol	S.I. Unit	Symbol
Length	ℓ	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Temperature	T	kelvin	K
Luminous intensity	I_u	candela	cd
Amount of the substance	n	mole	mol

(B) Prefixes Used With Units

The S.I. system recommends the multiples such as 10^3, 10^6, 10^9 etc. and fraction such as 10^{-3}, 10^{-6}, 10^{-9} etc., i.e. the powers are the multiples of 3. These are indicated by special prefixes. These along with some other fractions or multiples in common use, along with their prefixes are given below in Table and illustrated for length (m)

TABLE : SOME COMMONLY USED PREFIXES WITH THE BASE UNITS.

Prefix	Symbol	Multiplication Factor	Example
deci	d	10^{-1}	1 decimetre (dm) = 10^{-1} m
centi	c	10^{-2}	1 centimetre (cm) = 10^{-2} m
milli	m	10^{-3}	1 millimetre (mm) = 10^{-3} m
micro	μ	10^{-6}	1 micrometre (μm) = 10^{-6} m
nano	n	10^{-9}	1 nanometre (nm) = 10^{-9} m
pico	p	10^{-12}	1 picometre (pm) = 10^{-12} m
femto	f	10^{-15}	1 femtometre (fm) = 10^{-15} m
atto	a	10^{-18}	1 attometre (am) = 10^{-18} m
deca	da	10^{1}	1 dekametre(dam) = 10^{1} m
hecto	h	10^{2}	1 hectometre (hm) = 10^{2} m
kilo	k	10^{3}	1 kilometre (km) = 10^{3} m
mega	M	10^{6}	1 megamerte(Mm) = 10^{6} m
giga	G	10^{9}	1 gigametre (Gm) = 10^{9} m
tera	T	10^{12}	1 teramerte (Tm) = 10^{12} m
peta	P	10^{15}	1 petametre (Pm) = 10^{15} m
exa	E	10^{18}	1 exametre (Em) = 10^{18} m

As volume is very often expressed in litres, it is important to note that the equivalence in S.I. units for volume is as under: **1 litre (1L) = 1 dm³ = 1000 cm³**

and **1 millilitre (1mL) = 1 cm³ = 1cc**

(C) SOME IMPORTANT UNIT CONVERSIONS

1. Length :
1 mile = 1760 yards
1 yard = 3 feet
1 foot = 12 inches
1 inch = 2.54 cm
1Å = 10^{-10}m or 10^{-8} cm

2. Mass :
1 Ton = 1000 kg
1 Quintal = 100 kg
1 kg = 2.205 Pounds (lb)
1 kg = 1000 g

		1 gram	=	1000 milli gram
		1 amu	=	1.67×10^{-24} g
3.	**Volume :**	1 L	=	1 dm^3 = $10^{-3} m^3$ = 10^3 cm^3 = 10^3 mL = 10^3 cc
		1 mL	=	1 cm^3 = 10^{-6} m^3
			=	1 cc
4.	**Energy :**	1 calorie	=	4.184 joules $\simeq$ 4.2 joules
		1 joule	=	10^7 ergs
		1 litre atmosphere (L-atm)	=	101.3 joule
		1 electron volt (eV)	=	1.602×10^{-19} joule
5.	**Pressure :**	1 atmosphere (atm)	=	760 torr
			=	760 mm of Hg
			=	76 cm of Hg
			=	1.01325×10^5 pascal (Pa)
			=	1.01325×10^5 N/m^2

6. **Temperature :** ℃ + 273.15 = K ; $\frac{5}{9}$(°F – 32) = °C

Some More Prefixes :

Semi	=	$\frac{1}{2}$	Mono	=	1
Sesqui	=	$\frac{3}{2}$ = 1.5	Di or Bi	=	2
Tri	=	3	Tetra	=	4
Penta	=	5	Hexa	=	6
Hepta	=	7	Octa	=	8
Nona	=	9	Deca	=	10
Undeca	=	11	Do deca	=	12
Trideca	=	13	Tetra deca	=	14
Pentadeca	=	15	Hexa deca	=	16
Hepta deca	=	17	Octa deca	=	18
Nonadeca	=	19	Eicoso/Icoso	=	20

POINTS TO REVISE

- The unit named after a scientist is started with a small letter and not with a capital letter e.g. unit of force is written as newton and not as Newton.

  Likewise unit of heat and work is written as joule and not as Joule.
- Symbols of the units do not have a plural ending like 's'. For example we have 10 cm and not 10 cms.
- Words and symbols should not be mixed e.g. we should write either joules per mole or J mol^{-1} and not joules mol^{-1}
- Prefixes are used with the basic units e.g. kilometer means 1000 m (because meter is the basic unit).

  Exception. Though kilogram is the basic unit of mass, yet prefixes are used with gram because in kilogram, kilo is already a prefix.
- A unit written with a prefix and a power is a power for the complete unit e.g. cm^3 means $(\text{centimeter})^3$ and not centi $(\text{meter})^3$.

MOLE CONCEPT

In SI Units we represent mole by the symbol 'mol'. It is defined as follows :

(i) A mole is the amount of a substance that contains as many entities (atoms, molecules or other particles) as there are atoms in exactly 12g of the carbon - 12 isotope.

It may be emphasised that the mole of a substance always contains the same number of entities, no matter what the substance may be. In order to determine this number precisely, the mass of a carbon-12 atom was determined by a mass spectrometer and found to be equal to 1.992648×10^{-23}g Knowing that 1 mole of carbon weighs 12g, the number of atoms in it is equal to :

$$\frac{12\text{g/mol } C^{12}}{1.992648 \times 10^{-23}\text{g/ } C^{12}\text{ atom}} = 6.0221367 \times 10^{23} \text{ atoms/mol}$$

(ii) In a simple way, we can say that mole has 6.0221367×10^{23} entities (atoms, molecules or ions etc.)

The number of entities in 1 mole is so important that it is given a separate name and symbol, known as '**Avogadro constant**' denoted by N_A .

Here entities may represent atoms, ions, molecules or other subatomic entities. Chemists count the number of atoms and molecules by weighing. In a reaction we require these particles (atoms, molecules and ions) in a definite ratio. We make use of this relationship between numbers and masses of the particles for determining the stoichiometry of reactions .

Formula to get moles are following :

(i) $\boxed{\text{Number of moles (n)} = \frac{\text{weight (g)}}{\text{molar mass}}}$

Where molar mass = gram atomic mass or gram molecular mass or gram ionic mass

(ii) $\boxed{\text{Number of moles (n)} = \frac{V_{(L)}}{22.4}}$ (Where V = Volume of gas in L at NTP or STP)

(iv) $\boxed{\text{Number of moles (n)} = \frac{N}{N_A}}$ (Where N = Number of particles)

$$\text{No. of moles of atoms} = \frac{\text{number of atoms}}{N_A} \text{ and No. of moles of molecules} = \frac{\text{number of molecules}}{N_A}$$

SOME RELATED DEFINITIONS :

Atomic Mass (Relative Atomic Mass)

It is defined as the number which indicates how many times the mass of one atom of an element is heavier in comparison to $1/12^{th}$ part of the mass of one atom of C^{12}.

Atomic mass unit (amu) : The quantity $1/12^{th}$ mass of an atom of C^{12} is known as atomic mass unit.

Since mass of 1 atom of $C^{12} = 1.9924 \times 10^{-23}$ g

$$\therefore 1/12^{th} \text{ part of the mass of 1 atom} = \frac{1.9924 \times 10^{-23}\text{g}}{12} = 1.67 \times 10^{-24} \text{ g} = 1 \text{ a.m.u.} = \frac{1}{6.023 \times 10^{23}}$$

It may be noted that the atomic masses as obtained above are the relative atomic masses and not the actual masses of the atoms. These masses on the atomic mass scale are expressed in terms of atomic mass units (abbreviated as amu). Today, 'amu' has been replaced by 'u' which is known as unified mass.

One atomic mass unit (amu) is equal to $1/12^{th}$ of the mass of an atom of C^{12} isotope.

Thus the atomic mass of hydrogen is 1.008 amu while that of oxygen is 15.9994 amu (or taken as 16 amu).

Gram Atomic Mass (or Mass of 1 Gram Atom)

When numerical value of atomic mass of an element is expressed in grams then the value becomes gram atomic mass.

gram atomic mass = mass of 1 **gram atom** = mass of 1 **mole atom**

= mass of N_A atoms = mass of 6.023×10^{23} atoms.

Ex. gram atomic mass of oxygen = mass of 1 **g atom** of oxygen = mass of 1 **mol atom** of oxygen.

$$= \text{mass of } N_A \text{ atoms of oxygen.} = \left(\frac{16}{N_A} g\right) \times N_A = 16 \text{ g}$$

Molecular Mass (Relative Molecular Mass)

The number which indicates how many times the mass of one molecule of a substance is heavier in comparison to $1/12^{th}$ part of the mass of an atom of C^{12}.

Gram Molecular Mass (Mass of 1 Gram Molecule)

When numerical value of molecular mass of the substance is expressed in grams then the value becomes gram molecular mass.

gram molecular mass = mass of 1 **gram molecule** = mass of 1 **mole molecule**

= mass of N_A molecules = mass of 6.023×10^{23} molecules

Ex. gram molecular mass of H_2SO_4 = mass of 1 **gram molecule** of H_2SO_4

= mass of 1 **mole molecule** of H_2SO_4

= mass of N_A molecules of H_2SO_4

$$= \left(\frac{98}{N_A} g\right) \times N_A = 98 \text{ g}$$

Actual Mass

The mass of one atom or one molecule of a substance is called as actual mass.

Ex. (i) Actual mass of O_2 = 32 amu = $32 \times 1.67 \times 10^{-24}$ g $\rightarrow$ Actual mass

(ii) Actual mass of H_2O = (2 + 16) amu = $18 \times 1.67 \times 10^{-24}$g = 2.99×10^{-23} g

Atomicity – Total number of atoms in **a molecule** of elementary substance is called as atomicity.

Ex.

Molecule	Atomicity
H_2	2
O_2	2
O_3	3
NH_3	4

RELATION BETWEEN MOLECULAR WEIGHT AND VAPOUR DENSITY :

Vapour density (V.D) : Vapour density of a gas is the ratio of densities of gas & hydrogen at the same temperature & pressure.

$$\text{Vapour Density (V.D)} = \frac{\text{Density of gas}}{\text{Density of hydrogen}} = \frac{d_{gas}}{d_{H_2}} \qquad \left\{ d = \frac{m(\text{mass})(g)}{V(\text{Volume})(mL)} \right.$$

$$\text{V.D} = \frac{(m_{gas})\text{for certain V litre volume}}{(m_{H_2})\text{for certain V litre volume}}$$

If N molecules are present in the given volume of a gas and hydrogen under similar condition of temperature and pressure.

$$\text{V.D.} = \frac{(m_{gas})\text{ of N molecules}}{(m_{H_2})\text{of N molecules}} = \frac{(m_{gas})\text{ of 1 molecule}}{(m_{H_2})\text{of 1 molecule}} = \frac{\text{Molecular mass of gas}}{2}$$

$\therefore$ $\boxed{\textbf{Molecular mass of gas } (M_w) = 2 \times \text{V.D}}$

RELATION BETWEEN MOLAR MASS (M_W) & VOLUME :

$$\text{At STP. } M_{W.} = 2 \times \text{V.D} = 2 \times \frac{d_{gas}}{d_{H_2}} = 2 \times \frac{(m_{gas})\text{for certain V litre volume}}{(m_{H_2})\text{for certain V litre volume}}$$

$$\text{or} \quad M_W = 2 \times \frac{\text{mass of 1 litre gas}}{\text{mass of 1 litre } H_2}$$

$$\text{or} \quad M_W = 2 \times \frac{\text{Mass of 1 litre gas}}{0.089g}$$

$$M_W(g) = 22.4 \times \text{mass of 1 litre gas}$$

$$d_{H_2} = 0.000089 \frac{g}{mL} = \frac{m}{V} = \frac{m}{1000mL}$$

$V = 1\text{ L} = 1000\text{ mL}$

then $m_{H_2} = 0.089g$

$\boxed{\textbf{M}_\textbf{W}\textbf{(g) = Mass of 22.4 litre gas}}$ **or** $\boxed{\textbf{M}_\textbf{w}\textbf{(g)} \equiv \textbf{22.4 litre (at STP)}}$

GRAM MOLECULAR VOLUME (GMV)

At NTP, the volume of 1 mole of gaseous substance is 22.4 litre is called as gram molecular volume.

At NTP, d_{H_2} = 0.000089 g/mL = mass/volume = mass/1000 mL

If volume = 1 L = 1000 mL then mass = 0.089 g

$\because$ 0.089g H_2 occupies = 1 L at STP

$\therefore$ 1 g H_2 occupies = $\frac{1\text{litre}}{0.089}$ at STP

$\therefore$ 2 g or 1 mol H_2 occupies = $\frac{1\text{litre}}{0.089} \times 2 = 22.4$ L at STP

1 mole of any gaseous substance occupy 22.4 litre of volume at NTP or STP

$\boxed{\textbf{1 mol} \equiv \textbf{22.4 L (at STP)}}$

POINTS TO REVISE

- Term molar mass means mass of 1 mol particles.
- Vapour density is calculated with respect to H_2 gas under similar conditions of temperature and pressure.
- Relative density = $\frac{\text{Density of gas A}}{\text{Density of gas B}}$
- Specific gravity : It is density of material with respect to water.
- Vapour density, relative density and specific gravity are ratios so they are unitless.
- The term STP means 273.15 K (0°C) and 1 bar pressure. The term NTP means 273.15 K (0°C) and 1 atm.

PERCENTAGE COMPOSITION, EMPIRICAL FORMULA & MOLECULAR FORMULA

Percentage formula (% by mass)

(In a molecule or compound) Mass $\%$ of an element $= \frac{\text{Number of atom (Atomicity)} \times \text{atomic mass}}{\text{molecular mass}} \times 100$

If number of atom =1 : Molecular mass = **minimum molecular mass**

Empirical Formula

The empirical formula of a compound express the *simplest whole number ratio of atoms* of various elements present in 1 molecule of the compound.

Ex.					
	Molecular Formula →	H_2O_2	CH_4	C_2H_6	$C_2H_4O_2$
		2 : 2	1 : 4	2 : 6	2 : 4 : 2
		1 : 1	1 : 4	1 : 3	1 : 2 : 1
	Empirical Formula →	$\boxed{HO}$	$\boxed{CH_4}$	$\boxed{CH_3}$	$\boxed{CH_2O}$

Molecular Formula

The molecular formula of a compound represents the **actual number of atoms** present in 1 molecule of the compound i.e. it shows the real formula of its 1 molecule.

Relationship between Empirical & Molecular Formula

Molecular Formula = n × Empirical Formula

[Where n = natural no. (1, 2, 3,.........)]

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

Determination of Empirical Formula

Following steps are involved to determine the empirical formula of the compounds –

(1) First of all find the % by weight of each element present in 1 molecule of the compound

(2) The % by weight of each element is divided by its atomic weight. It gives atomic ratio of elements present in the compounds.

(3) Atomic ratio of each element is divided by the minimum value of atomic ratio as to get simplest ratio of atoms.

(4) If the value of simplest atomic ratio is fractional then raise the value to the nearest whole number.

or Multiply with suitable coefficient to convert it into nearest whole number.

(5) Write the Empirical formula as we get the simplest ratio of atoms.

Illustrations

Illustration Find out percentage composition of each element present in glucose ?

Solution % of C = $\frac{12\times6}{180}\times100 = 40\%$

% of H = $\frac{12\times1}{180}\times100 = 6.67\%$

% of O = $\frac{16\times6}{180}\times100 = 53.33\%$

Illustration In a compound x is 75.8% and y is 24.2% by weight present. If atomic weight of x and y are 24 and 16 respectively. Then calculate the empirical formula of the compound.

Solution

Elements	%	Atomic weight	$\frac{\%}{\text{Atomic weight}}$	Simplest ratio	Ratio
x	75.8%	24	$\frac{75.8}{24}=3.1$	$\frac{3.1}{1.5}=2$	2
y	24.2%	16	$\frac{24.2}{16}=1.5$	$\frac{1.5}{1.5}=1$	1

Empirical formula = x_2y

Illustration In a compound Carbon is 52.2%, Hydrogen is 13%, Oxygen is 34.8% are present and molecular mass of the compound is 92. Calculate molecular formula of the compound ?

Solution

Elements	%	Atomic weight	$\frac{\%}{\text{Atomic weight}}$	Simplest ratio	Ratio
C	52.2	12	$\frac{52.2}{12}$ =4.35 = 4.4	$\frac{4.4}{2.2}=2$	2
H	13	1	$\frac{13}{1}=13$	$\frac{13}{22}$=5.9	6
O	34.8	16	$\frac{34.8}{16}=2.2$	$\frac{2.2}{2.2}=1$	1

Empirical formula = C_2H_6O

Empirical formula mass = $12\times2 + 16 + 6 = 46$

$$n = \frac{\text{Molecular formula mass}}{\text{Empirical formula mass}} = \frac{92}{46} = 2$$

molecular formula = $2\times(C_2H_6O) = C_4H_{12}O_2$

STOICHIOMETRY BASED CONCEPT (PROBLEMS BASED ON CHEMICAL REACTION)

One of the most important aspects of a chemical equation is that when it is written in the balanced form, it gives quantitative relationships between the various reactants and products in terms of moles, masses, molecules and volumes. This is called stoichiometry (Greek word, meaning 'to measure an element'). For example, a balanced chemical equation along with the quantitative information conveyed by it is given below:

Ex.

$CaCO_{3(s)}$	+	$2HCl_{(aq)}$	$\longrightarrow$	$CaCl_{2(aq)}$	+	$H_2O_{(l)}$	+	$CO_{2(s)}$
1 Mol		2 Mol		1 Mol		1 Mol		1 Mol
$40 + 12 + 3\times16$		$2(1 + 35.5)$		$40 + 2\times35.5$		$2\times1 + 16$		$12 + 2\times16$
= 100 g		= 73 g		= 111 g		= 18 g		= 44 g or 22.4 L at STP

Thus,

(i) 1 mole of calcium carbonate reacts with 2 moles of hydrochloric acid to give 1 mole of calcium chloride, 1 mole of water and 1 mole of carbon dioxide.

(ii) 100 g of calcium carbonate react with 73 g hydrochloric acid to give 111 g of calcium chloride, 18 g of water and 44 g (or 22.4 litres at STP) of carbon dioxide.

Ex. $\boxed{1 \quad\quad 3 \quad\quad 2}$ Stoichiometric coefficient

$N_{2(g)}$	+	$3H_{2(g)}$	$\rightarrow$	$2NH_{3(g)}$
1 mol	+	3 mol	$\rightarrow$	2 mol
22.4 L	+	3×22.4 L	$\rightarrow$	2×22.4 L (at STP)
1 L	+	3 L	$\rightarrow$	2 L
1000 mL	+	3000 mL	$\rightarrow$	2000 mL
1 mL	+	3 mL	$\rightarrow$	2 mL
28 g	+	6 g	$\rightarrow$	34 g (According to the law of conservation of mass)

- Gram can not be represented according to stoichiometry.

The quantitative information conveyed by a chemical equation helps in a number of calculations. The problems involving these calculations may be classified into the following two different types :

(a) Single reactant based (b) More than one reactant based

(A) SINGLE REACTANT BASED :

(1) Mass - Mass Relationships i.e. mass of one of the reactants or products is given and the mass of some other reactant or product is to be calculated.

(2) Mass - Volume Relationships i.e. mass/volume of one of the reactants or products is given and the volume/mass of the other is to be calculated.

(3) Volume - Volume Relationships i.e.volume of one of the reactants or the products is given and the volume of the other is to be calculated.

General method : Calculations for all the problems of the above types consists of the following steps :-

(i) Write down the balanced chemical equation.

(ii) Write the relative number of moles or the relative masses (gram atomic or molecular masses) of the reactants and the products below their formula.

(iii) In case of a gaseous substance, write down 22.4 litres at STP below the formula in place of 1 mole

(iv) Apply unitary method to make the required calculations.

Quite often one of the reactants is present in larger amount than the other as required according to the balanced equation. The amount of the product formed then depends upon the reactant which has reacted completely. This reactant is called the limiting reactant. The excess of the other is left unreacted.

Combustion reaction : (Problem based on combustion reactions) :

For balancing the combustion reaction: First of all balance C atoms, Then balance H atom, Finally balance Oxygen atom.

For Example : Combustion reaction of C_2H_6 : $C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$ (skeleton equation)

First balance C atoms $C_2H_6 + O_2 \longrightarrow 2CO_2 + H_2O$

Now balance H atoms $C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$

Now balance Oxygen atoms $C_2H_6 + {}^{7}\!/_{2}\, O_2 \longrightarrow 2CO_2 + 3H_2O$

Illustrations

TYPE-I (INVOLVING MASS-MASS RELATIONSHIP)

Illustration How much iron can be theoretically obtained in the reduction of 1 kg of Fe_2O_3

Solution

$$\boxed{\begin{array}{cccc} 1 & 3 & 2 & 3 \end{array}}$$

$$Fe_2O_3 + 3H_2 \longrightarrow 2Fe + 3H_2O$$

$$n = \frac{\text{weight}}{M_w} = \frac{1000}{160}\text{mol}$$

The equation shows that 2 mol of iron are obtained from 1 mol of ferric oxide.

Hence, the obtained no. of moles of Fe = $\frac{2 \times 1000}{160}$ =12.5 mol = $\frac{\text{weight}}{\text{Atomic weight}} = \frac{\text{weight}}{56}$

Weight of iron obtained = 12.5 × 56 g = 700 g

Illustration What amount of silver chloride is formed by the action of 5.850 g of sodium chloride on an excess of silver nitrate?

Solution

$$\boxed{\begin{array}{cccc} 1 & 1 & 1 & 1 \end{array}}$$

$$NaCl + AgNO_3 \longrightarrow AgCl + NaNO_3$$

$$n = \frac{\text{weight}}{M_w} = \frac{5.85}{58.5} = 0.1 \text{ mol}$$

1 mol of AgCl is obtained from 1 mol of NaCl
Hence, the number of moles of AgCl obtained with 0.1 mol of NaCl = 0.1 mol

$$\because \quad n = \frac{\text{weight}}{M_W} \Rightarrow 0.1 \text{ mol} = \frac{\text{weight}}{M_w} = \frac{\text{weight}}{143.5} \Rightarrow \text{weight} = 0.1 \times 143.5 \text{ g} = 14.35 \text{ g.}$$

TYPE-II (MASS - VOLUME RELATIONSHIP)

Illustration For complete combustion of 3g ethane the required volume of O_2 & produced volume of CO_2 at STP will be.

Solution

$$\boxed{\begin{array}{cccc} 2 & 7 & 4 & 6 \end{array}}$$

$$2C_2H_{6(g)} + 7O_{2(g)} \longrightarrow 4CO_{2(g)} + 6H_2O_{(g)}$$

$$n = \frac{\text{weight}}{M_w} = \frac{3}{30} = \frac{1}{10} = 0.1 \text{ mol}$$

(a) Required moles of $O_2 = \frac{7}{2} \times 0.1 = 0.35$ mol

volume of O_2 at STP = 0.35 × 22.4 = 7.84 L

(b) Produced moles of $CO_2 = \frac{4}{2} \times 0.1 = 0.2$ mol

volume of CO_2 at STP = 0.2 ×22.4 = 4.48 L

Illustration In the following reaction, if 10 g of H_2 reacts with N_2. What will be the volume of NH_3 at STP.

$N_2 + 3H_2 \longrightarrow 2NH_3$

Solution

1	3	2

$$\underset{}{N_2} + \underset{10g}{3H_2} \longrightarrow 2NH_3$$

$$n = \frac{\text{weight}}{M_w} = \frac{10}{2} = 5 \text{ mol.}$$

Produced moles of $NH_3 = \frac{2}{3} \times 5 = \frac{10}{3}$, Volume of NH_3 at STP $= \frac{10}{3} \times 22.4 = 74.67$ litre

TYPE-III (VOLUME-VOLUME RELATIONSHIP)

Illustration For complete combustion of 1.12 L of butane (C_4H_{10}), the produced volume of $H_2O_{(g)}$ & $CO_{2(g)}$ at STP will be.

Solution

1	13/2	4	5

$$\underset{1.12L}{C_4H_{10(g)}} + 13/2\,O_{2(g)} \longrightarrow 4CO_{2(g)} + 5H_2O_{(g)}$$

Volume of $H_2O_{(g)}$ at STP = 5 × 1.12 = 5.6 L

Volume of $CO_{2(g)}$ at STP = 4 × 1.12 = 4.48 L

Illustration For complete combustion of 5 mol propane (C_3H_8). The required volume of O_2 at STP will be.

Solution For C_3H_8, the combustion reaction is

1	5	3	4

$$\underset{5\,mol}{C_3H_{8(g)}} + 5O_{2(g)} \longrightarrow 3CO_{2(g)} + 4H_2O_{(l)}$$

Required moles of $O_2 = 5 \times 5 = 25 \text{ mol} = \frac{V}{22.4}$

volume of O_2 gas at STP = 25 × 22.4 = 560 L

(B) MORE THAN ONE REACTANT BASED :

Limiting reagent (L.R.) concept

Limiting Reagent (L.R.) : The reactant which is completely consumed in a reaction is called as limiting reagent.

Ex.

	1		2		1		2	⟵ Stoichiometry
	A	+	2B	⟶	C	+	2D	
given	3 mol		9mol					
	3 – 3 = 0 mol		9 – 6 = 3 mol		3 mol		6 mol	

L.R. = A

$$X = \frac{\text{given value (may moles, volume, or molecules)}}{\text{Stoichiometry Co-efficient}}$$

Reactants having least value of x are limiting reagents.

Ex. A + 2B $\rightarrow$ P

$3/1 = 3$ $9/2 = 4.5$

$3 < 4.5$ So A is L.R.

Illustrations

Illustration $A + 5B \longrightarrow C + 3D$ In this reaction which is a L.R.

Solution $A + 5B \longrightarrow C + 3D$ Given 10 mol of A and 10 mol of B.

10 mol 10 mol

$x = \frac{10}{1} = 10$ $x = \frac{10}{5} = 2$

$2 < 10$ So B is L.R.

Illustration $H_{2(g)} + \frac{1}{2} O_{2(g)} \longrightarrow H_2O_{(g)}$; In the above reaction what is the volume of water vapour at STP.

Given 4 g of H_2 and 32 g of O_2

Solution

$$\begin{array}{lll} 1 & 1/2 & 1 \\ H_{2(g)} + & \frac{1}{2} O_{2(g)} \longrightarrow & H_2O_{(g)} \\ 4\text{ g} & 32\text{ g} & \end{array}$$

For H_2 For O_2

$n = \frac{4}{2} = 2$ mol $n = \frac{32}{32} = 1$ mol $\frac{2}{1} = 2$ $\frac{1}{\frac{1}{2}} = 2$ mol

Moles of $H_2O_{(g)}$ = 2 mol = $\frac{V}{22.4}$ $2 = 2$ So Both H_2 & O_2 are L.R.

Volume of $H_2O_{(g)}$ at STP = $22.4 \times 2 = 44.8$ litre

Illustration At NTP, In a container 100 mL N_2 and 100 mL of H_2 are mixed together. Then find out the produced volume of NH_3.

Solution Balanced equation will be $N_2 + 3H_2 \longrightarrow 2NH_3$.

Given 100mL 100mL

For determination of Limiting reagent. Now divide the given quantities by stoichiometry coefficients

$\frac{100}{1} = 100$ $\frac{100}{3} = 33.3$ (Limiting reagent)

In this reaction H_2 is limiting reagent so reaction will proceed according to H_2.

As per stoichiometry from 3 mL of H_2 produces ; volume of NH_3 = 2 mL

That is from 100 mL of H_2 produced volume of $NH_3 = \frac{2}{3} \times 100 = 66.6$ mL

CONCENTRATION TERMS :

The concentration of a solution is the amount of solute dissolved in a known amount of the solvent or solution. Solution can be described as dilute or concentrated solution as per their concentration. A dilute solution has a very small quantity of solute while concentrated solution has a large quantity of solute in solution. Various concentration terms are as follows.

Mass percentage :

It may be defined as the number of parts of mass of solute per hundred parts by mass of solution.

$$\textbf{\% by mass}\left(\frac{w}{W}\right): = \frac{\text{wt. of solute}}{\text{wt. of solution}} \times 100$$

[X % by mass means 100 gm solution contains X gm solute ; $\therefore$ (100 – X) gm solvent]

Mass-volume percentage (W/V %) :

It may be defined as the mass of solute present in 100 cm^3 of solution. For example, If 100 cm^3 of solution contains 5 g of sodium hydroxide, than the mass-volume percentage will be 5% solution.

$$\%\left(\frac{w}{V}\right) = \frac{\text{wt. of solute}}{\text{volume of solution}} \times 100 \text{ [for liq. solution]}$$

[X % $\left(\frac{w}{V}\right)$ means 100 ml solution contains X gm solute]

Volume Percent :

It can be represented as % v/v or % volume and used to prepare such solutions in which both components are in liquids state. It is the number of parts of by volume of solute per hundred parts by volume of solution. Therefore,

$$\%\left(\frac{v}{V}\right) = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$$

Mole % $= \dfrac{\text{Moles of solute}}{\text{Total moles}} \times 100$

- For gases **% by volume** is same as **mole %**

Mole Fraction (X) :

Mole fraction may be defined as the ratio of number of moles of one component to the total number of moles of all the components (solute and solvent) present in solution. It is denoted by letter X and the sum of all mole fractions in a solution always equals one.

$$\text{Mole fraction (X)} = \frac{\text{Moles of solute}}{\text{Total moles}}$$

Mole fraction does not depend upon temperature and can be extended to solutions having more than two components.

Molarity (M) :

Molarity is most common unit for concentration of solution. It is defined as the number of moles of solute present in one litre or one dm^3 of the solution or millimol of solute present in one mL of solution.

$$\text{Molarity (M)} = \frac{\text{Mole of solute}}{\text{volume of solution in litre}}$$

Molality (m) : The number of gram mole of the solute present in 1000 g of the solvent is known as molality of solution. It represented by letter 'm'.

$$\text{Molality (m)} = \frac{\text{Moles of solute}}{\text{Mass of solvent (in kg)}}$$

The unit of molality is mol/kg and it does not effect by temperature.

Parts per million (ppm) : The very low concentration of solute in solution can be expressed in ppm. It is the numbers of parts by mass of solute per million parts by mass of the solution.

$$\text{Parts per million (ppm)} = \frac{\text{Mass of solute}}{\text{Mass of solvent}} \times 10^6 \cong \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^6$$

◆ Get yourselves very much confortable in their inter conversion. It is very handy.

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 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 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 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	$X_A = \frac{\text{Mole of A}}{\text{Mole of A} + \text{Mole of B} + \text{Mole of C} + ...}$ $X_B = \frac{\text{Mole of B}}{\text{Mole of A} + \text{Mole of B} + \text{Mole of C} + ...}$	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 litre of solution.
Molality	$m = \frac{\text{Mass of solute} \times 1000}{\text{Molar mass of solute} \times \text{Mass of solvent(g)}}$	Moles of solute in one kg of solvent

Ex. ***Calculate the mole fractions of the components of the solution composed by 92 g glycerol and 90 g water ?(M (water) = 18 ; M (glycerol) = 92)***

Ans. Moles of water = 90 g / 18 g = 5 mol water

Moles of glycerol = 92 g / 92 g = 1 mol glycerol

Total moles in solution = 5 + 1 = 6 mol

Mole fraction of water = 5 mol / 6 mol = 0.833

Mole fraction of glycerol = 1 mol / 6 mol = 0.167

Ex. ***What will be the Molarity of solution when water is added to 10 g $CaCO_3$ to make 100 mL of solution?***

Ans. Mol of $CaCO_3$ = 10 / 100 = 0.1

Molarity = Mole of solute / Volume of solution (L) = 0.10 mol / 0.10 L

Therefore ; Molarity of given solution = 1.0 M

Ex. ***Calculate the molality of a solution containing 20 g of sodium hydroxide (NaOH) in 250 g of water?***

Ans. Moles of sodium hydroxide = 20 / 40 = 0.5 mol NaOH

250 gm = 0.25 kg of water

Hence molality of solution = Mole of solute / Mass of solvent (kg)= 0.5 mol / 0.25 kg

or Molality(m) = 2.0 m

Ex. ***Calculate the grams of copper sulphate ($CuSO_4$) needed to prepare 250.0 mL of 1.00 M $CuSO_4$?***

Ans. Moles of $CuSO_4$ = $M \times V = 1 \times \dfrac{250}{1000}$

Molar mass of copper sulphate = 159.6 g/mol

Hence Mass of copper sulphate (gm) = Moles of $CuSO_4 \times$ Molar mass of copper sulphate.

$$= 1 \times \frac{250}{1000} \times 159.6 \text{ g/mol}$$

= 39.9 gm of Copper sulphate

Ex. ***How many grams of H_2SO_4 are present in 500 ml of 0.2M H_2SO_4 solution ?***

Ans. $M = \dfrac{\text{moles}}{\text{vol.}} \Rightarrow$ moles of H_2SO_4 $= M \times V = 0.2 \times \dfrac{500}{1000} \text{L} = 0.1$

Mass of H_2SO_4 = $0.1 \times 98 = 9.8$ g

Ex. ***Calculate the ppm of mercury in water in given sample contain 30 mg of Hg in 500 ml of solution.***

Ans. Parts per million $= \dfrac{\text{Mass of solute} \times 10^6}{\text{Mass of solution}}$

Mass of Hg = 30 mg

Mass of water = 500/1 = 500g = 50×10^4 mg

(density = mass / volume ; density of water 1 g / ml) $w = \dfrac{v}{d}$

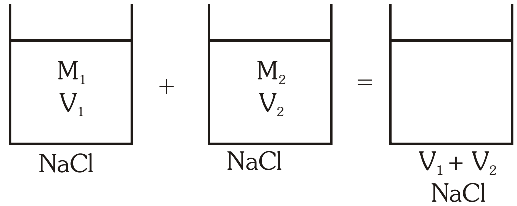
Therefore, ppm of mercury = $\dfrac{30 \times 10^6}{50 \times 10^4}$ = 60 ppm of mercury

MIXING OF SOLUTIONS :

It is based on law of conservation of moles.

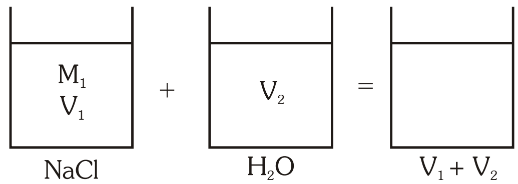
(i) Two solutions having same solute

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



(ii) Dilution Effect : When a solution is diluted, the moles of solute do not change but molarity changes while on taking out a small volume of solution from a larger volume, the molarity of solution do not change but moles change proportionately.

Final molarity = $\dfrac{M_1V_1}{V_1 + V_2}$



n-fold or n-times dilution

$\Rightarrow$ **Final volume = $V_1 + V_2 = n(V_1)$**

Ex. ***50 ml 0.2 M H_2SO_4 is mixed with 50 ml 0.3M H_2SO_4. Find molarity of final solution.***

Ans. $M_f = \dfrac{\text{Total moles of } H_2SO_4}{\text{Total volume}} = \dfrac{50 \times 0.2 \times 10^{-3} + 50 \times 10^{-3} \times 0.3}{(50 + 50) \times 10^{-3}} = \boxed{0.25\,M}$

Ex. ***Find final molarity in each case :***

Ans. **(i)** 500 ml 0.1 M HCl + 500 ml 0.2M HCl

$M_f = \dfrac{500 \times 0.1 + 500 \times 0.2}{500 + 500} = \boxed{0.15\,M}$

(ii) 50 ml 0.1M HCl + 150 ml 0.3MHCl + 300 ml H_2O

$M_f = \dfrac{50 \times 0.1 + 150 \times 0.3}{50 + 150 + 300} = \dfrac{50}{500} = 0.1$ M

(iii) 4.9g H_2SO_4 +250 ml H_2O + 250 ml 0.1 M H_2SO_4

$M_f = \dfrac{\dfrac{4.9}{98} + \dfrac{250}{1000} \times 0.1}{\left(\dfrac{250 + 250}{1000}\right)} = \dfrac{50 + 25}{500} = \boxed{0.15\,M}$

Ex. ***How much water should be added to 2M HCl solution to form 1 litre of 0.5 M HCl ?***

Ans. Let V be initial volume

Then mol of HCl = constant

$2 \times V = 1 \times 0.5 \Rightarrow V = 0.25$ L

Volume of water added = 1 – 0.25 = 0.75 L

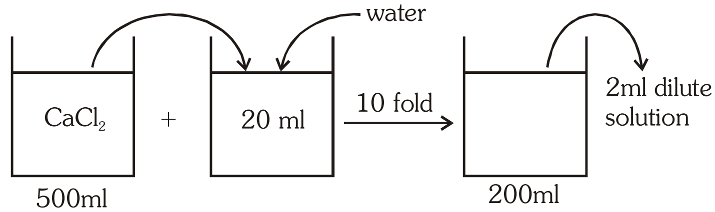
Ex. ***Find number of Na^+ & PO_4^{-3} ions in 250 ml of 0.2M Na_3PO_4 solution.***

Ans. $Na_3PO_4 + \text{aq.} \longrightarrow 3Na^+(aq) + PO_4^{-3}(aq)$ [Ionic compound when added to water ionize completely].

50 millimoles (m.m.) 150 mm 50 mm

No. of Na^+ ions = $150 \times 10^{-3} \times N_A$; No. of PO_4^{-3} ions = $50 \times 10^{-3} \times N_A$

Ex. ***1.11g $CaCl_2$ is added to water forming 500 ml of solution. 20 ml of this solution is taken and diluted 10 folds. Find moles of Cl^- ions in 2 ml of diluted solution.***



Ans. $\frac{1.11}{111} = 0.01$ mol $CaCl_2$

Moles of $CaCl_2$ in 20ml solution $= \frac{0.01}{500} \times 20 = \frac{0.01}{25}$

In 200 ml solution moles of $CaCl_2 = \frac{0.01}{25}$ [Note : Dilution does not change moles of solute]

In 2 ml of dilute solution moles of $CaCl_2 = \frac{\frac{0.01}{25}}{200} \times 2 = \frac{0.01}{2500} = 8 \times 10^{-6}$

$\therefore$ moles of $Cl^- = 2 \times 8 \times 10^{-6} = 1.6 \times 10^{-5}$

Ex. ***What volumes of 1M & 2M H_2SO_4 solution are required to produce 2L of 1.75M H_2SO_4 solution?***

Ans. Let XL be vol. of 1M solution.

$\therefore$ (2 – X)L is vol. of 2M solution.

Moles of $H_2SO_4 = 2 \times 1.75 = 1(X) + (2 - X)2$

$3.5 = 4 - X$; $X = 0.5$ L

i.e. 0.5L of 1M & 1.5 L of 2M solution required.

Ex. ***80g NaOH was added to 2L water. Find molality of solution if density of water = 1g/mL***

Ans. $m = \frac{\text{moles of NaOH}}{\text{mass of } H_2O} \times 1000 = \frac{80/40}{2 \times 1000} \times 1000 = \boxed{1\,\text{molal}}$

Ex. ***A 100g NaOH solution has 20g NaOH. Find molality.***

Ans. $m = \frac{20/40}{100-20} \times 1000 = \frac{500}{80} = \boxed{6.25\,\text{mol/kg}}$

Ex. ***Find molality of aqueous solution of CH_3COOH whose molarity is 2M and density d = 1.2 g/mL.***

Hint : $\boxed{\dfrac{\mathbf{1000 \times M}}{\mathbf{1000 \times d - MMs}}}$

where d = density in gL^{-1} , M = Molarity, m = molality, M_S = molar mass of solute.

Ans. $m = \frac{2}{1200 - 2 \times 60} \times 1000 = \boxed{1.85\,\text{m}}$

Ex. ***A solution is made by mixing 300 ml 1.5M $Al_2(SO_4)_3$ + 300 ml 2M $CaSO_4$ + 400 ml 3.5M $CaCl_2$ Find final molarity of (1) SO_4^{-2}, (2) Ca^{2+}, (3) Cl^- . [Assume complete dissociation of these compounds].***

Ans. (1) $[SO_4^{-2}]_f = \frac{\text{Total moles}}{\text{Total volume}} = \frac{300 \times 1.5 \times 10^{-3} \times 3 + 300 \times 2 \times 10^{-3}}{(300 + 300 + 400) \times 10^{-3}} = 1.95\text{M}$

(2) $[Ca^{+2}]_f = \dfrac{300 \times 2 + 400 \times 3.5}{1000} = 2M$

(3) $[Cl^-]_f = \dfrac{400 \times 3.5 \times 2}{1000} = 2.8M$

Ex. ***A solution has 80% $\frac{w}{w}$ NaOH with density $2gL^{-1}$. Find (a) Molarity (b) Molality of solution.***

Ans. Let V_{lit} be vol. of solution

$$\text{Mass of solute} = (d \times V) \times \frac{\left(\%\frac{w}{w}\right)}{100} = 2 \times V \times \frac{80}{100} = 1.6V$$

(a) $M = \dfrac{1.6V/40}{V} = \boxed{0.04M}$ (b) $m = \dfrac{1.6V/40}{2V - 1.6V} \times 1000 = \boxed{100\,\text{mol kg}^{-1}}$

Ex. ***4.450 g 100 per cent sulphuric acid was added to 82.20 g water and the density of the solution was found to be 1.029 g/cc at 25°C and 1 atm pressure. Calculate (a) the weight percent, (b) the mole fraction , (c) the mole percent, (d) the molality , (e) the molarity of sulphuric acid in the solution under these conditions.***

Ans. Sulphuric acid = 4.450 g , Water = 82.20 g $\Rightarrow$ Wt. of solution = 86.65 g

$\therefore$ Density of solution = 1.029 g/cc.

(a) Weight percent $= \dfrac{\text{wt. of solute}}{\text{wt. of solution}} \times 100 = \dfrac{4.450}{86.65} \times 100 = 5.14$

(b) Mole fraction :

Mole of solute $= \dfrac{\text{wt. of solute}}{\text{mol wt. of solute}} = \dfrac{4.45}{98} = 0.0454$

Mole of solvent $= \dfrac{82.20}{18} = 4.566$

Total moles in solution = 0.0454 + 4.566 = 4.6114

Mole fraction of solute $= \dfrac{0.0454}{4.6114} = 0.0098$

(c) Mole percent $= \dfrac{\text{moles of solute}}{\text{Total moles in solution}} \times 100$

$= \text{mole fraction of solute} \times 100 = 0.0098 \times 100 = 0.98$

(d) Molality $= \dfrac{\text{moles of solute}}{\text{mass of solvent (in gm)}} \times 1000$

$= \dfrac{0.0454 \times 1000}{82.2} = 0.552$

(e) Molarity = $\frac{\text{moles of solute}}{\text{litre of solution}}$

Volume of solution = $\frac{\text{Mass}}{\text{Density}} = \frac{86.65}{1.029}\text{ml}$

$= \frac{86.65}{1.029 \times 1000}\text{litre}$

Molarity = $\frac{0.0454}{\frac{86.54}{1.029 \times 1000}} = \frac{0.0454 \times 1000 \times 1.029}{86.65} = 0.539$

Ex. ***A solution of KCl has a density of 1.69 g mL^{-1} and is 67% by weight. Find the density of the solution if it is diluted so that the percentage by weight of KCl in the diluted solution is 30%.***

Ans. Let the volume of the KCl solution be 100 mL,

Weight of KCl solution = 100 × 1.69 = 169 g

100 g of solution contains = 67 g of KCl

169 g of solution = $\frac{67}{100} \times 169 = 113.23\text{g}$

Lex x mL of H_2O be added.

New volume of solution = (100 + x) mL

New weight of solution = (169 + x) g

(Since x mL of H_2O = x g of H_2O, d_{H_2O} = 1)

New percentage of the solution = 30%

% by weight = $\frac{\text{weight of solute} \times 100}{\text{weight of solution}}$

$30 = \frac{113.23}{(169 + x)} \times 100$

x = 208.43 mL = 208.43 g

New density $= \frac{\text{New weight of solution}}{\text{New volume of solution}}$

$= \frac{(169 + x)}{(100 + x)}$

$\frac{(169 + 208.43)}{(100 + 208.43)} = \frac{377.43}{308.43}$

∴ d = 1.224

SOME TYPICAL CONCENTRATION TERMS

PERCENTAGE LABELLING OF OLEUM :

Labelled as '% oleum' , it means maximum amount of H_2SO_4 that can be obtained from 100 gm of such oleum (mix of H_2SO_4 and SO_3) by adding sufficient water. For ex. 109 % oleum sample means, with the addition of sufficient water to 100 gm oleum sample 109 gm H_2SO_4 is obtained.

% labelling of oleum sample = (100 + x)%

x = mass of H_2O required for the complete conversion of SO_3 in H_2SO_4

Ex. Find the mass of free SO_3 present in 100 gm , 109 % oleum sample.

Sol. 109 % means, 9 gm of H_2O is requried.

$$SO_3 \quad + \quad H_2O \longrightarrow H_2SO_4$$

9gm

1/2mole 1/2mole

40gm

$\therefore$ Mass of free SO_3 = 40 gm , Mass of H_2SO_4 = 60 gm

Note: Work out, what are the maximum and minimum value of the % labelling.

Ex. ***Find the % labelling of 100 gm oleum sample if it contains 20 gm SO_3.***

Sol. % labelling of oleum sample = (100 + x)%

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

20gm

1/4mole 1/4mole

4.5gm

$\therefore$ % labelling of oleum sample = (100 + 4.5) % = 104.5%

VOLUME STRENGTH OF H_2O_2 SOLUTION :

Labelled as 'volume H_2O_2 , it means volume of O_2 (in litre) at STP that can be obtained from 1 litre of such a sample when it decomposes according to

$$H_2O_2 \rightarrow H_2O + \frac{1}{2}O_2$$

Volume Strength of H_2O_2 Solution = 11.35 ×molarity

Ex. ***Find the % w/v of "10 V" H_2O_2 solution-***

Sol. Molarity (M) of solution $= \dfrac{\text{volume strength}}{11.35} = \dfrac{10}{11.35}$

$$\%\left(\frac{w}{v}\right) = \frac{M \times \text{mol. wt. of solute}}{10} = \frac{10}{11.35} \times \frac{34}{10} = 3\%$$

EQUIVALENT WEIGHT

The equivalent weight of a substance is the number of parts by mass of the substance that combine with or displaces directly or indirectly 1.008 parts by mass of hydrogen or 8 parts by mass of oxygen or 35.5 parts by mass of chlorine or 108 parts by weight of Ag.

(a) Calculation of Equivalent Weight

(i) Equivalent weight $= \dfrac{\text{Atomic weight}}{\text{Valency factor}}$

(ii) Equivalent weight of ions $= \dfrac{\text{formula weight of ion}}{\text{Valency}}$

(iii) Equivalent weight of ionic compound = equivalent weight of cation + equivalent weight of anion

Ex. Equivalent weight of H_2SO_4 = Equivalent weight of H^++Equivalent weight of Anion (SO_4^{-2})

$= 1 + 48 = 49$

(iv) Equivalent weight of acid / base $= \dfrac{\text{Molecular weight}}{\text{Basicity/Acidity}}$

(v) Equivalent weight of salt $= \dfrac{\text{Molecular weight}}{\text{Total charge on cation or anion}}$

Ex. Na_2SO_4 (salt) $\rightarrow 2Na^+ + SO_4^{-2}$

Total charge on cation or anion is 2

molecular weight of Na_2SO_4 is $= (2 \times 23 + 32 + 16 \times 4) = 142$

Equivalent weight of Na_2SO_4 $= \dfrac{142}{2} = 71$

(vi) Equivalent weight of an oxidizing or reducing agent

$$= \frac{\text{Molecular weight of the substance}}{\text{Number of electrons gain/lost by one molecule}}$$

(b) Concept of gram equivalent and law of chemical equivalence :

Number of gram equivalent $= \dfrac{W_{(gram)}}{E}$

$$= \frac{W_{(gram)} \times \text{Valence factor}}{M}$$

$= n \times$ valence factor

According to it, in a reaction equal number of gram equivalents of reactants react to give equal number of gram equivalents of products.

For a reaction

$$aA + bB \longrightarrow cC + dD$$

Number of gram equivalents of A = Number of gram equivalents of B = Number of gram equivalents of C = Number of gram equivalents of D

(c) METHODS FOR DETERMINATION OF EQUIVALENT WEIGHT

(i) Hydrogen displacement method : This method is used for those elements which can evolve hydrogen from acids i.e. active metals.

$$\text{equivalent weight of metal} = \frac{\text{weight of metal}}{\text{weight of } H_2 \text{ gas (displaced)}} \times 1.008$$

(ii) Oxide formation method : A known mass of the element is changed into oxide directly or indirectly. The mass of oxide is noted.

Mass of oxygen = (Mass of oxide – Mass of element)

$$\text{equivalent weight of element} = \frac{\text{weight of element}}{\text{weight of oxygen}} \times 8$$

(iii) Chloride formation method : A known mass of the element is changed into chloride directly or indirectly. The mass of the chloride is determined.

$$\text{equivalent weight of element} = \frac{\text{weight of element}}{\text{weight of chlorine}} \times 35.5$$

(iv) Metal to metal displacement method : More active metal can displace less active metal from its salt's solution. The mass of the displaced metal bear the same ratio as their equivalent weights.

$$\frac{m_1}{m_2} = \frac{E_1}{E_2}$$

(v) Double decomposition method : This method is based on the following points -

(a) The mass of the compound reacted and the mass of product formed are in the ratio of their equivalent masses.

(b) The equivalent mass of the compound (electrovalent) is the sum of equivalent masses of its radicals.

(c) The equivalent mass of a radical is equal to the formula mass of the radical divided by its charge.

$$AB + CD \longrightarrow AD \text{ (ppt.)} + CB$$

$$\frac{\text{Mass of AB}}{\text{Mass of AD}} = \frac{\text{Equivalent mass of AB}}{\text{Equivalent mass of AD}} = \frac{\text{Equivalent mass of A} + \text{Equivalent mass of B}}{\text{Equivalent mass of A} + \text{Equivalent mass of D}}$$

(vi) Silver salt method : This method is used for finding the equivalent weight of carbonic (organic) acids. A known mass of the RCOOAg is changed into Ag through combustion. The mass of Ag is determined.

$$\frac{\text{Equivalent weight of RCOOAg}}{\text{Equivalent weight of Ag}} = \frac{\text{weight of RCOOAg}}{\text{weight of Ag}}$$

$$\text{equivalent weight of RCOOAg} = \frac{\text{weight of RCOOAg}}{\text{weight of Ag}} \times 108$$

(vii) By electrolysis : $\frac{w_1}{w_2} = \frac{E_1}{E_2}$

Where w_1 & w_2 are deposited weight of metals at electrodes and E_1 and E_2 are equivalent weight respectively.

METHODS FOR CALCULATION OF ATOMIC WEIGHT AND MOLECULAR WEIGHT

(a) Methods for Determination of Atomic Weight

(i) Atomic weight = equivalent weight × valency

(ii) Dulong and Petit's law - This law is applicable only for solids (except Be, B, Si, C)

Atomic mass × specific heat (Cal g^{-1} °C) ≈ 6.4

or atomic mass (approximate) = $\dfrac{6.4}{\text{specific heat}}$

(iii) Law of isomorphism : Isomorphous substances form crystals which have same shape and size and can grow in the saturated solution of each other.

Examples of isomorphous compounds -

(1) H_2SO_4 and K_2CrO_4 (2) $ZnSO_4.7H_2O$ and $FeSO_4.7H_2O$ and $MgSO_4.7H_2O$

(3) $KClO_4$ and $KMnO_4$ (4) $K_2SO_4.Al_2(SO_4)_3.24H_2O$ and $K_2SO_4.Cr_2(SO_4)_3.24H_2O$

Conclusions -

- Masses of two elements that combine with same mass of other elements in their respective compounds are in the ratio of their atomic masses.

$$\frac{\text{Mass of one elements (A) that combines with a certain mass of other element}}{\text{Mass of other element (B) that combines with the same mass of other element}} = \frac{\text{Atomic mass of A}}{\text{Atomic mass of B}}$$

- The valencies of the elements forming isomorphous compounds are the same.

(iv) Volatile chloride method

Required condition – chloride of element should be vapour.

Required data - (i) Vapour density of chloride. (ii) Equivalent weight of element.

Let the valency of the element be x. The formula of its chloride will be MCl_x.

Molecular weight = Atomic weight of M + 35.5 x

∵ Atomic weight = Equivalent weight × valency or A = E × x

∴ Molecular weight = E x + 35.5 x or 2 ×V.D. = x(E + 35.5) or $x = \dfrac{2 \times V.D.}{E+35.5}$

(v) Specific heat method : If $\dfrac{C_P}{C_V} = \gamma$ is given, then

Case I. If $\gamma = 5/3 = 1.66$ Atomicity will be one

Case II. If $\gamma = 7/5 = 1.4$ Atomicity will be two

Case III. If $\gamma = 4/3 = 1.33$ Atomicity will be three

Atomic weight = $\dfrac{\text{Molecular weight}}{\text{Atomicity}}$

(b) Method for Determination of Molecular Weight :

(i) Molecular weight = 2 × V.D.

(ii) Victor Mayer's method is used to determine molecular weight of volatile compound.

POINTS TO REVISE

- Equivalent weight of a species changes with reaction in which it gets involved.
- Amount of substance which loses or gains 1 mole electrons or 96500 coulomb electricity will always be its equivalent weight.

LAWS OF CHEMICAL COMBINATION

(a) Law of Mass Conservation (Law of Indestructibility of Matter)

"It was given by **Lavoisier** and tested by **Landolt"**

According to this law, the mass can neither be created nor be destroyed in a balanced chemical reaction or physical reaction. But one form is changed into another form is called as law of mass conservation.

If the reactants are completely converted into products, then the sum of the mass of reactants is equal to the sum of the mass of products.

$$\boxed{\textbf{Total mass of reactants = Total mass of products.}}$$

If reactants are not completely consumed then the relationship will be :

$$\boxed{\textbf{Total mass of reactants = Total mass of products + Mass of unreacted reactants}}$$

$$H_2 + Cl_2 \longrightarrow 2HCl$$

Ex. Mass in (g) $\underbrace{2 \quad 71}_{2\,+\,71\,=\,73\text{ g}} \qquad \underbrace{2(1 + 35.5)}_{73\text{ g}}$

(b) Law of Definite Proportion / Law of Constant Composition

"It was given by **Proust**."

According to this law, a compound can be obtained from different sources. But the ratio of each component (by weight) remain same.i.e. it does not depend on the method of its preparation or the source from which it has been obtained.

For example :- molecule of ammonia always has the formula NH_3. That is one molecule of ammonia always contains, one atom of nitrogen and three atoms of hydrogen or 17 g of NH_3 always contains 14 g of nitrogen and 3 g of hydrogen.

Ex. Water can be obtained from different sources but the ratio of weight of H and O remains same.

$\boxed{H_2}\ \boxed{O}$ $\longleftarrow$ $2H_2O \longleftarrow 2H_2 + O_2$; Sea water; Tap water; River water; Ganga Jal; Rain water

2 : 16
or 1 : 8

(c) Law of Multiple Proportion

"It was given by **John Dalton**"

According to law of Multiple proportion if two elements combine to form more than one compound then the different mass of one element which combine with a fixed mass of other element bear a simple ratio to one another.

The following examples illustrate this law.

(i) **Nitrogen and oxygen combine to form five oxides, which are :** Nitrous oxide (N_2O), nitric oxide (NO), nitrogen trioxide (N_2O_3), nitrogen tetraoxide (N_2O_4) and nitrogen pentaoxide (N_2O_5).

Weight of oxygen which combine with the fixed weight of nitrogen in these oxides are calculated as under:

Oxide Ratio of weight of nitrogen and oxygen in each compound

N_2O 28 : 16	NO 14 : 16	N_2O_3 28 : 48
N_2O_4 28 : 64	N_2O_5 28 : 80	

Number of parts by weight of oxygen which combine with 14 parts by weight of nitrogen from the above are 8,16,24,32 and 40 respectively. Their ratio is 1 : 2 : 3 : 4 : 5, which is a simple ratio. Hence, the law is illustrated.

(ii) Sulphur combines with oxygen to from two oxides SO_2 and SO_3, the weights of oxygen which combine with a fixed weight of sulphur, i.e. 32 parts by weight of sulphur in two oxides are in the ratio of 32 : 48 or 2 : 3 which is a simple ratio. Hence the law of multiple proportions is illustrated.

(d) Law of Gaseous Volume

"It was given by **Gay Lussac**"

According to this law, in the gaseous reaction, the reactants are always combined in a simple ratio by volume and form products, which is **simple ratio by volume** at same temperature and pressure.

Ex. One volume of hydrogen combines with one volume of chlorine to produce 2 volumes of hydrogen chloride.

Simple ratio = 1 : 1 : 2.

$$\boxed{1 \qquad\qquad 1 \qquad\qquad 2} \longleftarrow \text{Stoichiometry}$$

$$\begin{array}{ccccc} H_{2(g)} & + & Cl_{2(g)} & \longrightarrow & 2HCl_{(g)} \\ \text{1 Volume} & & \text{1 Volume} & & \text{2 Volume} \end{array}$$

Ex. One volume of nitrogen combines with 3 volumes of hydrogen to from 2 volumes of ammonia.

Simple ratio 1 : 3 : 2

$$\boxed{1 \qquad\qquad 3 \qquad\qquad 2}$$

$$\begin{array}{ccccc} N_{2(g)} & + & 3H_{2(g)} & \longrightarrow & 2NH_{3(g)} \\ \text{1 Volume} & & \text{3 Volume} & & \text{2 Volume} \end{array}$$

Special Note : This law is used only for gaseous reaction. It relate volume to mole or molecules. But not relate with mass.

EUDIOMETRY :

Eudiometry or gas analysis involves the calculations based on gaseous reactions or the reactions in which at least two components are gaseous, in which the amounts of gases are represented by their volumes, measured at the same pressure and temperature. Some basic assumptions related with calculations are:

1. Gay-Lussac's law of volume combination holds good. According to this law, the volumes of gaseous reactants reacted and the volumes of gaseous products formed, all measured at the same temperature and pressure, bear a simple ratio.

$N_2 (g) + 3H_2 (g) \longrightarrow 2NH_3 (g)$

1 vol. 3 vol. 2 vol.

Problem may be solved directly is terms of volume, in place of mole. The stoichiometric coefficients of a balanced chemical reactions gives the ratio of volumes in which gaseous substances are reacting and products are formed, at same temperature and pressure.

2. The volumes of solids or liquids is considered to be negligible in comparison to the volume of gas. It is due to the fact that the volume occupied by any substance in gaseous state is even more than thousand times the volume occupied by the same substance in solid or liquid states.

$2H_2 (g) + O_2 (g) \longrightarrow 2H_2O (l)$

2 mole 1 mole 2 mole

2 vol. 1 vol. 0 vol.

3. Air is considered as a mixture of oxygen and nitrogen gases only. It is due to the fact that about 99% volume of air is composed of oxygen and nitrogen gases only.

4. Nitrogen gas is considered as a non- reactive gas. It is due to the fact that nitrogen gas reacts only at very high temperature due to its very high thermal stability. Eudiometry is performed in an eudiometer tube and the tube can not withstand very high temperature. This is why, nitrogen gas can not participate in the reactions occurring in the eudiometer tube.

5. The total volume of non-reacting gaseous mixture is equal to sum of partial volumes of the component gases **(Amagat's law)**.

$V = V_1 + V_2 + \ldots\ldots\ldots\ldots\ldots$

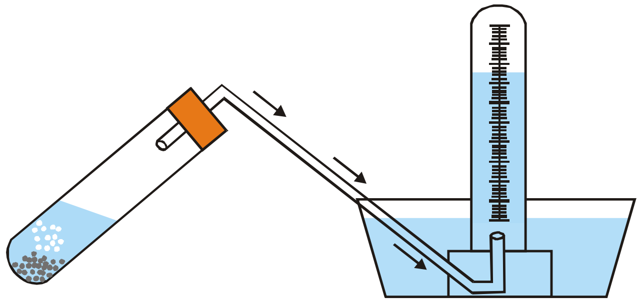
Partial volume of gas in a non-reacting gasesous mixture is its volume when the entire pressure of the mixture is supposed to be exerted only by that gas.

6. The volume of gases produced is often given by certain solvent which absorb contain gases.

Solvent	**Gases absorb**
KOH	CO_2, SO_2, Cl_2
Ammonical Cu_2Cl_2	CO
Turpentine oil	O_3
Alkaline pyrogallol	O_2
water	NH_3, HCl
$CuSO_4/CaCl_2$	H_2O

EUDIOMETER

An eudiometer is a laboratory device that measures the change in volume of a gas mixture following a physical or chemical change.



Scheme of eudiometer

To use a eudiometer, it is filled with water, inverted so that its open end is facing the ground (while holding the open end so that no water escapes), and then submersed in a basin of water. A chemical reaction is taking place through which gas is created. One reactant is typically at the bottom of the eudiometer (which flows downward when the eudiometer is inverted) and the other reactant is suspended on the rim of the eudiometer, typically by means of a platinum or copper wire (due to their low reactivity). When the gas created by the chemical reaction is released, it should rise into the eudiometer so that the experimenter may accurately read the volume of the gas produced at any given time. Normally a person would read the volume when the reaction is completed

Ex. 10 ml of CO is mixed with 25 ml air (20% O_2 by volume) in a container at 1 atm. Find final volume (in ml) of container at 1 atm after complete combustion. (Assume that temperature remain constant).

Ans. (30)

$$\underset{}{\overset{10\,ml}{CO}} + \overset{5\,ml}{\frac{1}{2}O_2} \longrightarrow \underset{10\,ml}{CO_2}$$

$V_f = V_{CO_2}$ + Volume of remaining air = 10 + 20 = 30 ml

Ex. A 3 L gas mixture of propane (C_3H_8) and butane (C_4H_{10}) on complete combustion at 25°C produced 10 L CO_2. Assuming constant P and T conditions what was volume of butane present in initial mixture ?

Ans. (1)

$$C_3H_8(g) + 5O_2 \longrightarrow 3CO_2(g) + 4H_2O(l)$$

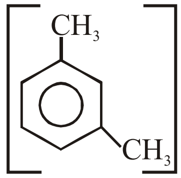
x L 3x L

$$C_4H_{10}(g) + \frac{13}{2}O_2(g) \longrightarrow 4CO_2(g) + 5H_2O(l)$$

(3-x) L 4(3 – x) L

from question $3x + 4(3 - x) = 10 \Rightarrow x = 2$

$\therefore$ Volume of butane , $C_4H_{10} = (3 - x) = 1$ L

Ex. **100 ml gaseous meta Xylene**  **undergoes combustion with excess of oxygen at room temperature and pressure. Volume contraction / expansion (in ml) during reaction is :**

Ans. (350)

$$C_4H_{10}(g) + \frac{21}{2}O_2(g) \longrightarrow 8CO_2(g) + 5H_2O(l)$$

$$100\text{ ml} \qquad \frac{21}{2}\times 100 = 1050\text{ml} \qquad 800\text{ ml} \qquad 0$$

$\therefore$ Conctraction in volume = (100 + 1050) – 800 = 350 ml

Ex. **30 ml gaseous mixture of methane and ethylene in volume ratio X : Y requires 350 ml air containing 20% of O_2 by volume for complete combustion. If ratio of methane and ethylene changed to Y : X. What will be volume of air (in ml) required for complete reaction under similar condition of temperature and pressure.**

Ans. 400

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O\ (l)$$

$$V_1\text{ml} \qquad 2V_1\text{ml} \qquad V_1\text{ml} \qquad 0$$

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O\ (l)$$

$$V_2\text{ml} \qquad 3V_2\text{ml} \qquad 2V_2\text{ml} \qquad 0$$

For given data : $V_1 + V_2 = 30$

and $2V_1 + 3V_2 = 350 \times \frac{20}{100} = 70$

$\therefore V_1 = 20 , V_2 = 10$

For required data : $V_1 = 10$ and $V_2 = 20$

$\therefore$ Volume of O_2 required = $2V_1 + 3V_2 = 80$ ml and volume of air required = $80 = \frac{100}{20} = 400\text{ml}$

Ex. **An alkene upon combustion produces CO_2(g) and H_2O(g). In this combustion process if there is no volume change occurs then the no. of C atoms per molecule of alkene will be :**

Ans.(2)

$$C_nH_{2n}(g) + \frac{3n}{2}\ O_2(g) \rightarrow nCO_2(g) + nH_2O(g)$$

if there no volume changes i.e. $\Delta_{ng} = 0$

$$(n + n) - \left(1+\frac{3n}{2}\right) = 0 \Rightarrow n = 2$$

Ex. A gaseous hydrocarbon (C_xH_y) requires 6 times of its own volume of O_2 for complete oxidation and produces 4 times of its volume of CO_2. Find out the volume of x + y.

Ans. (012)

$$C_xH_y + \left(x + \frac{y}{4}\right) O_2 \longrightarrow xCO_2 + \frac{y}{2}H_2O(l)$$

Vol a $\quad a\left(x+\frac{y}{4}\right) \quad$ ax

Given that : $a(x + y/4) = 6a$

vol of CO_2 = 4 vol of C_xH_y

$ax = 4(a)$

$x = 4$ ...(2)

from (1) $x + \frac{y}{4} = 6$

$\therefore x + y = 4 + 8 = 12$

Ex. On heating 60 ml mixture containing equal volume of chlorine gas and it's gaseous oxide, volume becomes 75 ml due complete decomposition of oxide. On treatment with KOH volume becomes 15 ml. What is the formula of oxide of chlorine ?

Ans. Let oxide of Cl is Cl_xO_y

So in 60 mL $\Rightarrow$ 30 mL Cl_xO_y and 30 mL Cl_2.

Now,

$$Cl_xO_y \longrightarrow \frac{x}{2}Cl_2 + \frac{y}{2}O_2$$

30mL $\quad \frac{30.x}{2}$mL $\quad \frac{30.y}{2}$mL

Given :

$$75 = 30 + \frac{30x}{2} + \frac{30y}{2} \Rightarrow x + y = 3 \text{ ....(i)}$$

KOH absorbs Cl_2 and volume becomes 15 mL so,

$$(75 - 15) = V_{Cl_2} = 30 + \frac{30x}{2} \Rightarrow x = 2 \text{ and } y = 1$$

So the oxide : Cl_2O

Ex. 5 L of A (g) & 3 L of B(g) measured at same T & P are mixed together which react as follows

$2A(g) + B(g) \rightarrow C(g)$

What will be the total volume (in litre) after the completion of the reaction at same T & P.

Ans. (3)

Sol. $2A(g) + B(g) \longrightarrow C(g)$

5L $\quad$ 3L

L.R. is A

So, volume of C produced = $\frac{1}{2}\times 5$ = 2.5 L

and, volume of B reacted = $\frac{1}{2}\times 5$ = 2.5 L

So, volume fo B remained = 3 − 2.5 = 0.5 L

Hence, $V_{total} = V_C + V_B = 2.5 + 0.5 = 3$ L

Ex. 5 ml of gas containing only carbon and hydrogen was mixed with an excess of oxygen (30 ml) and the mixture exploded by means of an electric spark. After the explosion , the volume of the mixed gases remaining was 25 ml. On adding a concentrated solution of potassium hydroxide, the volume further diminished to 15 mole the residual gas being pure oxygen. All volumes have been reduced to NTP. Calculate the molecular formula of the hydrocarbon. **[JEE-1979]**

Sol. Volume of oxygen taken = 30 ml

Volume of unused oxygen = 15 ml

Volume of O_2 used = Volume of O_2 added - volume of O_2 left.

= 30 – 15 = 15 ml.

or volume of CO_2 produced = 25 – 15 = 10 ml

General equation of the combustion of a hydrocarbon is as following lows.

$$C_xH_y + \left(x+\frac{y}{4}\right) O_2 \rightarrow xCO_2 + \left(\frac{y}{2}\right) H_2O$$

(Hydrocarbon) $5\left(x+\frac{y}{2}\right)$ml 5x

5 ml

$\therefore$ Volume of CO_2 produced

= 5x, since volume of CO_2 = 10 ml

$\therefore$ 5x = 10 $\Rightarrow$ x = 2, volume of O_2 used = 15

$\therefore\ 5\left(x+\frac{y}{2}\right) = 15 \Rightarrow x + \frac{y}{4} = 3$

$\Rightarrow 2 + \frac{y}{4} = 3$ $(\because = 2)$

$\Rightarrow 8 + y = 12$ $\therefore\ y = 4$

Ex. A 20.0 ml mixture of CO , CH_4 and He gases is exploded by an electric discharge at room temperature wth excess of oxygen. The volume contraction is found to be 13.0 cm^3. A further contraction of 14.0 cm^3 occurs when the residual gas is treated with KOH solution. Find out the composition of the gaseous mixture in terms of volume percentage. **[JEE-1995]**

Sol. $CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g)$

$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(l)$

'x' is the volume of CO and y is the volume of CH_4

Thus, $\frac{1}{2}x + y = 13$ (1)

x + y = 14 (2)

x = 10cc and y = 4cc

Thus, % CH_4 = 20, % CO = 50, % He = 30