Mole Concept

1. MEASUREMENT IN CHEMISTRY

1.1 Every measured physical quantity consists of two parts viz a numerical value and the unit.
Physical quantity = numerical value $\times$ unit

1.2 The numerical value of a physical quantity is determined experimentally. Every scientific measurement has some degree of uncertainty due to two reasons.
(i) Skill of the observer
(ii) Limitation of measuring instrument

1.3 There are two ways of expressing this uncertainty.
(i) One method of expressing it is to use the notation $\pm$ 1 along with the doubtful digit.

Ex.1 : 63.7 $\pm$ 0.1cm.
(ii) Another method is expressing it is to use the concept of significant figures.

1.4 Rules for determining the number of significant figures:
Rule (i) All nonzero digits are significant

Ex.2 : 6.324 has four significant figures.
7.92 has three significant figures
1.2 has two significant figures.

Rule (ii) A zero appearing at a beginning of a number is not significant

Ex.3 : 0.32 has two significant figures
0.018 has two significant figures.
0.0004 has one significant figure.
0.324 has three significant figures.

Rule (iii) A zero appearing in the middle of a number or at the end of a number that included a decimal point is significant.

Ex.4 : 6.023 has four significant figures.
3.01 has three significant figures.
1.050 has four significant figures.
1.5000 has five significant figures.

Rule (iv) If a number ends in zeros but these zeros are not to the right side of a decimal point, then they may or may not be significant.

Ex.5 : 6500 may have two or three or four significant figures if 6500 is expressed as
(i) 6.5×10^3 then it has two significant figures.
(ii) 6.50×10^3 then it has three significant figures.
(iii) 6.500×10^3 then it has four significant figures.
(iv) In order to avoid the ambiguity in the rule
(v) generally very large and very small numbers are expressed in exponential form or scientific notation.
In this notation a number is expressed as $N \times 10^n$ where N = a number with at least a single nonzero digit to the left of the decimal point. n = an integer.

Ex.6 : In Avogadro's number 6.023×10^{23}. 6.023 is a number of four significant figures i.e., in the exponential notation, the numerical portion represents the number of significant figures.

1.5 Calculations involving significant figures.

Rule 1 : The result of addition or subtraction should be reported to the same number of decimal places as that of the term with least number of decimal places.

Addition example: 6.23 + 2.1 + 1.712 = 10.042
This value should be taken as 10.0 only because 2.1 has only one decimal place.

Subtraction example: 16.3215 – 2.706 = 13.6155
This value should be taken as 13.615 because 2.706 have only three decimal places.

Rule 2 : The result of multiplication or division should be reported to the same number of significant figures as possessed by the least precise term.

Ex.7 **: for multiplication:** 6.102 x 2.1 = 12.8142.
This value should be taken as 12 because 2.1 has two significant figures.

Ex.8 **: for division** : 5.2765/1.25 = 4.2212.
This value should be taken as 4.22 because 1.25 have three significant figures.

2. DIMENSIONAL ANALYSIS

2.1 The units of mass, length and time are independent units and as they are not derived from any other unit, these are called fundamental units.
System International (S.I.) has seven basic units.

Quantity	**Unit**	**Symbol**
Length	Meter	m
Mass	Kilogram	Kg
Time	Second	s
Temperature	Kelvin	K
Amount of substance	mole	mol.
Electric Current	ampere	A
Luminous intensity	Candela	Cd

2.2 **Derived Units :**

(i)	Area	=	length $\times$ breadth	=	m x m		= m^2
(ii)	Velocity	=	displacement / time	=	m / s		= ms^{-1}
(iii)	Acceleration	=	velocity / time	=	ms^{-1}/s		= ms^{-2}
(iv)	Force	=	mass acceleration	=	m x a		= Kg ms^{-2}
(v)	Pressure	=	Force / area	=	Kg m^{-1} s^{-2}		= Nm^{-2}
(vi)	Volume	=	length cube	=	m x m $\times$ m		= m^3
(vii)	Density	=	mass / volume	=	Kg/m^3		= Kg m^{-3}
(viii)	Energy /Work	=	Force $\times$ distance travelled	=	Kg m s^{-2} $\times$ m		= Kg m^2 s^{-2}
	Energy and work have same units.						
(ix)	Frequency	=	no. of cycles per second	=	cycles/second		= s^{-1} or Hz
(x)	Electric charge	=	current $\times$ time	=	ampere x second = A$\times$s = coulomb.		
(xi)	Power	=	Energy / time	=	Kg m^2 s^{-3}		= Watt
(xii)	Potential difference	=	Power / Ampere	=	Kg m^2 s^{-3} A^{-1}		= $JA^{-1}s^{-1}$

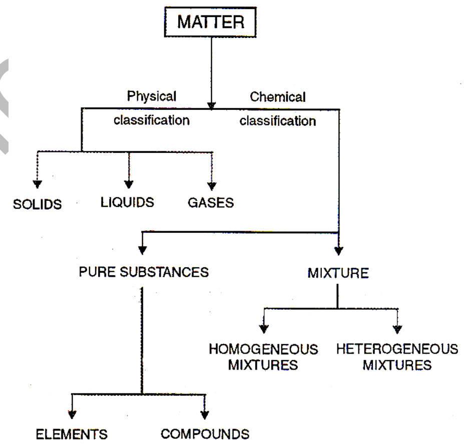
3. RELATION BETWEEN VARIOUS UNITS

Units of length :			
	1 mile	=	1760 yards
	1 yard	=	3 ft
	1 ft	=	12 inch
	1 inch	=	2.54 cm

	1 mile	=	1.609 km		
	1 mile	=	5280 ft		
Units of mass :	1 kg	=	10^3 g		
	1 lb	=	453.6 g		
	1 metric ton	=	1000 kg		
Units of volume :	1 m^3	=	10^3 L		
	1 dm^3	=	1 L		
	1 cm^3	=	1c.c	=	10^{-3} L
	1 ft^3	=	28.32 L		
	1 qt	=	0.9464 L		
Units Of Energy :	1 dyne	=	10^{-5} N		
	1 calorie	=	4.184 J		
	1 erg	=	10^{-7} J		
	1 e.V	=	1.6022×10^{-19} J		

4. CHEMICAL CLASSIFICATION OF MATTER

Anything that occupies space, possesses mass and the presence of which can be felt by any one or more of our five senses is called matter.



4.1 Matter may be

(i) homogeneous (uniform composition)

(ii) heterogeneous (nonuniform composition)

4.2 Homogeneous matter may be

(i) non fixed composition or solutions.

(ii) fixed composition or pure substances.

4.3 Pure substances may be

(i) elements (can not be decomposed)

(ii) compounds (can be decomposed by chemical reactions)

Element : A pure substance which can neither be decomposed into nor built from simpler substances is called Element.

Ex. 9 **:** Oxygen, sodium, aluminium, ferrum etc.,
Elements are further classified into i) Metals ii) Nonmetals, iii) Metalloids.

Compound : A substance which can be obtained by the union of two or more elements in a definite proportion by weight and into which it may be decomposed by suitable chemical methods.

Ex.10 **:** Carbon dioxide, water, methane etc.

5. LAWS OF CHEMICAL COMBINATION

5.1. The Law of conservation of mass

(i) This law was given by Lavoiser.

(ii) This law was verified by Landolt.

(iii) This law states "matter can neither be created nor destroyed in a chemical reaction, the mass of the reactants is equal to the mass of the products".

(iv) The exception to this law is nuclear reactions where Einstein equation is applicable.

Ex.11

$$2Ca \quad + \quad O_2 \quad \rightarrow \quad 2CaO$$

$$2 \times 40 \qquad 1 \times 32 \qquad 2(40+16)$$

$$= \ 80g \ = \ 32g \qquad = 112g$$

Total mass reactant = Total mass product = 112g.

5.2 The law of constant composition or definite proportion

(i) This law was given by Proust.

(ii) This law was verified by Stass & Richards.

(iii) This law states that "All pure samples of the same chemical compound contain the same elements combined in the same proportion by mass, irrespective of the method of preparation"

Ex.12 Different samples of carbon di oxide contain carbon and oxygen in the ratio of 3:8 by mass. Similarly in water ratio of weight of hydrogen to oxygen is 1:8.

5.3 The Law of multiple proportion

(i) This law was given by John Dalton.

(ii) The law was verified by Berzelius.

(iii) This law states that "when two elements A and B combine together to form, more than one compound, then several, masses of A which separately combine with a fixed mass of B, are in a simple ratio".

Ex.13

CO		and	CO_2
12:16			12:32
ratio	=	16 : 32	
		1 : 2	
H_2S		and	H_2S_2
2:32			2:64
ratio	=	32 : 64	
		1 : 2	

5.4 The Law of reciprocal proportions

(i) This law was given by Ritcher.

(ii) This law states that "when two elements combines separately with third element and form different types of molecules, their combining ratio is directly reciprocated if they combine directly."

Ex.14 C combines with O to form CO_2 and with H to form CH_4. In CO_2 12g of C reacts with 32g of O, whereas in CH_4 12g of C reacts with 4g of H. Therefore when O combines with H, they should combine in the ratio of 32 : 4 (i.e. 8:1) or in simple multiple of it. The same is found to be true in H_2O molecules. The ratio of weight of H and O in H_2O is 1:8.

5.5 The Law of Gaseous volume.

(i) This law was given by Gaylussac.

(ii) This law states that "when gas combine, they do so in volume which bear a simple ratio to each other and also to the product formed provided all gases are measured under similar conditions." Or in other words volume of reacting gasses and product gases have a simple numerical ratio to one another.

Ex.15

H_2 (g)	+	Cl_2 (g)	$\rightarrow$	2HCl (g)
1 unit vol.		1 unit vol.		2 unit vol.
				ratio = 1 : 1 : 2
N_2	+	$3H_2$	$\rightarrow$	$2NH_3$
1 unit vol.		3 unit vol.		2 unit vol.
				ratio = 1 : 3 : 2

5.6 The Avogadro Law

(i) This law states that "equal volume of all gaseous under similar conditions of temperature and pressure contain equal number of molecules".

$2H_2$	+	O_2	$\rightarrow$	$2H_2O$
2 vol.		1 vol.		2 vol.
				ratio 2 : 1 : 2

(ii) This law helped to remove anomaly between Dalton's atomic theory and Gay lussac's law of volume by making a clear distinction between atoms and molecules.

(iii) It reveals that common elements gases like hydrogen, nitrogen, oxygen etc. are diatomic.

(iv) It provides a method to determine the atomic mass of gaseous elements.

(v) It provides a relationship between vapour density and molecular mass of substances.

2 × vapour density (VD) = molecular mass of gas.

SOLVED PROBLEMS

1. 23.4g of NaCl on reacting with 68g of $AgNO_3$ formed 57.4g of AgCl and 34g of $NaNO_3$. This is in accordance with

(1) The law of conservation of mass. (2) The law of constant composition.

(3) The law of reciprocal proportion. (4) None of these

Ans. (1)

$$\underset{23.4g}{NaCl} + \underset{68g}{AgNO_3} \rightarrow \underset{57.4g}{AgCl} + \underset{34g}{NaNO_3}$$

mass of reactant = mass of product = 91.4g. Hence the law of conservation of mass is obeyed.

2. 8g of $CaCO_3$ on heating gave 3.25g CO_2 gas. The mass of residue left is -

(1) 4g (2) 4.48g (3) 12g (4) 16g

Ans. (2)

$$\underset{8g}{CaCO_3(s)} \rightarrow \underset{x}{CaO(s)} + \underset{3.52g}{CO_2(g)}$$

According to law of conservation of mass

mass of reactants = mass of products.

8 = x + 3.52 (where x is mass of residue)

x = 8 – 3.52 = 4.48g

3. In one experiment 12g of Mg combine with 8g of O_2 to form 20g of MgO. In another experiment when 150g of Mg combine with 100g of O_2 then 250g of MgO is formed. Above two experiment follows-

(1) The law of conservation of mass (2) The law of constant composition

(3) The law of definite proportion (4) All of the above

Ans. (4) I experiment $\dfrac{\text{mass of Mg reacted}}{\text{mass of } O_2 \text{ reacted}} = \dfrac{12}{8} = \dfrac{3}{2}$

II experiment $\dfrac{\text{mass of Mg reacted}}{\text{mass of } O_2 \text{ reacted}} = \dfrac{150}{100} = \dfrac{3}{2}$

Hence both law of conservation of mass and constant composition is obeyed.

4. H combines with O to form two compounds water (H_2O) and hydrogen peroxide (H_2O_2). If 2g of H reacts with O completely to form 18g of water and 34g of H_2O, then what is ratio of mass of O combining with H.

(1) 3 : 9 (2) 4 : 16 (3) 1 : 2 (4) None of these **Ans. (3)**

The ratio of masses of oxygen which combine with 2g of hydrogen to give H_2O and H_2O_2 are 16 : 32 = 1 : 2. That is, one mass of O is a multiple of the other mass of O combining with the same mass of H to form different compounds.

5. The vapour density of a gas is 11.2, then 11.2g of this gas at N.T.P. will occupy a volume-

(1) 1.12 litres (2) 0.112 litres (3) 112 litres (4) 11.2 litres **Ans. (4)**

Vapour density of any gas at N.T.P occupies a volume of 11.2 litres.

6. DALTON'S ATOMIC THEORY

6.1 Postulates.

(i) Every element consists of large number of small particles called atoms.
(ii) Atoms are invisible, indivisible, can neither be created nor destroyed.
(iii) Atoms of same element are identical in all respects such as size, shape, weight and properties.
(vi) Atoms of different elements combine in simple ratio such as 1 : 1, 1 : 2, 2 : 3 etc., to form compound atoms.
(v) The compound atoms (molecules according to modern approach) of the same compound are identical.

6.2 Limitations.

(i) It does not distinguish between the ultimate particles of an element and that of a compound.
(ii) It fails to explain the laws of combining volumes of gases.
(iii) It does not give the idea of isotopes and isobars.

7. ATOM, MOLECULES AND MOLECULAR FORMULA

7.1 Atom : It is the smallest particle of an element that takes part in a chemical reaction and is not capable of independent existence.

7.2 Molecule : It is the smallest particle of matter which is capable of independent existence. A molecule is generally an assembly of two or more tightly bonded atoms.

Homo atomic molecules : Molecules of an element containing one type of atoms only eg. H_2,N_2 etc.
Hetero atomic molecules : Molecules of compounds containing more that one type of atom
eg. H_2O, NH_3, SO_2 etc.

7.3 Mole :

(i) The quantitative aspect of a chemical reaction in chemistry is done by using the mole concept.
(ii) Mole is the chemical counting unit. It expresses the amount of a substance.
(iii) The word 'mole' (Latin = heap or pile) was introduced by Wilhelm Ostwald in 1896.
(iii) A mole is the amount of substance that contains an many entities (atoms, molecules or other particles) as there are atoms in exactly 12g (or 0.012 kg) of carbon-12 (C^{12}) isotope.
(iv) 1 mole has 6.022137 x 10^{23} entities (atoms, molecules, ions, protons, electrons etc.) This number of entities is constant and is known as Avogadro constant. It is represented by symbol N_A or N
(v) Mole is SI unit, mole can be used with a prefix.

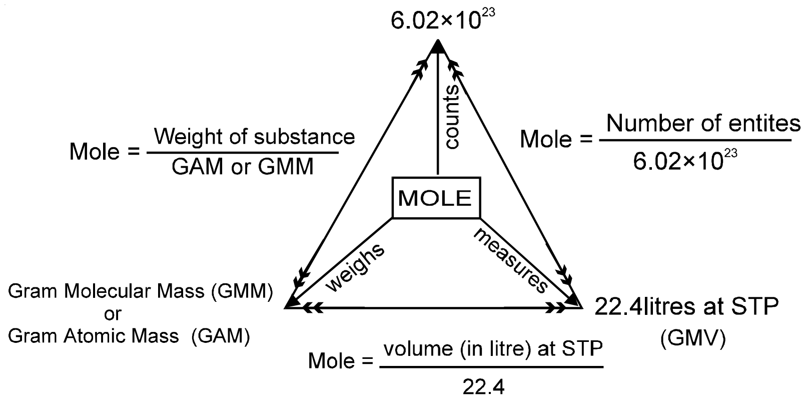
Ex.16 1mmol $= 10^{-3}$ mol
1µmol $= 10^{-6}$ mol
1 nmol $= 10^{-9}$ mol

IMPORTANT POINTS

☞ 1 Mole = 6.023 × 10^{23} particles
☞ 1 mole atoms = 6.023 × 10^{23} atoms
☞ One mole molecule = 6.023 × 10^{23} molecules
☞ Mass of one mole of atoms = Gram atomic mass (GAM)
☞ Mass of one mole of molecules = Gram molecular mass (GMM)
☞ Volume occupied by 1 mole of a gas at N.T.P. = 22.4 litres (Molar Volume)
☞ 1 mole of different substances have different masses.

Ex.17 1 mole C = 12 g
1 mole N_2 = 28g
1 mole of different substances have same volume at STP.
1 mole CO_2, N_2, H_2, He will have volume 22.4 L at STP.

8. MASS, MOLE AND NUMBER CONVERSION



9. ATOMIC MASS SCALE

Carbon as standard : The modern standard reference for atomic mass is carbon isotope of mass number 12.

$$\text{Atomic mass of an element} = \frac{\text{Mass of 1 atom of element}}{\frac{1}{12} \text{ x Mass of 1 atom of } C^{12}}$$

IMPORTANT POINTS

☞ Atomic mass is not a mass (weight) but a number

☞ Atomic mass is not absolute but relative to the mass of the standard reference element C^{12}.

10. MOLECULAR MASS

It is number of times a molecule is heavier than 1/12th of an atom of C^{12}.

$$\text{Molecular Mass} = \frac{\text{Mass of 1 molecule}}{\frac{1}{12} \text{ x Mass of one atom of } C^{12}}$$

IMPORTANT POINTS

☞ Molecular mass is not a mass (weight) but a number.

☞ Molecular mass is relative and not absolute.

☞ Molecular mass expressed in grams and is called Gram Molecular Mass (GMM).

☞ Molecular mass is calculated by adding all the atomic mass of all the atoms in a molecule.

Ex.18 $CO_2 = 12 + (2 \times 16) = 44$, $NH_4Cl = (14 \times 1) + (1 \times 4) + (1 \times 35.5) = 53.5$

11. MORE ABOUT ATOMIC MASS AND MOLECULAR MASS

11.1 Atomic Mass :

The relative atomic mass (atomic weight) of an element is the mass of one atom of the element compared with the mass of an atom of ${}_6C^{12}$ (carbon -12 isotope) taken as 12.0000 units.

Gram Atomic Mass (GAM) : The gram atomic mass of an element is the atomic mass of it, expressed in g.

Ex.19 1 gram atom of hydrogen = 1.008g
1 gram atom of carbon = 12g
1 gram atom of chlorine = 35.5g

NOTE : (i) Atoms of the same element which have different relative masses are called isotopes.
(ii) In case of isotopes, atomic mass of the elements is average of relative masses of different isotopes of the element.

Ex.20 There are two isotopes of chlorine.

	$_{17}Cl^{35}$	and	$_{17}Cl^{37}$
relative mass	35		37
relative abundance	3	:	1

$$\text{At. mass of element} = \frac{(\text{At. mass of I isotope} \times \text{relative abundance}_{I}) + (\text{At. mass of II isotope} \times \text{relative abundance}_{II})}{\text{Total relative abundance}}$$

$$\text{At. mass of Cl} = \frac{(35 \times 3) + (37 \times 1)}{3 + 1} = \frac{105 + 37}{4} = 35.5$$

11.1.1 Determination of Atomic mass :

atomic mass can be determined by Dulog and Petit's method (1819). It states the "the product of atomic mass and specific heat of an element is 6.4 approximately".

Mathematically : - Atomic mass × specific heat $\cong$ 6.4

Ex.21 The specific heat of metal is 1J/g/K. If equivalent mass of metal is 9. Calculate its exact atomic mass.

$$\text{specific heat} = 1\text{J/g/K} = \frac{6.4}{0.24} \text{ or } 0.24 \text{ cal /g/K}$$

$$\text{atomic mass (app.)} = \frac{6.4}{\text{specific heat}} = \frac{6.4}{0.24} = 26.75$$

$$\text{now valency} = \frac{\text{Atomic mass (app)}}{\text{Equivalent mass}} = \frac{26.75}{9} = 2.9 \simeq 3. \ (\because \text{ valency is integer})$$

Exact atomic mass = Equivalent mass × valance

= 9 × 3 = 27

Atomic mass can also be determined from molecular mass and atomicity.

Atomicity :- It may be defined as the number of atoms present in a molecule of an element.

$$\text{Atomic mass} = \frac{\text{molecular mass}}{\text{atomicity}}$$

11.2 MOLECULAR MASS

(i) Molecular mass (the relative molecular mass) :- The relative molecular mass (weight) of an element or compound is the mass of one molecule of the element or compound compared with the mass of atom of $_6C^{12}$ which is arbitrarily assigned as 12.0000.

(ii) Gram molecular mass (GMM) :- The molecular mass expressed in grams is called gram molecular mass (GMM), or gram mole or mole.

(iii) Molar volume or gram molecular volume (GMV) : - The volume occupied by one gram mole or one mole of a gas at STP is called molar volume or gram molecular volume (GMV).

11.2.1 DETERMINATION OF MOLECULAR MASS

(i) Regnault method

(a) By this method vapour density (V.D) of the gas are determined by direct weighing.

(b) $$\text{V.D.} = \frac{\text{weight of a certain vol. of gas or vapour under certain temp. and pressure}}{\text{weight of the same vol. of H under same temp. and pressure}}$$

(ii) Diffusion method :-

(a) It is based on Graham's Law of diffusion.

(b) Graham's Law states that "the rate of diffusion of different gases, under similar conditions of temperature and pressure, are inversely proportional to the square roots of their densities (molecular mass)"

(c) Mathematically : – $\frac{r_1}{r_2} = \sqrt{\frac{d_1}{d_2}} = \sqrt{\frac{M_1}{M_2}}$

(iii) Ideal Gas Law : -

(a) Mathematically $PV = \frac{w}{M}RT$ or $PV = nRT$

Here P, V,w, M, T, n and R are pressure, volume, mass of substance, molecular mass, absolute temperature, moles and ideal gas constant.

SOLVED PROBLEMS

6. Calculate the number of molecules of dinitrogen oxide in 0.044 kg of the gas

(1) 6.02×10^{23} (2) 3.01×10^{23} (3) 12×10^{23} (4) None of these

Ans. (1)

Gram molecular mass of dinitrogen oxide (N_2O) = 44g

Given mass - 0.044kg or 44g

A gram molecular mass of any gas contain Avogadro number of molecules = 6.023×10^{23}

0.044kg of dinitrogen oxide contain 6.023×10^{23} molecules

7. Which of the following contains the least number of molecules and which the highest -

(1) 17.75g of Cl_2 (2) 8g of CO (3) 4g of He (4) 28g of Fe

Ans. (3)

(1) Number of moles of $Cl_2 = \frac{\text{mass}}{\text{GMM}} = \frac{17.75}{71} = 0.25$

(2) Number of moles of $CO = \frac{\text{mass}}{\text{GMM}} = \frac{8}{28} = 0.285$

(3) Number of moles of $He = \frac{\text{mass}}{\text{GMM}} = \frac{4}{4} = 1$

(4) Number of moles of $Fe = \frac{\text{mass}}{\text{GMM}} = \frac{28}{56} = 0.5$

8. Atomic mass of neon is 20. (i) Calculate the number of atoms in 1g of neon and (ii) 1g atom of neon.

Sol.

(i) 20g of neon contains 6.023×10^{23} atoms

$\therefore$ 1g of neon contains = $\frac{6.02 \times 10^{23}}{20} = 3.01 \times 10^{22}$ atoms

(ii) 1g atom means 1 mole of neon atom, therefore number of neon atoms is 6.02×10^{23}

9. What is the mass of 1 molecule of dry ice ?

Sol. Gram molecular mass of CO_2 (dry ice) = 12 + 32 = 44g

6.023×10^{23} molecules of CO_2 weighs 44g

therefore 1 molecule of CO_2 weighs $\frac{1 \times 44}{6.02 \times 10^{23}} = 7.30 \times 10^{-23}$g

10. Calculate the number of atoms in each of the following
(i) 52 mole of He (ii) 52 amu of He (iii) 52g of He

Sol. (i) 1 mole He contain 6.02×10^{23} atoms
52 mole of He contain = $52 \times 6.02 \times 10^{23}$ atoms

(ii) Atomic mass of He = 4 amu

52 amu of He contain = $\frac{52}{4}$ = 13 atoms of He

(iii) Number of moles of He in 52g of = 13 moles
number of atoms in 52g of He i.e. 13 moles = $13 \times 6.02 \times 10^{23}$ atoms
= 78.26×10^{23} atoms

11. 6×10^{20} molecules of SO_2 are removed from 320 milligram of SO_2. What are the remaining moles of SO_2.
(1) 4×10^{-3} moles (2) 5×10^{-3} moles (3) 2×10^{-3} moles (4) 6×10^{-3} moles

Ans. (1) Mole in 320mg. of $SO_2 = \frac{320 \times 10^{-3}}{62} = 5 \times 10^{-3}$ moles ;

Moles of SO_2 removed are = $\frac{6 \times 10^{20}}{6 \times 10^{23}} = 10^{-3}$ moles

Remaining moles of $SO_2 = [5 \times 10^{-3} - 10^{-3}] = 4 \times 10^{-3}$ moles

12. CHEMICAL FORMULA

It is of two types -

12.1 Molecular formula : Chemical formula that indicate the actual number and type of atoms in a molecule are called molecular formula eg. - Molecular formula of benzene is C_6H_6

12.2 Empirical formulae : The chemical formulae that give only the relative number of atoms of each type in a molecule are called empirical formulae eg. - empirical formula of benzene is CH.

12.3 Determination of Chemical Formulae :

Determination of empirical formulae :

Step - I : Determination of percentage
Step - II : Determination of mole ratio
Step - III : Making it whole number ratio
Step - IV : Removal of fractions from mole ratio (to obtain empirical formula)

12.4 Determination of molecular formula :

Step - V : Molecular formula = (Empirical formula) n

$$\text{where} \quad n = \frac{\text{Molecular mass (weight)}}{\text{Emperical mass (weight)}}$$

Note : To find percentage of oxygen in organic compound add percentage of all other atoms and subtract it from 100
As for above example % of O = 100 - (51.4 + 4.3 + 12.8 + 9.8 + 7.0) = 14.7

Ex.22 A compound of carbon, hydrogen and nitrogen contains three elements in the respective ratio of 9 : 1 : 3.5. Calculate the empirical formula. If the molecular weight is 108, what is molecular formula.

Sol.

Element	Ratio by weight	Atomic ratio	Least ratio
C	9	$\frac{9.0}{12} = 0.75$	3
H	1	$\frac{1.0}{1} = 1$	4
N	3.5	$\frac{3.5}{14} = 0.25$	1

Empirical formula = C_3H_4N

Empirical formula mass = (12 × 3) + (1 × 4) + (14 × 1) = 54

Molecular formula mass = 108

where $N = \dfrac{\text{molecular mass}}{\text{emperical mass}} = 2$

hence molecular formula = 2 × C_3H_4N = $C_6H_8N_2$

13. CHEMICAL EQUATION

Representation of the chemical change in terms of symbol and formulae of the reactants and products is called a chemical equation.

13.1 Information conveyed by a chemical equation :

(i) Qualitatively, a chemical equation tells us the names of the various reactants

(ii) Quantitatively, it express

(a) The relative number of molecules of reactants and products

(b) The relative number of moles of reactant and products.

(c) The relative masses of reactants and products.

(d) The relative volumes of gaseous reactants and products

14. STOICHIOMETRY

(i) *Stoichiometry* is a Greek word (*stoicheio* = element and *metron* = element)

(ii) Stoichiometry is a calculation of the quantities of reactant and product involved in a chemical reaction.

(iii) Stoichiometry can be classified into two groups -

(a) Gravimetric Analysis (b) Volumetric analysis

14.1 Stoichiometry and Problem Solving

In problem solving we shall first discuss gravimetric analysis of chemical reaction. In gravimetric analysis we relate the weights of two substances or a weight of a substance with a volume of a gas or volumes of two or more gases.

14.2 Problem involving Mass-Mass relationship-

Ex.23 What amount of MgO is formed when 12g of Mg reacts with oxygen completely.

Sol. Following are the steps to solve the above problem where mass of reactant is given and mass of product is to be calculated.

Step 1 Write balance equation

$Mg + O_2 \rightarrow MgO$ (unbalanced reaction)

$2Mg + O_2 \rightarrow 2MgO$ (balanced reaction)

Step 2 Write the moles below the formula

$2Mg + O_2 \rightarrow 2MgO$

moles 2 1 2 (this represents simplest molar ratio among reagents.)

Step 3 Write the relative weights of the reactant and product

2Mg	+ O_2	$\rightarrow$	2MgO
(2 x 24)			2 x (24+16)
= 48g			= 80g

Step 4 Apply unitary method

48g of Mg gives 80g of MgO

$\therefore$ 12g of Mg gives $\dfrac{12 \times 80}{48}$ g of MgO = 20g of MgO

14.3 Problem involving Mass-Volume relationship.

Ex.24 By heating 10g $CaCO_3$, 5.6g of CaO is formed. What volume of CO_2 obtained in this reaction at STP.

Sol. Step 1 Write balance equation

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Step 2 Write the moles below the formula

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

moles 1 1 1

Step 3 Write the relative weights of the reactant and volume of product

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

100g 22.4L at STP

Step 4 Apply unitary method

100g of $CaCO_3$ gives 22.4L of CO_2

$\therefore$ 10g $CaCO_3$ gives $\frac{10 \times 22.4}{100} = 2.24L$ of CO_2 at STP

14.4 Problem involving Volume-Volume relationship-

Ex.25 Hydrogen reacts with nitrogen to produce ammonia according to this equation.

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

Sol. Determine how much ammonia would be produced if 200L of hydrogen react completely with nitrogen to form ammonia.

Step 1 Write moles below the balance equation

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

3 1 2

Step 2 Write relative volume of reactants and product

$3H_2(g)$	+	$N_2(g)$	$\rightarrow$	$2NH_3(g)$
3×22.4		1×22.4		2×22.4
= 67.2L		= 22.4L		= 44.8L

Step 3 Apply unitary method

67.2L of H_2 gives 44.8L of NH_3

$\therefore$ 200L of H_2 gives $\frac{200 \times 44.8}{67.2} = 133.3 \text{ L of } NH_3$

Note : Quantity of a substance consumed or produced can be determined only if we use a balance chemical equation.

15. LIMITING REAGENT

(i) Limiting Reagent (reactant) : The reactant which is completely consumed during the reaction.

(ii) Excess Reagent (reactant) : The reactant that is not completely consumed in a reaction.

The moles of product formed are always determined by the initial moles of limiting reagent.

12. Calculate the weight of iron which will be converted into its oxide by the action of 18g of steam.

$$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$$

Sol.

$$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$$

3Fe	$4H_2O$	Fe_3O_4
3×56	4×18	1×232
= 168	= 72	= 232

now 72g of steam (H_2O) reacts with 168g of Fe

$\therefore$ 18g of steam will react with $\frac{18 \times 168}{72} = 42g$ of Fe

13. Calculate the weight of lime (CaO) obtained by heating 200kg of 90% pure limestone ($CaCO_3$)

Sol.

100kg impure sample has $CaCO_3$ = 90 kg

$\therefore$ 200kg impure sample has $CaCO_3 = \frac{90 \times 200}{100}$ = 180 kg

now $\underset{100g}{CaCO_3} \rightarrow \underset{56g}{CaO} + \underset{44g}{CO_2}$

100kg $CaCO_3$ gives 56kg of CaO

$\therefore$ 180kg of $CaCO_3$ gives $\frac{56 \times 180}{100} = 100.8$kg of CaO

14. Oxygen is prepared by catalytic decomposition of potassium chlorate ($KClO_3$). Decomposition of $KClO_3$ gives potassium chloride (KCl) and oxygen (O_2). If 4.2 mole of oxygen is needed for an experiment, how many grams of $KClO_3$ must be decomposed.

Sol.

Step 1 $KClO_3(s) \rightarrow KCl(s) + O_2(g)$

Step 2 $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$
2 mole 3 mole

Step 3 (2 x 122.5) 3 mole
= 245g 3 mole

Step 4 If 2.45g of $KClO_3$ gives 3 mole of O_2 then 4.2 mole of O_2 will be obtained from

$= \frac{4.2 \times 245}{100}$ = 343g of $KClO_3$

15. A gaseous alkane is exploded with oxygen. The volume of O_2 for complete combustion to CO_2 formed is in the ratio of 7 : 4. Deduce molecular formula of alkane.

Sol. Let the formula of alkane be C_nH_{2n+2}

$$C_nH_{2n+2} + \left[n + \frac{n+1}{2}\right]O_2 \rightarrow nCO_2 + (n+1)H_2O(l)$$

Given $\frac{\text{volume of } O_2 \text{ used}}{\text{volume of } CO_2 \text{ formed}} = \frac{7}{4}$

$$\frac{n + (n+1)/2}{n} = \frac{7}{4}$$

n = 2. The alkane is C_2H_6.

16. 10 moles SO_2 and 15 moles O_2 were allowed to react over a suitable catalyst. 8 moles of SO_3 were formed. The remaining moles of SO_2 and O_2 respectively are-

(1) 2 moles, 11 moles (2) 2 moles, 8 moles (3) 4 moles, 5 moles (4) 8 moles, 2 moles

Ans. (1)

$2SO_2$	+	O_2	$\rightarrow$	$2SO_3$
10		15		0
10-2x		15-x		2x
2x =8		x = 4		

Hence remaining SO_2 = 10 - 8 = 2 moles, O_2 = 15- 4 = 11 moles

17. If 0.5 mole of $BaCl_2$ is mixed with 0.2 mole of Na_3PO_4 the maximum amount of $Ba_3(PO_4)_2$ that can be formed is-

(1) 0.70 mol (2) 0.50 mol (3) 0.20 mol (4) 0.10 mol

Ans. (4)

	$3\,BaCl_2$ +	$2Na_3PO_4$	$\rightarrow$	$6NaCl$	+	$Ba_3(PO_4)_2$
molar ratio	3	2		6		1
initial moles	0.5	0.2		0		0

Limiting reagent is Na_3PO_4 hence it would be consumed and the yield would be decided by its initial moles. 2 moles of Na_3PO_4 give 1 mole of $Ba_3(PO_4)_2$

$\therefore$ 0.2 moles of Na_3PO_4 would give 0.1 mole of $Ba_3(PO_4)_2$.

16. STOICHIOMETRY OF REACTIONS IN SOLUTIONS

(i) Reactions of solutions are by the most common .

(ii) Volumetric analysis is the method to deal with quantitative analysis involving solution.

(iii) The most commonly used concept to express the composition of solution is **molarity**.

16.1 Molarity

The number of gram moles of the solute dissolved per litre of the solution. It is denoted by 'M'.

$$M = \frac{\text{Weight}}{\text{Molecular weight}} \times \frac{1}{\text{V (litre)}}$$

$$M = \frac{\text{Weight}}{\text{Molecular weight}} \times \frac{1000}{\text{V (ml)}}$$

(i) If density and weight percentage of the solution is given then

$$M = \frac{10 \times d \times \text{percent}}{\text{GMM}}$$ where d = density of solution,

Number of millimoles = M x V (ml)

$\frac{M}{2}$	$\frac{M}{5}$	$\frac{M}{10}$	$\frac{M}{100}$	$\frac{M}{1000}$	M	5M	10M
Semi molar solution	Penti molar solution	Deci molar solution	Centi molar solution	Milli molar solution	Molar solution	Penta molar solution	Deca molar solution
$\frac{1}{2}$ mol/L	$\frac{1}{5}$ mol/L	$\frac{1}{10}$ mol/L	$\frac{1}{100}$ mol/L	$\frac{1}{1000}$ mol/L	1 mol/L	5mol/L	10mol/L

Ex.26 4g NaOH is present in 100ml of its aqueous solution. What is the molarity :-

(1) 2M (2) 1M (3) 10M (4) 0.1M **Ans. (2)**

Sol. $\text{Molarity} = \frac{w}{\text{GMM}} \times \frac{1000}{\text{volume (mL)}} = \frac{4}{40} \times \frac{1000}{100} = 1$

1M solution of NaOH

Ex.27 The solution of H_2SO_4 contains 80% by mass. Specific gravity (density) of solution is 1.71 g/cc. Find its Molarity.

Sol. $M = \frac{10 \times d \times \text{percent}}{\text{GMM}}$ $M = \frac{10 \times 1.71}{98} \times 80 = 13.95$

Ex.28 To neutralizes 20mL NaOH, the volume of 1M HNO_3 is–

(1) 4mL (2) 3mL (3) 2mL (4) 1mL

Sol. NaOH HNO_3

$M_1V_1 = M_2V_2$

$0.2 \times 20 = 1 \times V_2$ $V_2 = 4mL$

17. MORE ABOUT EXPRESSION OF STRENGTH/ CONCENTRATION OF SOLUTION

"The amount of solute which is dissolved in unit volume of solution is called concentration of solution."

$$\text{Concentration} = \frac{\text{amount of solute}}{\text{volume of solution}}$$

17.1 Weight - weight percent (w/W) : Weight of solute present in 100g of the solution.

$$\text{Weight percent} = \frac{\text{weight of solute (g)}}{\text{weight of solution (g)}} \times 100 . \qquad \text{weight percent } \frac{w}{W} \times 100$$

Ex.29 What is the weight percentage of NaCl solution in which 20g NaCl is dissolved in 60g of water.

(1) 10% (2) 5% (3) 25% (4) 15%

Ans. (4) $\text{weight percentage of NaCl} = \frac{\text{weight of NaCl}}{\text{weight of solution}} \times 100$

$$= \frac{20}{20+60} \times 100 = 25\% \text{ NaCl solution (w/W)}$$

17.2 Volume - volume percent (v/V) : (In liquid - liquid solution)

Volume of solute in ml. present in 100ml of the solution is called volume - volume percentage.

$$\text{Volume - volume percentage} = \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times 100$$

$$\text{volume percent} = \frac{v}{V} \times 100$$

Ex.30 A solution is prepared by mixing of 10ml ethanol with 120ml of methanol. What is volume percentage of ethanol:-

(1) 10% (2) 7.7% (3) 20% (4) 15%

Ans. (2) $\text{Volume percentage of ethanol} = \frac{\text{volume of ethanol}}{\text{volume of solution}} \times 100 = \frac{10}{10+120} \times 100 = 7.7\%$

17.3 Weight - volume percent (w/V) :

Weight of solute in g present in 100mL of the solution is called weight - volume percentage.

$$\text{weight - volume percentage} = \frac{\text{weight of solute (g)}}{\text{volume of solution (ml)}} \times 100$$

$$\text{percent of strength} = \frac{w}{V} \times 100$$

Ex.31 What is weight volume percentage of a solution in which 7.5g of KCl is dissolved in 100mL of the solution–

(1) 7.5% (2) 92.5% (3) 50% (4) none

Sol. 7.5% of KCl (w/V) : 7.50g KCl present in 100mL of the solution.

$$\frac{7.5}{100} \times 100 = 7.5\%$$

17.4 Normality.

The number of gram equivalents of the solute dissolved per litre of the solution. It is denoted by 'N' :

$$N = \frac{\text{Weight}}{\text{Equivalent weight}} \times \frac{1}{V(\text{litre})}$$

$$N = \frac{\text{Weight}}{\text{Equivalent weight}} \times \frac{1000}{V(\text{ml})}$$

(i) If density and weight percent of the solution is given then :-

$$N = \frac{10 \times d}{\text{GEM}} \times \text{percent} \quad \text{(where d = density of solution)}$$

$$\text{or } N = \frac{10 \times d}{\text{GEM}} \text{ x } \frac{100}{W}$$

$$\because \text{percent} = \frac{\text{weight of solute (w)}}{\text{weight of solution (W)}} \times 100$$

Number of equivalent = N x V(litre)
Number of milli equivalent = N x V(ml)
GEM = Gram equivalent mass of solute

$\frac{N}{2}$	$\frac{N}{5}$	$\frac{N}{10}$	$\frac{N}{100}$	$\frac{N}{1000}$	N	5N	10N
Semi normal	Penti normal	Deci normal	Centi normal	Milli normal	normal	Penta normal	Deca normal
$\frac{1}{2}$ eq./L	$\frac{1}{5}$ eq./L	$\frac{1}{10}$ eq./L	$\frac{1}{100}$ eq./L	$\frac{1}{1000}$ eq./L	1eq./L	5 eq./L	10eq./L

Ex.32 0.56g KOH is present in 100 mL of the solution what is the normality :-

(1) 1N (2) 0.1N (3) 2N (4) 0.2N

Ans - (2)

$$\text{Normality} = \frac{w}{\text{GEM}} \times \frac{1000}{V(\text{mL})} = \frac{0.56}{56} \times \frac{1000}{10} = 0.1\text{N}$$

Ex.33 Find the number of milliequivalent of H_2SO_4 persent in 50mL of N/20 H_2SO_4–

Sol. meq = N $\times$ V(mL) = 1/20 $\times$ 50 = 2.5

Ex.34 To prepare 600mL of 2N solution of NH_4OH, what volume of 10N NH_4OH is required –

Sol. $N_1V_1 = N_2V_2$

$2 \times 600 = 10 \times V_2$

V_2 = 120mL

Ex.35 To dissolve 3.3g of certain metal 110mL of N H_2SO_4. Find the equivalent mass of metal–

(1) 15g (2) 30g (3) 20g (4) 10g

Sol. meq. of H_2SO_4 = 110 $\times$ 1 = 110

Equivalent of H_2SO_4 = 110×10^{-3}

now according to law of equivalents of H_2SO_4 = equivalents of metal

$\therefore$ equivalent of metal = 110×10^{-3}

now equivalent = $\frac{\text{weight}}{\text{GEM}}$ $\quad 110 \times 10^{-3} = \frac{3.3}{\text{GEM}}$ $\quad$ GEM = 30

Ex.36 What volume of 2N and 5N H_2SO_4 should be mixed so that the resultant solution of 1L has normality = 3

Sol. $N_1V_1 = N_2V_2 = N_3V_3$ where $V_3 = V_2 + V_1$

let us assume the volume of 2N H_2SO_4 is x litre

then the volume of 6N H_2SO_4 is (1 – x)L

$2 \times x + 6(1–x) = 3 \times 1$; $2x + 6 – 6x = 3$; x = 3/4 or 0.75L

17.5 Molality :

The number of gram moles of solute dissolved in 1000g or 1kg of the solvent. It is denoted by 'm'.

$$m = \frac{\text{Weight}}{\text{Molecular weight}} \times \frac{1000}{V(\text{ml})}$$

$$m = \frac{\text{Weight}}{\text{Molecular weight}} \times \frac{1}{V(\text{litre})}$$

Ex.37 10g HCl dissolved in 250mL of its aqueous solution. If density of the solution is 1.2g/mL. than molality of the solution will be :-

(1) 1 (2) 0.34 (3) 0.945 (4) 3.4

Ans - (3)

Weight of solute = 10g Volume of solution = 250mL

Density of solution = 1.2g/mL. $\therefore$ Weight of solution = $250 \times 1.2 = 300$g

$\therefore$ Weight of solvent = weight of solution - weight of solute = 300 - 10 = 290g

$$\therefore \quad m = \frac{w}{GMM} \times \frac{1000}{W} = \frac{10}{36.5} \times \frac{1000}{290} = 0.945$$

17.6 Formality

The number of gram formula weight of a solute dissolved per litre of the solution is called formality of the solution. It is denoted by 'F'.

$$\text{Formality} = \frac{\text{mas of solute(g)}}{\text{formula mass of solute}} \times \frac{1}{\text{Volume of solution(L)}}$$

(i) $F = \frac{w}{f} \times \frac{1}{V(L)}$ (ii) $F = \frac{w}{f} \times \frac{1000}{V(mL)}$ (iii) $F = n_f \times \frac{1}{V(L)}$

where w = mass of solute, f = formula mass of solute, V = volume of solution,

n_f = number of gram formula mass.

Ex.38 CH_3COOH exists as dimer in benzene 1.2g of the acid was dissolved and the volume was made up to one litre by benzene, what is the formality -

(1) 0.1F (2) 0.01F (3) 1F (4) 10 F

Ans - (2)

Molecular mass of CH_3COOH = 60

Formula mass of the associated molecule of the acid = $2 \times 60 = 120$

mass of CH_3COOH = 1.2g

volume of solution = 1L Formality $= \frac{1.2}{120} \times \frac{1}{1} = 0.01F$

17.7 Mole Fraction

The mole fraction of a component in a solution is the ratio of the number of moles of that component to the total number of moles present in the solution.

Suppose : – $\left.\begin{matrix} \text{A – Solute} \\ \text{B - Solvent} \end{matrix}\right\}$ Solution n_A = number of moles of solute n_B = number of moles of solvent

Then mole fraction of solute = $X_A = \frac{n_A}{n_A + n_B}$

Mole fraction of solvent = $X_B = \frac{n_B}{n_A + n_B}$ $X_A + X_B = 1$

Ex.39 1 molal aqueous solution of any solute will have mole fraction–
(1) 1 (2) 1.8 (3) 18 (4) 0.018

Sol. Mole fraction = $\frac{n_A}{n_A + n_B}$ for 1 molal solution number of A = 1

mass of H_2O in aqueous solution of 1000g.

$\therefore\ n_B = \frac{1000}{18} = 55.4 \quad = \frac{1}{1+55.4} = \frac{1}{56.4} = 0.018$

For gaseous mixture :

A binary system of two gases A & B

P_A = Partial pressure of A , P_B = Partial pressure of B

$P = P_A + P_B$ = Total pressure of gaseous mixture

Mole fraction of gas A $X_A = \frac{P_A}{P_A + P_B} = \frac{P_A}{P}$

Mole fraction of gas B $X_B = \frac{P_B}{P_A + P_B} = \frac{P_B}{P}$

Mole Percentage :

Mole percentage = Mole fraction x 100

Mole percent of A = $X_A \times 100$

Mole percent of B = $X_B \times 100$

17.8 ppm (Part per million)

The parts of the component per million parts (10^6) of the solution.

$$ppm = \frac{w}{w + W} \times 10^6$$

where w = weight of solute, W = weight of solvent

18. FORMULA FOR VOLUMETRIC CALCULATIONS

1. (i) Concentration (g/L) = $\frac{\text{weight of solute(g)}}{\text{volume of solution(L)}}$

(ii) Concentration (g/L) = Molarity $\times$ Molecular weight

(iii) Concentration (g/L) = Normality $\times$ Equivalent weight

(iv) Normality $\times$ Equivalent weight = Molarity $\times$ Molecular weight = Acidity of base
Normality of acid = Molarity $\times$ Basicity
Normality of base = Molarity $\times$ Acidity

(v) Normality = Molarity = Formality
The relation is true for a substance having
Eq. wt = Mol. wt. and the substance does not undergo association or dissociation.

(vi) Molarity is independent of Temperature.

Solved Example

Exp–1 How many g atom and no. of atoms are there in (a) 60 g carbon (b) 224.4 g Cu ?
Given At. weights of C and Cu are 12 and 63.6 respectively. Avogadro's no. = 6.02×10^{23}.

Sol. $\because$ g atom = $\frac{\text{wt.}}{\text{at. wt.}}$ and No. of atoms = $\frac{\text{wt.} \times \text{Av. No.}}{\text{at. wt.}}$

(a) $\therefore$ **For 60 g C :** g atom = $\frac{60}{12} = 5$

No. of atoms = $\frac{60 \times 6.02 \times 10^{23}}{12}$ = $\mathbf{30.1 \times 10^{23}}$

(b) For 224.4 g Cu : g atom = $\frac{224.4}{63.6} = 3.53$

No. of atoms = $\frac{224.4 \times 6.02 \times 10^{23}}{63.6}$ = $\mathbf{21.24 \times 10^{23}}$

Exp–2 Find the no. of g atoms and weight of an element having 2×10^{23} atoms. At. weight of element is 32.

Sol. $\because$ N atoms have 1 g atom

$\therefore$ 2×10^{23} atoms have = $\frac{2 \times 10^{23}}{6.023 \times 10^{23}}$ = **0.33g atom**

$\because$ N atoms of element weigh 32 g

$\therefore$ 2×10^{23} atoms of element weight = $\frac{32 \times 2 \times 10^{23}}{6.023 \times 10^{23}}$ = **10.63 g**

Exp–3 Calculate the no. of atoms and volume of 1 g He gas at NTP.

Sol. $\because$ 4 g He has 6.023×10^{23} atoms

$\therefore$ 1 g He has $\frac{6.023 \times 10^{23}}{4}$ atoms = $\mathbf{1.506 \times 10^{23}}$ **atoms**

Also,

$\because$ 4 g He has volume at NTP = 22.4 litre

$\therefore$ 1 g He has volume at NTP = $\frac{22.4}{4}$ = **5.6 litre**

Exp–4 How many mole and molecules of O_2 are there in 64 g O_2 ? What is the mass of one molecule of O_2 ?

Sol. $\because$ 32 g O_2 has mole = 1

$\therefore$ 64 g O_2 has mole = $\frac{64 \times 1}{32}$ = **2 mole**

$\because$ 32 g O_2 contain 6.023×10^{23} molecules

$\therefore$ 64 g O_2 contain $\frac{6.023 \times 10^{23} \times 64}{32}$ = $\mathbf{12.04 \times 10^{23}}$ **molecules**

$\therefore$ N molecules of O_2 weight 32 g

$\therefore$ 1 molecules of O_2 weights = $\frac{32}{6.023 \times 10^{23}}$ = $\mathbf{5.313 \times 10^{-23}}$ **g**

Exp–5 How many years it would take to spend Avogadro's number of rupees at the rate of 10 lac rupees per second ?

Sol. Total rupees to be expanded = 6.023×10^{23}

Rupees spent per second = 10^6

$\therefore$ Rupees spent per year = $10^6 \times 60 \times 60 \times 24 \times 365$

$\because$ $10^6 \times 60 \times 60 \times 24 \times 365$ Rupees are spent in 1 year

$\therefore$ 6.023×10^{23} " " " $= \dfrac{6.023 \times 10^{23}}{10^6 \times 60 \times 60 \times 24 \times 365}$ = **1.9099 × 10¹⁰ year**

Exp–6 Weight of one atom of an element is 6.644×10^{-23} g. Calculate g atom of element in 40 kg.

Sol. Wt. of 1 atom of element = 6.644×10^{-23} g

$\therefore$ Wt. of N " " " $= 6.644 \times 10^{-23} \times 6.023 \times 10^{23} = 40$

$\because$ 40 g weight of element has a 1 g atom

$\therefore$ 40×10^3 g " " " $= \dfrac{40 \times 10^3}{40}$ = **10^3 g atom**

Exp–7 How many g of S are required to produce 100 mole and 100 g H_2SO_4 separately ?

Sol. $\because$ 1 mole of H_2SO_4 has = 32 g S

$\therefore$ 100 mole of H_2SO_4 = 32 x 100 = **3200 g S**

$\because$ 98 g H_2SO_4 has S = 32 g

$\therefore$ 100 g H_2SO_4 has S = $\dfrac{32 \times 100}{98}$ = **32.65 g S**

Exp–8 An alloy has Fe, Co and Mo equal to 71%, 12% and 17% respectively. How many cobalt atoms are there in a cylinder of radius 2.50 cm and a length of 10.0cm. The density of alloy is 8.20 g/mL. Atomic weight of cobalt = 58.9.

Sol. Weight of alloy cylinder = Volume x density

$= \pi r^2 h \times d$

$= \dfrac{22}{7} \times (2.5)^2 \times 10 \times 8.20 = 1610.7\,g$

Weight of cobalt in alloy = $\dfrac{1610.7 \times 12}{100} = 193.3\,g$

$\because$ 58.9 g cobalt has atoms = 6.023×10^{23}

$\therefore$ 193.3 g cobalt has atoms = $\dfrac{6.023 \times 10^{23} \times 193.3}{58.9}$ = **19.8×10^{23}**

Exp–9 The dot at the end of this sentence has a mass of about one microgram. Assuming that black stuff is carbon, calculate approximate atoms of carbon needed to make such a dot.

Sol. Mass of carbon in dot = 1×10^{-6} g

$\because$ 12 g C has 6.023×10^{23} atoms

$\therefore$ 1×10^{-6} g C has $\dfrac{6.023 \times 10^{23} \times 10^{-6}}{12}$ = **5×10^{16} atoms of C**

Exp–10 What is the molecular weight of a substance, each molecule of which contains 9 carbon atoms. 13 hydrogen atoms and 2.33×10^{-23} g of other component ?

Sol. Molecule has C, H and other component

$\therefore$ Wt. of 9 C atoms = 12 x 9 = 108 amu

$\because$ Wt. of 13 H atoms = 13 x 1 = 13 amu

Wt. of 2.33×10^{-23} g of other atom = $\dfrac{2.33 \times 10^{-23}}{1.66 \times 10^{-24}} = 14.04\,amu$

$\therefore$ Total weight of one molecule = 108 + 13 + 14.04 = **135.04 amu**

$\because$ Mol. weight = **135.04 g**

Exp–11 A plant virus is found to consist of uniform cylindrical particles of 150 Å in diameter and 5000 Å long. The specific volume of the virus is 0.75 cm^3/g. If the virus is considered to be single particle, find its molecular weight

Sol. Volume of virus

$$= \pi r^2 \ell = \frac{22}{7} \times \frac{150}{2} \times \frac{150}{2} \times 10^{-16} \times 5000 \times 10^{-8} = 0.884 \times 10^{-16}\ cm^3$$

$\therefore$ Weight of one virus = $\frac{0.884 \times 10^{-16}}{0.75}$ g = $\mathbf{1.178 \times 10^{-16}\ g}$

$\because$ Mol. wt. of virus = $1.178 \times 10^{-16} \times 6.023 \times 10^{23}$ = $\mathbf{7.095 \times 10^7}$

Exp–12 K-40 is a naturally occurring radioactive isotope having natural abundance 0.012% of potassium isotopes. How many K-40 atoms do you ingest by drinking one cup of whole milk containing 370 mg K?

Sol. Amount of K-40 in 370 mg K = $\frac{370 \times 0.012}{100}$ mg

= 0.0444 mg

$\because$ 40 g K-40 has atoms of K-40 = 6.023×10^{23}

$\therefore$ 0.0444×10^{-3} g K-40 has atoms = $\frac{6.023 \times 10^{23} \times 0.0444 \times 10^{-3}}{40}$ = $\mathbf{6.69 \times 10^{17}}$ **atoms**

Exp–13 Insulin contains 3.4% sulphur. Calculate minimum mol. wt. of insulin.

Sol. For minimum mol. wt., insulin must have at least one S atom in its one molecule.

$\because$ 3.4 g S then mol. wt. of insulin = 100

$\therefore$ 32 g S then mol. wt. of insulin = $\frac{100 \times 32}{3.4}$ = 941.176

$\therefore$ Minimum mol. wt. of insulin = **941.176**

Exp–14 Haemoglobin contains 0.25% iron by weight. The molecular weight of Haemoglobin is 89600. Calculate the no. of iron atom per molecule of Haemoglobin.

Sol. $\because$ 100 g Haemoglobin has = 0.25 g Fe

$\therefore$ 89600 g Haemoglobin has = $\frac{0.25 \times 89600}{100}$ = 224 g Fe

i.e. 1 mole or N molecules of Haemoglobin has

$$= \frac{224}{56} \text{ g atom Fe } = 4 \text{ atom Fe}$$

$\therefore$ **1 molecule of Haemoglobin has 4 atom of Fe.**

Exp–15 P and Q are two elements which forms P_2Q_3 and PQ_2. If 0.15 mole of P_2Q_3 weighs 15.9 g and 0.15 mole of PQ_2 weighs 9.3 g, what are atomic weight of P and Q ?

Sol. Let at. wt. of P and Q are a and b respectively.

$\therefore$ Mol. wt. of P_2Q_3 = 2a + 3b

and Mol. Wt. of PQ_2 = a + 2b

Now given that 0.15 mole of P_2Q_3 weigh 15.9 g

$$(2a + 3b) = \frac{15.9}{0.15} \qquad \left(\because \frac{\text{wt.}}{\text{mol. wt.}} = \text{mole}\right)$$

Similarly, $(a + 2b) = \frac{9.3}{0.15}$

Solving these two equations

$b = \mathbf{18}$ $\quad$ $a = \mathbf{26}$

Exp–16 Calculate the weight of iron which will be converted into its oxide by the action of 18 g of steam.

$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$

Sol. The reaction occurs as :

$$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$$

Mole ratio of reaction suggests :

$$\frac{\text{Mole of Fe}}{\text{Mole of } H_2O} = \frac{3}{4}$$

$\therefore$ Mole of Fe $= \frac{18}{18} \times \frac{3}{4} = \frac{3}{4}$

$\therefore$ Weigh of Fe $= \frac{3}{4} \times 56 = \mathbf{42\ g}$

Exp–17 Calculate the volume of O_2 and volume of air needed for combustion of 1 kg carbon at STP.

Sol. $C + O_2 \rightarrow CO_2$

$\because$ 12 g C requires O_2 = 22.4 litre = 1 mole = 32 g

$\therefore$ 1000 g C requires $O_2 = \frac{22.4 \times 1000}{12}$ litre

$= 1866.67$ litre O_2

$\therefore$ $V_{air} = 5 \times V_{O_2}$

$= 5 \times 1866.67 = \mathbf{9333.35\ litre}$

Exp–18 Calculate the weight of lime (CaO) obtained by heating 200 kg of 95% pure limestone ($CaCO_3$).

Sol. 100 kg impure sample has $CaCO_3$ = 95 kg

$\therefore$ 200 kg impure sample has $CaCO_3 = \frac{95 \times 200}{100} = 190$ kg

Now $\quad CaCO_3 \rightarrow CaO + CO_2$

M. wt. $\quad$ 100 g $\quad$ 56 g $\quad$ 44 g

$\because$ 100 kg $CaCO_3$ gives CaO = 56 kg

$\therefore$ 190 kg $CaCO_3$ gives CaO $= \frac{56 \times 190}{100} = \mathbf{106.4\ kg}$

Exp–19 Potassium selenate is isomorphous with potassium sulphate and contains 45.42% selenium by weight. Calculate the atomic weight of selenium. Also report the equivalent weight of potassium selenate.

Sol. Potassium selenate is isomorphous to K_2SO_4 and thus its molecular formula is K_2SeO_4.

Now Mol. wt. of $K_2SeO_4 = (39 \times 2 + a + 4 \times 16)$

$= (142 + a)$

Where a is at. wt. of Se.

$(142 + a)$ g K_2SeO_4 has Se = a g

$\therefore$ 100 g K_2SeO_4 has Se = $\frac{a \times 100}{142 + a}$

$\because$ % of Se = 45.42

$\therefore$ $\frac{a \times 100}{142 + a} = 45.42$

$\therefore$ a = **118.168** = **118.2**

Also Eq. wt. of $K_2SeO_4 = \frac{\text{Mol. wt.}}{2} = \frac{2 \times 39 + 118.2 + 64}{2} = \mathbf{130.1}$

Exp–20 One litre of CO_2 is passed over hot coke. The volume becomes 1.4 litre. Find the composition of products, assuming measurements at NTP.

Sol. $CO_2 + C \rightarrow 2CO$

Initial volume	1	0
Final volume	(1 - a)	2a

Given 1 - a + 2a = 1.4

$\therefore$ a = 0.4 litre

$\therefore$ CO_2 = 1 - 0.4 = **0.6 litre**

and CO = 2 x 0.4 = **0.8 litre**

Exp–21 One litre of a mixture of CO and CO_2 is passed through red hot charcoal in tube. The new volume becomes 1.4 litre. Find out % composition of mixture by volume. All measurements are made at same P and T.

Sol. On passing through charcoal only CO_2 reduces to CO.

$CO + C \rightarrow$ No reaction

Volume a

$CO_2 + C \rightarrow 2CO$

Volume before reaction	b	0
Volume after reaction	0	2b

As given a + b = 1 and a + 2b = 1.4

$\therefore$ b = **0.4 litre** $\therefore$ % of b = $\frac{0.4}{1} \times 100 = \mathbf{40\%}$

$\therefore$ a = **0.6 litre** $\therefore$ % of a = $\frac{0.6}{1} \times 100 = \mathbf{60\%}$

Exp–22 Find out equivalent weight of H_3PO_4 in the reaction.

$Ca(OH)_2 + H_3PO_4 \rightarrow CaHPO_4 + 2H_2O$

Sol. The reaction shows two H atoms replaced from H_3PO_4

basicity of H_3PO_4 = 2

Eq. wt. $H_3PO_4 = \frac{M}{2} = \frac{98}{2} = \mathbf{49}$

Exp–23 What volume of 0.20 M H_2SO_4 is required to produce 34.0 g of H_2S by the reaction :

$8KI + 5H_2SO_4 \rightarrow 4K_2HSO_4 + 4I_2 + H_2S + 4H_2O$

Sol. 1 mole of $H_2S \equiv 5$ mole of H_2SO_4

$\therefore$ $\frac{34}{34}$ = 1 mole of $H_2S \equiv 5$ mole of H_2SO_4

$\therefore$ $0.20 \times V = 5$

$\therefore$ $V = \frac{5}{0.20} =$ **25 litre**

Exp–24 The hydrated salt $Na_2SO_4.nH_2O$, undergoes 55.9% loss in weight on heating and becomes anhydrous. The value of n will be

(a) 5 (b) 3 (c) 7 (d) 10

Sol. $Na_2SO_4.nH_2O \xrightarrow{\Delta} Na_2SO_4 + nH_2O$

100 g 44.1 g 55.9 g

44.1 g of Na_2SO_4 has water with it = 55.9 g

$\therefore$ 142 g of Na_2SO_4 has water with it =

$$= \frac{55.9}{44.1} \times 142 = 179.99 \text{ g}$$

Thus, the value of n = $\frac{179.88}{18}$ = **10.0**

Hence the correct answer is (d).

Exp–25 An element (X) having equivalent mass E forms a general oxide X_mO_n, its atomic mass should be

(a) $\frac{2En}{m}$ (b) 2mEn (c) $\frac{E}{n}$ (d) $\frac{ME}{2n}$

Sol. The compound X_mO_n has n x 16 parts of oxygen combining with m x At. mass of X

$\therefore$ 8 parts of oxygen combines with X

$$= \frac{m \times \text{At.mass}}{n \times 16} \times 8$$

i.e.. E = $\frac{m \times \text{At.mass}}{n \times 2}$ or At. mass = $\mathbf{\frac{2E \times n}{m}}$

Hence the correct answer is **(a).**