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PRE-MEDICAL

PHYSICAL CHEMISTRY

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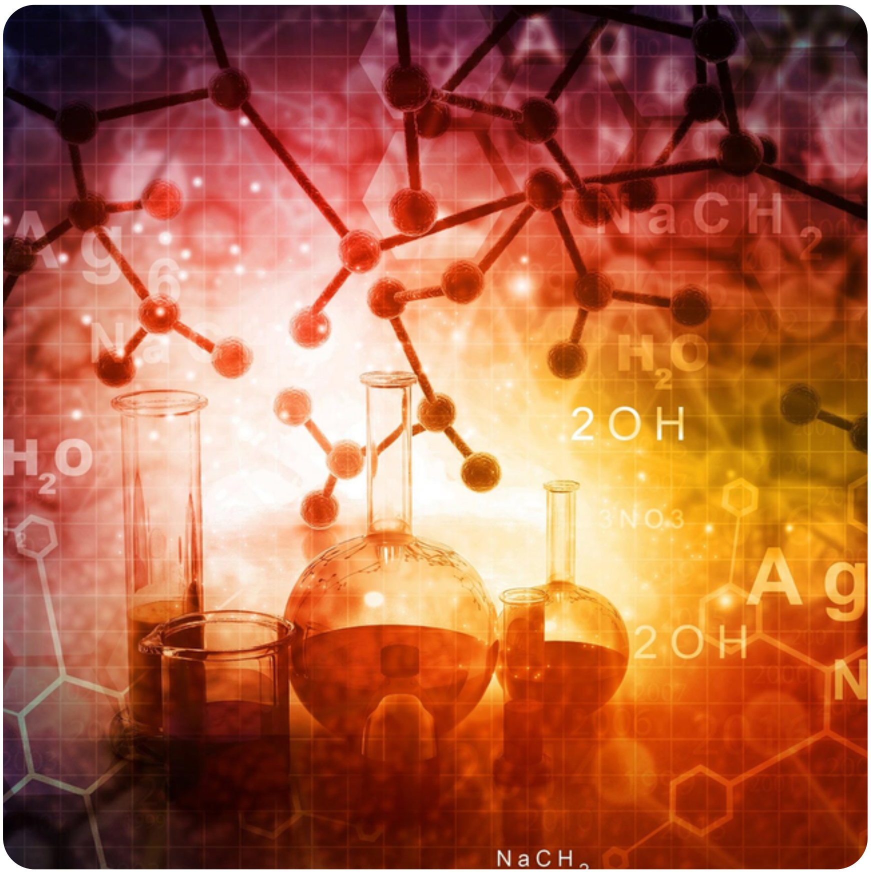

STUDY MATERIAL

Some Basic concept of chemistry

ENGLISH MEDIUM

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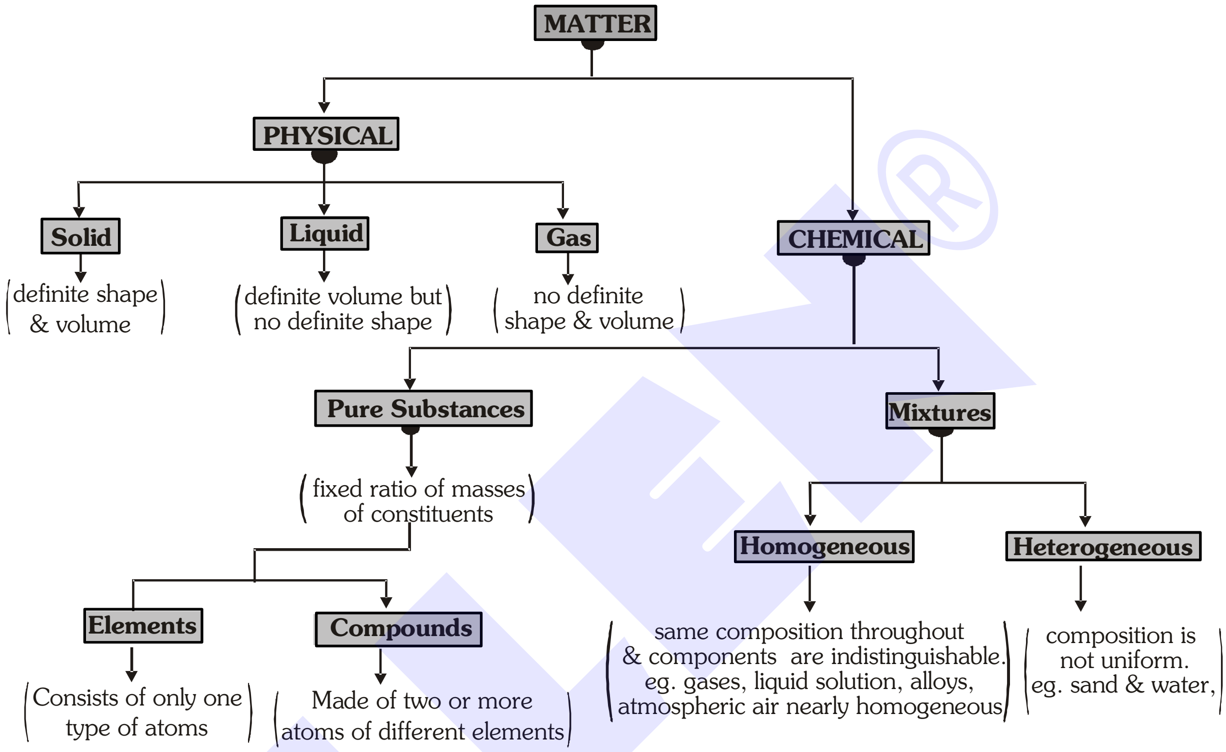
SOME BASIC CONCEPTS OF CHEMISTRY

1.0 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 placed.
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.

Illustrations

Illustration 1. Which is an example of matter according to physical state at room temperature and pressure.

(1) solid (2) liquid (3) gas (4) all of these

Solution **Ans. (4)** According to the physical state at room temperature and pressure, the matter is present in 3 state solid, liquid & gas

Illustration 2. Which of the following are the types of the compound.

(1) Organic compound (2) Inorganic compound

(3) Both (1) and (2) (4) None of these

Solution **Ans. (3)** Compound is divided into 2 types. Inorganic compound & Organic compound

Illustration 3. Which of the following is an example of a homogeneous mixture.

(1) Water + Alcohol (2) Water + Sand (3) Water + Oil (4) None of these

Solution **Ans. (1)** Water and alcohol are completely mixed and form uniform solution.

Illustration 4. Which of the following is a solution.

(1) Heterogeneous mixture (2) Homogeneous mixture

(3) Both (1) and (2) (4) None of these

Solution **Ans. (2)** Homogeneous mixture is a solution.

Illustration 5. Which of the following is a compound

(1) graphite (2) producer gas (3) cement (4) marble

Solution **Ans. (4)** Marble = $CaCO_3$ = compound.

Illustration 6. Which of the following statements is/are true :

(1) An element of a substance contains only one kind of atoms.

(2) A compound can be decomposed into its components.

(3) All homogeneous mixtures are solutions.

(4) All of these

Solution **Ans. (4)**

Illustration 7. A pure substance can only be :-

(1) A compound (2) An element

(3) An element or a compound (4) A heterogenous mixture

Solution **Ans. (3)**

Illustration 8. Which one of the following is not a mixture :

(1) Tap water (2) Distilled water (3) Salt in water (4) Oil in water

Solution **Ans. (2)**

1.1 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_v	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 : °C + 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

GOLDEN KEY POINTS

- 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 $(centimeter)^3$ and not centi $(meter)^3$.

Illustrations

Illustration 9. Which one of the following forms part of seven basic SI units :
(1) Joule (2) Candela (3) Newton (4) Pascal

Solution **Ans. (2)**

Illustration 10 Convert 2 litre atmosphere into erg.

Solution 2 litre atmosphere = 2×101.3 joule = $2\times101.3\times10^7$ erg. = 202.6×10^7 erg.
{1 litre atmosphere = 101.3J}

Illustration 11 Convert 2 atm into cm of Hg.

Solution 2 atm = 2×76 cm of Hg = 152 cm of Hg
{1 atmosphere = 76 cm of Hg}

Illustration 12 Convert 20 dm^3 into mL.

Solution 20 dm^3 = 20 L = 20×1000 mL = 2×10^4 mL
1 dm^3 = 1 L = 1000 mL

Illustration 13 Convert 59 F into ℃.

Solution $℃ = \frac{5}{9}(F-32) = \frac{5}{9}(59-32) = \frac{5}{9}\times 27 = 15℃$

1.2 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 :

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

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

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

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

(iii) $\boxed{\text{Number of moles (n)} = \dfrac{N}{N_A}}$ (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

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

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.

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

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

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

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 $= 18 \times 1.67 \times 10^{-24} g = 2.99 \times 10^{-23}\ g$

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

Illustrations

Illustration 14. Find out the volume and mole in 56 g nitrogen at STP

Solution Molar mass of N_2 is 28 g

(a) Calculation of volume : $\because$ 28 g of N_2 occupies = 22.4 L at STP

$\therefore$ 56 g of N_2 occupies $= \frac{22.4}{28} \times 56$ L = 44.8 L at STP

(b) Calculation of mole : $\because$ 28 g of N_2 = 1 mol of N_2

$\therefore$ 56 g of N_2 $= \frac{1}{28} \times 56 = 2$ mol of N_2

Illustration 15. Calculate the volume and mass of 0.2 mol of O_3 at STP.

Solution (a) Calculation of volume : $\because$ volume of 1 mole of O_3 at STP = 22.4 L

$\therefore$ volume of 0.2 mole of O_3 at STP = 22.4 × 0.2

= 4.48 L

(b) Calculation of mass : $\because$ mass of 1 mol of O_3 = 48 g

$\therefore$ mass of 0.2 mol of O_3 = 48 × 0.2 g = 9.6 g

Illustration 16. Find out the moles & mass in 1.12 L O_3 at STP.

Solution (a) Calculation of mole: $\because$ at STP 22.4 L of O_3 contain = 1 mol of O_3

$\therefore$ at STP 1.12 L of O_3 contain $= \frac{1}{22.4} \times 1.12$

= 0.05 mol of O_3

(b) Calculation of mass : Molecular weight of O_3 = 48 g

$\because$ weight of 22.4 L of O_3 at STP is = 48 g

$\therefore$ weight of 1.12 L of O_3 at STP is $= \frac{48}{22.4} \times 1.12 = 2.4$ g

Illustration 17. Find out the mass of 10^{21} atoms of Cu.

Solution Number of moles of Cu $= \frac{N}{N_A} = \frac{10^{21}}{6.023 \times 10^{23}} = \frac{\text{weight}}{\text{Atomic weight}} = \frac{\text{weight}}{63.5}$

weight of Cu $= \frac{10^{21}}{6.023 \times 10^{23}} \times 63.5 = 0.106$ g

Illustration 18. Calculate the number of molecules and number of atoms present in 1 g of nitrogen ?

Solution Number of moles (n) $= \frac{\text{weight}}{M_w} = \frac{1}{28} \Rightarrow$ Number of molecules (N) $= \frac{N_A}{28}$

$\because$ 1 molecule of N_2 gas contain = 2 atoms

$\therefore$ $\frac{N_A}{28}$ molecules of N_2 gas contain $= 2 \times \frac{N_A}{28} = \frac{N_A}{14}$ atoms

Illustration 19. Calculate the number of moles in 11.2 L at STP of oxygen.

Solution Number of moles of O_2 (n) $= \frac{V}{22.4} = \frac{11.2}{22.4} = 0.5$ mol

Illustration 20. For $\frac{1}{2}$ **g molecule** of oxygen. Find (i) mass, (ii) number of molecules, (iii) volume at STP. (iv) number of oxygen atoms.

Solution

(i) $n = \frac{1}{2}\text{ mol} = \frac{\text{weight}}{M_w} = \frac{\text{weight}}{32} \Rightarrow$ weight of oxygen = 16 g

(ii) $n = \frac{1}{2}\text{ mol} = \frac{N}{N_A} \Rightarrow$ Number of molecules of oxygen $(N) = \frac{N_A}{2}$

(iii) $n = \frac{1}{2}\text{ mol} = \frac{V}{22.4} \Rightarrow V = 11.2\text{ L}$

(iv) 1 molecule of O_2 contain = 2 oxygen atoms.

$\frac{N_A}{2}$ molecules of O_2 contain $= \frac{N_A}{2} \times 2 = N_A$ oxygen atoms.

BEGINNER'S BOX-1

1. The modern atomic weight scale is based on.

(1) C^{12} (2) O^{16} (3) H^{1} (4) C^{13}

2. Gram atomic weight of oxygen is

(1) 16 amu (2) 16 g (3) 32 amu (4) 32 g

3. Molecular weight of SO_2 is :

(1) 64 g (2) 64 amu (3) 32 g (4) 32 amu

4. 1 amu is equal to :-

(1) $\frac{1}{12}$ of C^{12} (2) $\frac{1}{14}$ of O^{16} (3) 1 g of H_2 (4) 1.66×10^{-24} kg

5. The actual molecular mass of chlorine is :

(1) 58.93×10^{-24} g (2) 117.86×10^{-24} g (3) 58.93×10^{-24} kg (4) 117.86×10^{-24} kg

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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 :

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

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

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

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

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

$V = 1\ L = 1000\ mL$

then $m_{H_2} = 0.089g$

$\boxed{M_W(g) = \text{Mass of 22.4 litre gas}}$ **or** $\boxed{M_w(g) = \text{22.4 litre (at STP)}}$

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 = $\dfrac{1\text{litre}}{0.089}$ at STP

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

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

$\boxed{\text{1 mol = 22.4 L (at STP)}}$

Illustrations

Illustration 21. Calculate the number of atoms of chlorine in 2.08 g of $BaCl_2$. (Atomic weight of Ba = 137, Cl = 35.5)

Solution Number of moles of $BaCl_2$ (n) = $\dfrac{\text{weight}}{M_w} = \dfrac{2.08}{208} = 0.01 \text{ mol} = \dfrac{N}{N_A}$

Number of molecules of $BaCl_2$ (N) = 0.01 N_A

1 molecule of $BaCl_2$ contain = 2 chlorine atoms.

0.01 N_A molecules $BaCl_2$ contain = $2 \times 0.01\ N_A$ Chlorine atoms. = $2 \times 10^{-2}\ N_A$ Chlorine atoms.

Illustration 22. Calculate the number of molecules and number of atoms present in 1.2 g of ozone.

Solution Number of moles of O_3 (n) = $\dfrac{\text{weight}}{M_w} = \dfrac{1.2}{48} = \dfrac{1}{40}$ mol

$\Rightarrow$ number of molecules of O_3 (N) = $\dfrac{N_A}{40}$

$\because$ 1 molecule of O_3 contain = 3 atoms, $\therefore$ $\dfrac{N_A}{40}$ molecules O_3 contain = $\dfrac{3N_A}{40}$ atoms.

Illustration 23. Calculate the number of atoms present in one drop of water having mass 1.8 g.

Solution Number of moles of H_2O (n) = $\frac{\text{weight}}{M_w} = \frac{1.8}{18} = 0.1$ mol

Number of molecules of H_2O (N) = 0.1 N_A

$\because$ 1 molecule of H_2O contain = 3 atoms

$\therefore$ 0.1 N_A molecules of H_2O contain = 3 × (0.1 N_A) = 0.3 N_A atoms.

Illustration 24. Calculate the number of atoms present in one litre of water (density of water is 1 g/mL).

Solution 1 litre = 1000 mL = 1000 g

Moles of H_2O (n) = $\frac{\text{weight}}{M_w} = \frac{1000}{18} = 55.5 \text{ mol} = \frac{N}{N_A}$

$\Rightarrow$ number of molecules of H_2O (N) = 55.5 N_A

$\because$ 1 molecule of H_2O contain = 3 atoms

$\therefore$ 55.5 N_A molecules H_2O contain = 3 × (55.5 N_A) atoms = 166.5 N_A atoms

Illustration 25. At NTP the density of a gas is 0.00445 g/mL then find out its V.D. and molecular mass.

Solution V.D. = $\frac{\text{Density of gas}}{\text{Density of } H_2} = \frac{0.004450}{0.000089} = 50$

Molecular mass = 2 × V.D.= 2 ×50 = 100

Illustration 26. Weight of 1 L gas is 2 g then find out its V.D. and molecular mass

Solution Density of gas = $\frac{\text{Mass}}{\text{Volume}} = \frac{2}{1000} = 0.002$ g/mL

V.D. = $\frac{\text{Density of gas}}{\text{Density of } H_2} = \frac{0.002000}{0.000089} = 22.4$

Molecular mass = 2 × V.D. = 44.8

GOLDEN KEY POINTS

- 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 ratio's 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.

BEGINNER'S BOX-2

1. Calculate the number of atoms in 11.2 L of SO_2 gas at STP :

(1) $\frac{N_A}{2}$ (2) $\frac{3N_A}{2}$ (3) $3N_A$ (4) N_A

2. Which of the following has maximum mass :

(1) 0.1 gram atom of carbon (2) 0.1 mol of ammonia

(3) 6.02×10^{22} molecules of hydrogen (4) 1120 cc of carbon dioxide at STP

3. The total number of electrons present in 18 mL of water :-

(1) 6.02×10^{22} (2) 6.02×10^{23} (3) 6.02×10^{24} (4) 6.02×10^{25}

4. The volume of 1.0 g of hydrogen at NTP is :

(1) 2.24 L (2) 22.4 L (3) 1.12 L (4) 11.2 L

5. 11 grams of a gas occupy 5.6 litres of volume at STP. The gas is :-

(1) NO (2) N_2O_4 (3) CO (4) CO_2

6. At NTP, 5.6 L of a gas weight 8 grams. The vapour density of gas is :-

(1) 32 (2) 40 (3) 16 (4) 8

7. The vapour densities of two gases are in the ratio of 1 : 3. Their molecular masses are in the ratio of :-

(1) 1 : 3 (2) 1 : 2 (3) 2 : 3 (4) 3 : 1

1.3 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 27. Find out percentage composition of each element present in glucose ?

Solution

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

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

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

Illustration 28. 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	$\frac{75.8}{24}=3.1$	$\frac{3.1}{1.5}=2.06$	2
y	24.2%	16	$\frac{24.2}{16}=1.5$	$\frac{1.5}{1.5}=1$	1

Empirical formula = x_2y

Illustration 29. 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	$\frac{13}{1}=13$	$\frac{13}{2.2}=5.9$	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 ×2 + 16 + 6 = 46

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

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

BEGINNER'S BOX-3

1. A hydrocarbon contain 80% C. The vapour density of compound is 30. Empirical formula of compound is :-

(1) CH_3 (2) C_2H_6 (3) C_4H_{12} (4) C_4H_8

2. Two elements X (Atomic weight = 75) and Y (Atomic weight = 16) combine to give a compound having 75.8% of X. The empirical formula of compound is :

(1) XY (2) X_2Y (3) X_2Y_2 (4) X_2Y_3

3. In a compound element A (Atomic weight = 12.5) is 25% and element B (Atomic weight = 37.5) is 75% by weight. The Empirical formula of the compound is :

(1) AB (2) A_2B (3) A_2B_2 (4) A_2B_3

1.4 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.1

$CaCO_{3(s)}$	+	$2HCl_{(aq)}$	$\longrightarrow$	$CaCl_{2(aq)}$	+	$H_2O_{(l)}$	+	$CO_{2(s)}$
1 Mol		2 Mol		1 Mol		1 Mol		1 Mol
$40 + 12 + 3 \times 16$		$2(1 + 35.5)$		$40 + 2 \times 35.5$		$2 \times 1 + 16$		$12 + 2 \times 16$
= 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.2

1		3		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 $\quad C_2H_6 + O_2 \longrightarrow 2CO_2 + H_2O$

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

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

Illustrations

TYPE-I (INVOLVING MASS-MASS RELATIONSHIP)

Illustration 30. How much iron can be theoretically obtained in the reduction of 1 kg of Fe_2O_3

Solution

1	3	2	3

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

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

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

Solution

1	1	1	1

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

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

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

Solution

2	7	4	6

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

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

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

volume of O_2 at STP = $0.35 \times 22.4 = 7.84$ L

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

volume of CO_2 at STP = $0.2 \times 22.4 = 4.48$ L

Illustration 33. 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

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

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

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 34. 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 = $5 \times 1.12 = 5.6$ L

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

Illustration 35. 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 = $5 \times 5 = 25$ mol $= \frac{V}{22.4}$

volume of O_2 gas at STP = $25 \times 22.4 = 560$ 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	9 mol			
	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. $A + 2B \rightarrow P$

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

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

Illustrations

Illustration 36. $A + 5B \longrightarrow C + 3D$ 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 $\quad$ 10 mol

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

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

Illustration 37. $H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(g)}$; 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

1	$1/2$	1		
$H_{2(g)}$ +	$\frac{1}{2}O_{2(g)}$ ⟶	$H_2O_{(g)}$		
4 g	32 g		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 = $\frac{V}{22.4}$ $\quad$ 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 38. 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

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

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

BEGINNER'S BOX-4

1. 1.5 mol of O_2 combine with Mg to form oxide MgO. The mass of Mg (At. mass 24) that has combined is :

(1) 72 g (2) 36 g (3) 24 g (4) 94 g

2. What quantity of lime stone on heating will give 56 kg of CaO :-

(1) 1000 kg (2) 56 kg (3) 44 kg (4) 100 kg

3. For reaction A + 2 B $\rightarrow$ C. The amount of product formed by starting the reaction with 5 mol of A and 8 mol of B is :

(1) 5 mol (2) 8 mol (3) 16 mol (4) 4 mol

1.5 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 $= \frac{\text{Atomic weight}}{\text{Valency factor}}$

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

(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 $= \frac{\text{Molecular weight}}{\text{Basicity/Acidity}}$

(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 = $(2\times23 + 32 + 16\times4) = 142$

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 = $\dfrac{W_{(gram)}}{E}$

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

$$= n \times \text{valence factor}$$

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.

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

(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)

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

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

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

(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.

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

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

(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.

1.6 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} $°C^{-1}$) ≈ 6.4

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

(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

$\because$ Atomic weight = Equivalent weight $\times$valency or A = E $\times$ x

$\therefore$ Molecular weight = E x + 35.5 x or 2 $\times$V.D. = x(E + 35.5) or $x = \frac{2 \times V.D.}{E + 35.5}$

(v) Specific heat method : If $\frac{C_P}{C_V} = \gamma$ 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

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

(b) Method for Determination of Molecular Weight :

(i) Molecular weight = 2 $\times$ V.D.

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

Illustrations

Illustration 39. Specific heat of a metal is $0.031 \frac{\text{Cal}}{°C\ g}$, and its equivalent weight is 103.6. Calculate the exact atomic weight of the metal.

Solution According to Dulong and Petit's law - approximate atomic weight = $\frac{6.4}{0.031} = 206.45$

$$\text{Valency of metal} = \frac{\text{Approximate atomic weight}}{\text{Equivalent weight}} = \frac{206.45}{103.6} = 1.99 \simeq 2$$

So, the exact atomic weight of the element = Equivalent weight $\times$ valency

= 103.6 $\times$2 = 207.2

Illustration 40. A chloride of an element contains 49.5% chlorine. The specific heat of the element is 0.064 cal g^{-1} $°C^{-1}$. Calculate the equivalent mass, valency and atomic mass of the element.

Solution Mass of chlorine in the metal chloride = 49.5

Mass of metal = (100 – 49.5) = 50.5

$$\text{Equivalent weight of metal} = \frac{\text{weight of metal}}{\text{weight of chlorine}} \times 35.5 = \frac{50.5}{49.50} \times 35.5 = 36.21$$

Now according to Dulong and Petit's law,

$$\text{Approximate at. wt. of the metal} = \frac{6.4}{\text{specific heat}} = \frac{6.4}{0.064} = 100$$

$$\text{Valency} = \frac{\text{Approximate atomic weight}}{\text{Equivalent weight}} = \frac{100}{36.21} = 2.7 \simeq 3$$

Hence, exact atomic weight = 36.21 $\times$3 = 108.63

Illustration 41. The oxide of an element contains 67.67% of oxygen and the vapour density of its volatile chloride is 79. Calculate the atomic weight of the element.

Solution Calculation of equivalent weight : weight of oxygen = 67.67 g

weight of element = 100 – 67.67 = 32.33 g

$\because$ 67.67 g of oxygen combines with 32.33 g of element

$\therefore$ 8 g of oxygen combines with $= \dfrac{32.33 \times 8}{67.67} = 3.82$ g of element

$\therefore$ Equivalent weight of the element = 3.82

Suppose M represents one atom of the element and x is its valency. The molecular formula of the volatile chloride would be MCl_x.

Formula weight of chloride = 3.82 × x + 35.5 x = 39.32 x

But molecular weight of Chloride = 2 × V.D. $\Rightarrow 39.32\,x = 2 \times 79 \Rightarrow x = \dfrac{2 \times 79}{39.32} = 4$

Now atomic weight = Equivalent weight × valency of element = 3.82 × 4 = 15.28

Illustration 42. Vapour density of a gas is 16. If the ratio of specific heat at constant pressure and specific heat at constant volume is 1.4. Then find out its atomic weight.

Solution Given : $\dfrac{C_P}{C_V} = 1.4 = \gamma$ and vapour density = 16

We know that Molecular weight = 2 × vapour density

$\therefore$ Molecular weight = 2 × 16 = 32

Here γ = 1.4 so atomicity will be 2.

$$\text{Atomic weight} = \frac{\text{Molecular weight}}{\text{Atomicity}} = \frac{32}{2} = 16$$

GOLDEN KEY POINTS

- 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.

BEGINNER'S BOX-5

1. Molecular weight of dibasic acid is W. Its equivalent weight will be :

(1) $\dfrac{W}{2}$ (2) $\dfrac{W}{3}$

(3) W (4) 3W

2. 0.126 g of an acid requires 20 ml of 0.1 N NaOH for complete neutralization. Eq. wt. of the acid is:

(1) 45 (2) 53

(3) 40 (4) 63

3. 1 mol O_2 will be equal to :

(1) 4 g equivalent oxygen (2) 2 g equivalent oxygen

(3) 32 g equivalent oxygen (4) 8 g equivalent oxygen

4. Volume of one gram equivalent of H_2 at NTP is :

(1) 5.6 L (2) 11.2 L

(3) 22.4 L (4) 44.8 L

1.7 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.

Total mass of reactants = Total mass of products.

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

Total mass of reactants = Total mass of products + Mass of unreacted reactants

Ex.

$$\begin{array}{lccc} & H_2 + Cl_2 & \longrightarrow & 2HCl \\ \text{Mass in (g)} & 2 + 71 & & 2(1+35.5) \\ & \downarrow & & \downarrow \\ & 2 + 71 = 73\text{ g} & & 73\text{ g} \end{array}$$

Illustrations

Illustration 43. What weight of $BaCl_2$ would react with 24.4 g of sodium sulphate to produce 46.6 g of barium sulphate and 23.4 g of sodium chloride ?

Solution Barium chloride and sodium sulphate react to produce barium sulphate and sodium chloride according to the equation :

$$\begin{array}{ccccc} BaCl_2 + Na_2SO_4 & \longrightarrow & BaSO_4 + 2NaCl \\ x\text{ g} \quad 24.4\text{ g} & & 46.6\text{ g} \quad 23.4\text{ g} \end{array}$$

Let the weight of $BaCl_2$ be x g. According to law of conservation of mass :

Total mass of reactants = Total mass of products

Total mass of reactants = (x + 24.4) g

Total mass of products = (46.6 + 23.4) g

Equating the two masses $\Rightarrow$ x + 24.4 = 46.6 + 23.4

x = 46.6 + 23.4 – 24.4 or x = 45.6 g

Hence, the weight of $BaCl_2$ is 45.6 g

Illustration 44. 10g of $CaCO_3$ on heating gives 4.4 g of CO_2 then determine weight of produced CaO in quintal.

Solution

$$\begin{array}{lccc} CaCO_3 \longrightarrow & CaO & + & CO_2 \\ 10\text{ g} & x\text{ g} & & 4.4\text{ g} \end{array}$$

According to law of conservation of mass

$$10 = 4.4 + x \qquad \begin{cases} 1\text{quintal} = 100\text{kg} \\ 1\text{kg} = 1000\text{g} \end{cases}$$

$$10 - 4.4 = x$$

$$x = 5.6\text{ g}$$

weight of CaO(x) $= 5.6 \times \frac{\text{kg}}{1000} = 5.6 \times 10^{-3}\text{kg} = 5.6 \times 10^{-3} \times \frac{1}{100}$ quintal $= 5.6 \times 10^{-5}$ quintal

(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

$\boxed{H_2}\ \boxed{O}$ ← $2H_2O \leftarrow 2H_2 + O_2$, Sea water, Tap water, River water, Ganga Jal, Rain water

2 : 16

or 1 : 8

Illustrations

Illustration 45. Weight of copper oxide obtained by treating 2.16 g of metallic copper with nitric acid and subsequent ignition was 2.70 g. In another experiment, 1.15 g of copper oxide on reduction yielded 0.92 g of copper. Show that the results illustrate the law of constant composition.

Solution

In I experiment

weight of Cu = 2.16 g

weight of CuO = 2.7 g

weight of Oxygen = 2.7 – 2.16 = 0.54 g

Cu : O

2.16 : 0.54

$\frac{2.16}{0.54} : \frac{0.54}{0.54}$

4 : 1

In II experiment

weight of CuO = 1.15 g

weight of Cu = 0.92 g

weight of Oxygen = 1.15 – 0.92 = 0.23 g

Cu : O

0.92 : 0.23

$\frac{0.92}{0.23} : \frac{0.23}{0.23}$

4 : 1

Thus the ratio of the masses of copper and oxygen in the two experiment are same. Hence the given data illustrate the law of constant proportion.

Illustration 46. In an experiment 2.4 g of FeO on reduction with hydrogen gives 1.68 g of Fe. In another experiment 2.9 g of FeO gives 2.03 g of Fe on reduction with hydrogen. Show that the above data illustrate the law of constant proportion.

Solution

In I experiment

Weight of FeO = 2.4 g

Weight of Fe = 1.68 g

Weight of Oxygen = 2.4 – 1.68 = 0.72 g

Fe : O

1.68 : 0.72

$\frac{1.68}{0.72} : \frac{0.72}{0.72}$

2.33 : 1

In II experiment

Weight of FeO = 2.9 g

Weight of Fe = 2.03 g

Weight of Oxygen = 2.9 – 2.03 = 0.87 g

Fe : O

2.03 : 0.87

$\frac{2.03}{0.87} : \frac{0.87}{0.87}$

2.33 : 1

Thus the ratio of the masses of iron and oxygen in the two experiment are same. Hence the given data illustrate the law of constant proportion.

(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.

Illustrations

Illustration 47. Hydrogen peroxide and water contain 5.93% and 11.2% of hydrogen respectively. Show that the data illustrate the law of multiple proportions.

Solution

Compound H_2O_2

H : O

5.93 : 94.07

$\frac{5.93}{5.93} : \frac{94.07}{5.93}$

1 : 15.86

Compound H_2O

H : O

11.2 : 88.8

$\frac{11.2}{11.2} : \frac{88.8}{11.2}$

1 : 7.92

Thus the ratio of weighs of oxygen which combine with the fixed weight (1.0 gram) of hydrogen in H_2O_2 and H_2O is 15.86 : 7.92 = 2 : 1 (Which is simple ratio). So the law of multiple proportion is illustrated.

Illustration 48. Carbon combines with hydrogen in P, Q and R. The % of hydrogen in P, Q and R are 25, 14.3, and 7.7 respectively. Which law of chemical combination is illustrated ?

Solution

P	Q	R
H : C	H : C	H : C
25 : 75	14.3 : 85.7	7.7 : 92.3
$1 : \frac{75}{25}$	$1 : \frac{85.7}{14.3}$	$1 : \frac{92.3}{7.7}$
1 : 3	1 : 6	1 : 12

Ratio of C in compounds P, Q and R is = 3 : 6 : 12 = 1 : 2 : 4

Which is a simple ratio so the data illustrate the law of multiple proportion.

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

Simple ratio = 1 : 1 : 2.

1		1		2	⟵ Stoichiometry
$H_{2(g)}$	+	$Cl_{2(g)}$	$\longrightarrow$	$2HCl_{(g)}$	
1 Volume		1 Volume		2 Volume	

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

Simple ratio 1 : 3 : 2

1		3		2
$N_{2(g)}$	+	$3H_{2(g)}$	$\longrightarrow$	$2NH_{3(g)}$
1 Volume		3 Volume		2 Volume

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

Illustrations

Illustration 49. For the gaseous reaction : $H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(g)}$. If 40 mL of hydrogen completely reacts with chlorine then find out the required volume of Chlorine & volume of produced $HCl_{(g)}$?

Solution According to Gay Lussac's Law :

1		1		2
$H_{2(g)}$	+	$Cl_{2(g)}$	$\longrightarrow$	$2HCl_{(g)}$

$\because$ 1 mL of $H_{2(g)}$ will react with 1 mL of $Cl_{2(g)}$ and 2 mL of $HCl_{(g)}$ will produce

$\therefore$ 40 mL of $H_{2(g)}$ will react with 40 mL of $Cl_{2(g)}$ and 80 mL of $HCl_{(g)}$ will produce

required volume of $Cl_{2(g)}$ = 40 mL

produced volume of $HCl_{(g)}$ = 80 mL

Illustration 50. For the gaseous reaction : $H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(g)}$. If initially 20 mL of $H_{2(g)}$ and 30 mL of $Cl_{2(g)}$ are present then find out the volume of $HCl_{(g)}$ and unreacted part of $Cl_{2(g)}$.

Solution According to Gay-Lussac's Law

1		1		2
$H_{2(g)}$	+	$Cl_{2(g)}$	$\longrightarrow$	$2HCl_{(g)}$

$\because$ 1 mL of $H_{2(g)}$ will react with 1 mL of $Cl_{2(g)}$ and 2 mL of $HCl_{(g)}$ will produce

$\therefore$ 20 mL of $H_{2(g)}$ will react with 20 mL of $Cl_{2(g)}$ and 40 mL of $HCl_{(g)}$ will produce

Given volume of $Cl_{2(g)}$ is 30 mL but its 20 mL reacts with $H_{2(g)}$. So 10 mL of $Cl_{2(g)}$ remains unreacted.

(e) Avogadro's law

"Equal volume of all gases contain equal number of molecules at same temperature and pressure."

Ex. | 1 | 1 | 2 | ⟵ Stoichiometry

$H_{2(g)}$	+ $Cl_{2(g)}$	$\longrightarrow$	$2HCl_{(g)}$
1 Volume	1 Volume		2 Volume
N molecules	N molecules		2N molecules
$\boxed{\frac{1}{2}\text{molecule}}$ (1 atom)	$\frac{1}{2}$ molecule (1 atom)		1 molecule

It is correct due to molecule is divisible.

ANSWER KEY

BEGINNER'S BOX-1	Que.	1	2	3	4	5		
	Ans.	**1**	**2**	**2**	**1**	**2**		

BEGINNER'S BOX-2	Que.	1	2	3	4	5	6	7
	Ans.	**2**	**4**	**3**	**4**	**4**	**3**	**1**

BEGINNER'S BOX-3	Que.	1	2	3
	Ans.	**1**	**4**	**1**

BEGINNER'S BOX-4	Que.	1	2	3
	Ans.	**1**	**4**	**4**

BEGINNER'S BOX-5	Que.	1	2	3	4	
	Ans.	**1**	**4**	**1**	**2**	