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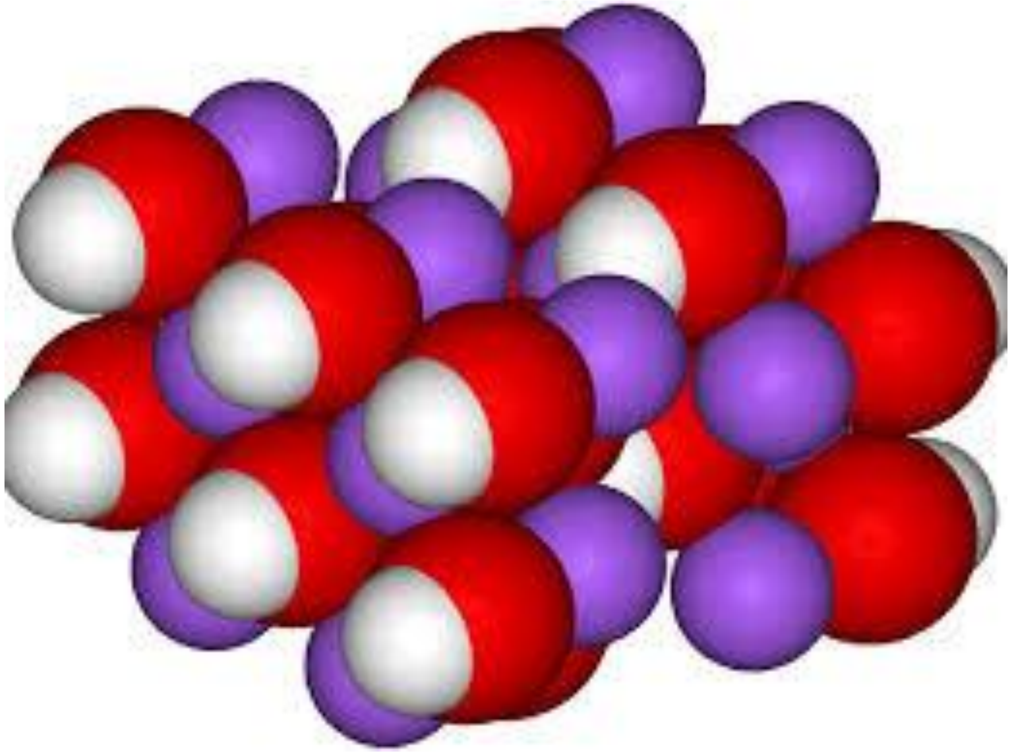
MOLE CONCEPT & STOICHEOMETRY



- Law of chemical combination
- Atomic & molecular mass
- Mole concept
- Determination of molecular formulae
- Stoichiometry
- Concentration terms

1. INTRODUCTION

There are a large number of objects around us which we can see and feel. **Anything that occupies space and has mass is called matter.**

Ancient Indian and Greek philosopher's believed that the wide variety of object around us are made from combination of five basic elements:

Earth, Fire, Water, Air and Sky.

The Indian philosopher Kanad (600 BC) was of the view that matter was composed of very small, indivisible particle called "parmanus".

Ancient Greek philosopher also believed that all matter was composed of tiny building blocks which were hard and indivisible.

The Greek philosopher Domocritus named these building blocks as atoms, meaning indivisible.

All these people had their philosophical views about matter; these views were never put to experimental test.

It was John Dalton who firstly developed a theory on the structure of matter, latter on which was known as **Dalton's atomic theory.**

<u>DALTON'S ATOMIC THEORY</u>:

- Matter is made up of very small indivisible particle called atoms.
- All the atoms of an element are identical in all respect i.e. mass, shape, size, etc. and atoms of different elements are different in nature.
- Atoms cannot be created or destroyed by any chemical process.

Classification of matter

- On the basis of physical behaviour
  - Solids
  - Liquids
  - Gases
- On the basis of chemical behaviour
  - Pure substances
    - Element
    - Compound
  - Mixtures

2. LAWS OF CHEMICAL COMBINATIONS

The combination of elements to form compounds is governed by the following five basic laws.

LAW OF CONSERVATION OF MASS:

It states that matter can neither be created nor destroyed.

This law was put forth by Antoine Lavoisier in 1789.

Lavoisier stated that

"during any physical or chemical change the total mass of the products produced is equal to the total mass of the reactants reacted".

He showed that when mercuric oxide was heated the total mass of mercury and oxygen produced was equal to the total mass of mercuric oxide.

Ex. 1: When 20 g of $NaHCO_3$ is heated, 12.62 g of Na_2CO_3 and 5.24g of CO_2 is produced. How many grams of H_2O is produced?

Solution: Total mass of $NaHCO_3$ heated = 20 gms;

Total mass Na_2CO_3 produced = 12.62 gms

Total mass of CO_2 produced = 5.24 gms

$\therefore$ Mass of H_2O produced = 20–12.62 –5.24 = 2.14 gms

LAW OF DEFINITE PROPORTIONS:

This law was given by, a French chemist, Joseph Proust.

He stated that a given compound always contains exactly the same proportion of elements by weight.

It is sometimes also referred to as **Law of definite composition**.

For example, if water is taken from difference sources, such as rivers, oceans, wells etc. they all contain hydrogen and oxygen, combined in the same proportion by weight in it.

Ex. 2: When 50 g of ammonia is heated it gives 41.18 g of Nitrogen.

When 10 g of Nitrogen is combined with required amount of hydrogen it produces 12.14g ammonia.

Show that the given data follows the law of constant compositions.

Solution:

If 50g of Ammonia gives 41.18g of Nitrogen, then the percentage of Nitrogen in

Ammonia is $\frac{41.18}{50} \times 100$ = 82.36%.

If 10g of Nitrogen gives 12.14 g of Ammonia then percentage of Nitrogen in ammonia is $\frac{10}{12.14} \times 100 = 82.37\%$.

LAW OF MULTIPLE PROPORTIONS:

This law was proposed by Dalton in 1803.

According to this law, **if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers.**

For example, carbon and oxygen combine to form CO and CO_2. In CO, 12 parts by mass of carbon combines with 16 parts by mass of oxygen while in CO_2 12 parts by mass of carbon combines with 32 parts by mass of oxygen. Therefore the ratio of the masses of oxygen that combines with a fixed mass of carbon is 16:32 that is 1:2.

Ex. 3: Sodium and oxygen combine to form two compounds of which one is Na_2O.

The percentage of sodium in the other compound is 59%. Find the formula of this compound.

Solution:

Percentage of sodium in Na_2O is $\frac{2 \times 23}{62} \times 100$ = 74.2% and percentage of oxygen is 25.8%.

Percentage of sodium in other compound is 59% while that of oxygen is 41%. This means that in the first compound (Na_2O) if we take 100 gm then 25.8 gm of oxygen will be present therefore the mass of sodium combining with 1g of oxygen would be $\frac{74.2}{25.8}$ = 2.87 g. Similarly in the second compound the mass of sodium combining with one gm of oxygen is $\frac{59}{41}$ = 1.44g. The ratio of masses of sodium combining with the fixed mass of oxygen is 2.87: 1.44 = 2:1. Therefore formula of the other compound is Na_2O_2.

LAW OF RECIPROCAL PROPORTIONS:

This law which was proposed by Ritcher (1792)

states that **"when two elements combine separately with fixed mass of third element then the ratio of their masses in which they do so is either the same or some whole number multiple of the ratio in which they combine with each other"**.

FOR EXAMPLE:

Carbon, Sulphur and Oxygen form CO_2, SO_2 and CS_2. In CO_2 12 parts by mass of carbon combine with 32 parts by mass of oxygen while in SO_2 32 parts by mass of Sulphur combine with 32 parts by mass of oxygen. Ratio of masses of carbon and sulphur which combine with fixed mass of oxygen is 12:32 or 3:8.In CS_2 12 parts by mass of carbon combines with 64 parts by mass of sulphur therefore the ratio of mass of carbon to sulphur in carbon disulphide is 12:64 i.e. 3:16.

Therefore, the ratio is $\frac{3}{8}$: $\frac{3}{16}$ or 2:1

GAY - LUSSAC LAW OF COMBINING VOLUMES:

This law which was proposed by Gay – Lussac states that,

the volumes of gaseous reactants reacted and the volumes of gaseous products formed, all measured at the same temperature and pressure bear a simple ratio.

FOR EXAMPLE:

The reaction involved in Haber's Process (Nitrogen and hydrogen gases react to form ammonia)

$$\begin{array}{lll} N_2(g) + & 3H_2(g) \rightarrow & 2NH_3(g) \\ 1\,\text{vol} & 3\,\text{vol} & 2\,\text{vol} \end{array}$$

It is observed that the ratio of the volumes of N_2 and H_2 reacted and volume of NH_3 produced is equal to **1:3:2** which is a simple ratio.

❖ **This law is applicable only for gaseous reactions and should not be used for non–gaseous reactants and products.**

LAWS OF CHEMICAL COMBINATIONS

1. Potassium combines with two isotopes of chlorine (^{35}Cl and ^{37}Cl) respectively to form two samples of KCl. Their formation follows the law of:
(A) constant proportions (B) multiple proportions
(C) reciprocal proportions (D) none of these

2. H_2S contains 5.88% hydrogen, H_2O contains 11.11% hydrogen while SO_2 contains 50% sulphur. These figures illustrate the law of:
(A)conservation of mass (B) constant proportions
(C)multiple proportions (D) reciprocal proportions

3. Irrespective of the source, pure sample of water always yields 88.89% mass of oxygen and 11.11% mass of hydrogen. This is explained by the law of:
(A)conservation of mass (B) constant composition
(C)multiple proportion (D) constant volume

ANSWER KEY

1. (D) 2. (D) 3. (B)

3. ATOMIC MASS & MOLECULAR MASS

In 1961 International Union of Pure and Applied Chemists (IUPAC) selected the most stable isotope of carbon, C–12 as the standard for comparison of atomic masses of elements. The mass of C–12 atom is taken as 12 atomic mass unit.

The scale in which the relative atomic masses of different elements are expressed is called atomic mass unit or amu.

Atomic mass Unit (amu) = $\frac{1}{12}$ the mass of a C - 12 atom = 1.660539 x10^{-24} gm

$$\text{Atomic mass of an element} = \frac{\text{Mass of one atom of the element}}{\frac{1}{12} \times \text{Mass of one atom of } {}^{12}_{6}C}$$

One amu is also called one Dalton (Da).
Nowadays amu has been replaced by ‘u’ which is known as unified mass.

MOLECULAR MASS:

Molecular Mass is the sum of atomic masses of the elements present in a molecule.

It is obtained by multiplying the atomic mass of each element by the number of its atoms and adding them together.

For example,

molecular mass of methane which contains one carbon atom and four hydrogen atoms can be obtained as follows:

Molecular mass of methane, CH_4 = (12.011) + 4 (1.008) = 16.043 u

Similarly, molecular mass of H_2O = 2 × atomic mass of hydrogen + 1 × atomic mass of oxygen

= 2 (1.008) + 16.00 = 18.016 u

4. MOLE CONCEPT

One mole is an amount of substance containing Avogadro's number of particles. Avogadro's number is equal to 602,214,199,000,000,000,000,000
or more simply, $6.02214199 \times 10^{23}$.

A mole (symbol mol) is defined as the amount of substance that contains as many atoms, molecules, ions, electrons or any other elementary entities as there are carbon atoms in exactly 12 gm of ^{12}C.
The number of atoms in 12 gm of ^{12}C is called Avogadro's number (N_A).

$$\boxed{N_A = 6.022 \times 10^{23}}$$

From mass spectrometer we found that there are 6.022×10^{23} atoms present in 12 gm of C – 12 isotope.

GRAM ATOMIC MASS:

- The atomic mass of an element expressed in grams is called gram atomic mass of that element.
- It is also defined as mass of 6.022×10^{23} atoms.
- It is also defined as the mass of one mole atoms.
- It is also defined as the mass of 1 gram atom of the element.

For example for oxygen atom:

Atomic mass of 'O' atom = mass of one 'O' atom = 16 amu

Gram atomic mass = mass of 6.022×10^{23} 'O' atoms

$= 16 \text{ amu} \times 6.022 \times 10^{23}$

$= 16 \times 1.66 \times 10^{-24} \times 6.022 \times 10^{23} \simeq 16$ gm/mole

$$\because 1.66 \times 10^{-24} \times 6.022 \times 10^{23} = 1$$

Ex. 4: How many atoms of oxygen are there in 16 g oxygen?

Sol. $x \times 1.66 \times 10^{-24} \times 16 = 16g$ $\qquad x = \frac{1}{1.66 \times 10^{-24}} = N_A$

<u>GRAM MOLECULAR MASS</u> :

- The molecular mass of a substance expressed in gram is called the gram-molecular mass of the substance.
- It is also defined as mass of 6.022×10^{23} molecules .r
- It is also defined as the mass of 1 mole molecules.
- It is also defined as the mass of 1 gram molecule.

For examples for 'O_2' molecules :

Molecular mass of 'O_2' molecule = mass of one 'O_2' molecule

= $2 \times$ mass of one 'O' atom

= 2×16 amu

= 32 amu

Gram molecular mass 'O$_2$' molecules $= \text{mass of } 6.022 \times 10^{23}$

$= 32 \text{ amu} \times 6.022 \times 10^{23}$

$= 32 \times 1.66 \times 10^{-24} \text{ gm} \times 6.022 \times 10^{23}$

$= 32 \text{ gm/mole}$

Ex. 5: The molecular mass of H_2SO_4 is 98 amu. Calculate the number of moles of each element in 294 g of H_2SO_4.

Sol. Gram molecular mass of H_2SO_4 = 98 gm

Moles of H_2SO_4 = $\frac{294}{98} = 3\,moles$

H_2SO_4	H	S	O
one molecule	2 atoms	one atom	4 atoms
1 x N_A	2 x N_A atoms	1 x N_A atoms	4 x N_A atoms
$\therefore$ one mole	2 mole	one mole	4 mole
$\therefore$ 3 mole	6 mole	3 mole	12 mole

AVOGADRO'S HYPOTHESIS:

Equal volumes of gases have equal number of molecules (not atoms) at same temperature and pressure condition.

S.T.P. (Standard Temperature and Pressure)

Temperature = $0^{o}C$ or 273 K

Pressure = 1 atm = 760 mm of Hg

Volume of one mole of gas at STP is found to be experimentally equal to 22.4 litres which is known as molar volume.

Ex. 6: Calculate the volume in litre of 20 g hydrogen gas at STP.

Sol. No. of moles of hydrogen gas = $\frac{\text{Given mass}}{\text{Molecular mass}} = \frac{20\,gm}{2\,gm} = 10\,mol$

Volume of hydrogen gas at STP = 10 x 22.4 lt = 224 lt.

FORMULA SHEET FOR MOLE CALCULATIONS:

S. No.	Items	Formula
1.	Molecules	$\frac{\text{Wt. in gm}}{\text{Molecular mass}}$
2.	Atom	$\frac{\text{Wt. in gm}}{\text{Atomic mass}}$
3.	Gases	$\frac{\text{Volume at STP}}{\text{Standard molar Volume at STP}}$
4.	Any Particle	$\frac{\text{Number of particles}}{\text{Avogardro number}}$
5.	Moles of A in A_xB_y	x
6.	For Gases at any P(in atm), V(in litres) & T(in K)	$n = \frac{PV}{RT}$ where R = 0.0821 lit-atm/mol K

IMPORTANT NOTES:

1. **STP or NTP conditions :** Standard conditions means that temperature is 0°C or 273K and pressure is one atmosphere or 760mm of Hg.
2. **1 gm - atom** is same as 1 mole of an atom & hence will have wt equal to atomic wt expressed in gms.
3. **1 gm – molecule** is same as 1 mole of the molecule & hence will have wt equal to molecular wt expressed in gms.
4. **1 gm – Ion** is same as 1 mole of an ion & hence will have wt equal to ionic wt

❖ Remember 1 gm of atom & 1gm– atom are two different phrases. Former is mentioning wt (equal to 1gm) & latter is mentioning moles.

e.g. (1) " x g atom of nitrogen " = x moles of N atom = $(x \times N_A)$ number of N atoms

(2) " x g molecule of nitrogen" = x moles of N_2 molecules = $(x \times N_A)$ molecules of N_2

= $(2x \times N_A)$ number of N atom

Ex. 7: **How many g atom and no. of atoms are there in (a) 60 g carbon (b) 224.4 g Cu?**

Given : **At. Weight of C and Cu are 12 and 63.6 respectively. Avogadro's no. = 6.02 x 10^{23}.**

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

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

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

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

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

Ex. 8: Find the number of g atoms and weight of an element having 2×10^{23} atoms. At. Weight of element is 32.

Solution $\because$ N_A atoms have 1 g atom

2×10^{23} atoms have $= \dfrac{2\times10^{23}}{6.022\times10^{23}} = 0.33g\ atom$

$\because$ N_A atoms of elements weigh 32 g

$\therefore$ 2×10^{23} **atoms of element weigh** $= \dfrac{2\times10^{23}\times32}{6.022\times10^{23}} = 10.628g$

Ex. 9: How many mole and molecules of O_2 are there in 64g O_2? What is the mass of one molecule of O_2?

Solution: $\because$ moles of O_2 in 32 g O_2 = 1

$\therefore$ In 64 g O_2 moles $= \frac{64 \times 1}{32} = 2\, mole$

$\because$ 32 g O_2 contains 6.022 x 10^{23} molecules

$\therefore$ 64 g O_2 contains $= \frac{6.022 \times 10^{23} \times 64}{32} = 12.044 \times 10^{23}$ molecules

$\therefore$ N_A molecules of O_2 weigh 32 gm

$\therefore$ 1 molecule of O_2 weighs $= \frac{32}{6.022 \times 10^{23}} = 5.3138 \times 10^{-23}\, gm$.

Ex. 10: From 200 mg of CO_2 , 10^{21} molecules are removed. How many g and mole of CO_2 are left?

Solution : $\because$ 6.022×10^{23} molecules of CO_2 = 44 g

$\therefore$ 10^{21} molecules of CO_2 $= \frac{44 \times 10^{21}}{6.022 \times 10^{23}} g$

$= 7.31 \times 10^{-2}$

= 73. 1 mg

$\therefore$ CO_2 left = 200 – 73.1 = 126.9 mg

Also Mole of CO_2 left $= \frac{wt.}{M.wt} = \frac{126.9 \times 10^{-3}}{44} = 2.88 \times 10^{-3}$

5. AVERAGE WEIGHT

AVERAGE ATOMIC WEIGHT:

For elements, which have atoms with different relative masses (isotopes) the atomic mass is taken as weighted mean of the atomic masses.

Average atomic weight = ∑ % of Isotope X Atomic Wt. of Isotope

For example, chlorine contains two isotopes of atomic masses 35 amu and 37 amu. The relative abundance of these two is in the ratio of 3:1.

Thus the atomic mass of chlorine is the average of different relative masses.

Therefore atomic mass of chlorine is equal to $\frac{35 \times 3 + 37 \times 1}{4}$ = 35.5 amu.

Ex. 11: Calculate average atomic wt. of silicon if relative abundance is 92.23% Si^{28}, 4.77%Si^{29}, 3% Si^{30}

Solution : Av at wt $= = \frac{92.23 \times 28 + 4.77 \times 29 + 3 \times 30}{100}$

= 28.1 amu

Ex. 12: Calculate % abundance of Ag^{109} if it is known that silver exist in only two isotopes Ag^{107} & Ag^{109} & average atomic weight = 108.5

Solution : let the % abundance of Ag^{109} be x

$$\therefore 108.5 = \frac{x \times 109 + (100 - x) \times 107}{100}$$

$$\Rightarrow 10850 - 10700 = 2x \quad \Rightarrow x = 75\%$$

Shortcut to calculate % abundance when an element X is having only two isotopes X^A & X^B & average atomic wt is X_{avg}

$$\Rightarrow \%\ of\ X^A = \frac{\left|X_{avg} - wt\ of\ X^B\right|}{\left|wt\ of\ X^A - wt\ of\ X^b\right|} \times 100$$

% obtained above is mole %.

AVERAGE MOLECULAR WEIGHT:

For homogenous mixture of several substances having number of moles, n_i & molecular mass $M_{(i)}$ for i^{th} species, the average molecular weight is given as

Average Molecular wt. = $\dfrac{n_1 \times M_1 + n_2 \times M_2 + ...}{n_1 + n_2 + ...}$

$$= \frac{Total\ weight}{Total\ no.of\ moles.} \qquad = \frac{\sum n_i M_i}{\sum n_i}$$

Ex. 13: Dry air has a molar composition as 20% O_2 & 80% N_2. Calculate average molecular wt. of dry air.

Solution :. Av. Molecular wt $= \dfrac{20 \times 32 + 80 \times 28}{100} = 28.8$

MINIMUM MOLECULAR WEIGHT:

It is the molecular weight of a compound shown by presence of minimum number of atoms

[i.e. for monomer = 1 , Dimer = 2 , Trimer = 3 , Tetramer = 4]

e.g. Insulin contains 3.4% S, Find its MMW.

100g insulin contains 3.4g S (sulphur). Assuming it to be monomer [as nothing specified] one sulphur atom $\approx$32 amu (atomic mass), We can write, 3.4 g S is contained by 100g insulin

32g S is contained by $\left(\frac{100}{3.4}\times 32\right)$ = MMW

MOLE CONCEPT

4. The best standard of atomic mass is:
(A)carbon-12 (B) oxygen-16
(C)hydrogen-1.008 (D) chlorine-35.5

5. The chemical formula of a particular compound represents:
(A)the size of its molecule
(B)the shape of its molecule
(C)the total number of atoms in a molecule
(D)the number of different types of atoms in a molecule

6. Two containers P and Q of equal volume (1 litre each) contain O_2 and SO_2 respectively at 300 K and 1 atmosphere. Then
(A)Number of molecules in P is less than that in Q
(B)Number of molecules in Q is less than that in P
(C)Number of molecules in P and Q are same
(D) Either (A) or (B)

7. 7.5 grams of a gas occupy 5.6 litres of volume at STP. The gas is
(A) NO (B) N_2O (C) CO (D) CO_2

8. The weight of a molecule of the compound $C_{60}H_{122}$ is
(A) 1.4×10^{-21} g (B) 1.09×10^{-21} g (C) 5.025×10^{23} g (D) 16.023×10^{23} g

9. 1.0 mole of CO_2 contains:
(A) 6.02×10^{23} atoms of C (B) 6.02×10^{23} atoms of O
(C) 18.1×10^{23} molecules of CO_2 (D) 3 g-atoms of CO_2

10. The number of atoms in 1.4 g nitrogen gas is:
(A) 6.02×10^{22} (B) 3.01×10^{22} (C) 1.20×10^{23} (D) 6.02×10^{23}

11. Which of the following has the smallest number of molecules?

(A) 22.4×10^3 ml of CO_2 gas (B) 22 g of CO_2 gas
(C) 11.2 litre of CO_2 gas (D) 0.1 mole of CO_2 gas

12. The number of grams of H_2SO_4 present in 0.25 mole of H_2SO_4 is
(A) 0.245 (B) 2.45 (C) 24.5 (D) 49.0

13. At NTP 1.0 g hydrogen has volume in litre:
(A) 1.12 (B) 22.4 (C) 2.24 (D) 11.2

14. 19.7 kg of gold was recovered from a smuggler. The atoms of gold recovered are: (Au = 197)
(A) 10 (B) 6.02×10^{23}
(C) 6.02×10^{24} (D) 6.02×10^{25}

15. The molecular mass of CO_2 is 44 amu and Avogadro's number is 6.02×10^{23}. Therefore, the mass of one molecule of CO_2 is:

(A) 7.31×10^{-23} (B) 3.65×10^{-23}
(C) 1.01×10^{-23} (D) 2.01×10^{-23}

16. The number of moles of H_2 in 0.224 litre of hydrogen gas at NTP is:

(A) 1 (B) 0.1 (C) 0.01 (D) 0.001

17. Choose the wrong statement:-

(A) 1 Mole means 6.02×10^{23} particles
(B) Molar mass is mass of one molecule
(C) Molar mass is mass of one mole of a substance
(D) Molar mass is molecular mass expressed in grams

18. 3 mol of ammonia contains:

(A) 18 gm of hydrogen (B) 42 gm of nitrogen
(C) both (D) None

19. Total no. of protons in 36 ml of water at 4°C (where ρ of water = 1 g/ml) is

(A) 20 (B) 16 (C) 20 N_A (D) $16N_A$

20. In which of the following pairs both members have same no. of atoms

(A) 1 gm O_2, 1 gm O_3 (B) 1 gm N_2, 2 gm N
(C) Both (D) None

21. The molecular wt. of green vitriol is M_0. The wt. of 10^{-3} N_A molecules of it is

(A) M_0 gm (B) M_0 mg (C) 10^3 M_0 gm (D) 10^{-3} M_0 mg

22. Which of the following has the highest mass?

(A) 1g atom of C (B) ½ mole of CH_4
(C) 10ml of H_2O (D) 3.011×10^{23} atom of oxygen.

23. Which one of the following samples contains the largest number of atoms?

(A) 2.5 mole CH_4 (B) 10 mole He (C) 4 mole SO_2 (D) 1.8 mole S_8

24. The no. of atoms in 52 amu. of He is

(A) 13×10^{23} (B) 1.3×10^{23} (C) 13 (D) 103

25. The no. of electrons in 2 gm ion of nitrate ion (NO_3^-) is

(A) 64 (B) $64N_A$ (C) 32 (D) $32N_A$

26. The mass of carbon present in 0.5 moles of $K_4[Fe(CN)_6]$ is

(A) 1.8 g (B) 18 g (C) 3.6 g (D) 36 g

27. The largest no. of molecules is in

(A) 28 g of CO_2 (B) 46 g of C_2H_5OH

(C) 36 g of H_2O (D) 54 g of N_2O_5

28. How many electrons are present in 180 gm. of water?

(A) 1 mole (B) 10 moles (C) 18 moles (D) 100 moles

29. How many molecules of H_2O are contained in 2.48 g of $Na_2S_2O_3.5H_2O$ (at.wt. of Na=23,S=32)

(A) 3×10^{20} (B) 3×10^{21} (C) 3×10^{22} (D) 3×10^{23}

30. The no. of silver atoms present in a 90% pure silver wire weighing 10 g. is (at.wt. of Ag=108)

(A) $8x10^{22}$ (B) $0.62x10^{23}$ (C) $5x10^{22}$ (D) $6.2x10^{29}$

31. The number of molecules of water in 333 g of $Al_2(SO_4)_3.18H_2O$ is

(A) $18\times6.02\times10^{23}$ (B) $9\times6.02\times10^{23}$ (C) 18 (D) 36

32. The number of water molecules present in a drop of water weighing 0.018 g is

(A) 6.02×10^{26} (B) 6.02×10^{23} (C) 6.02×10^{20} (D) 6.02×10^{19}

33. If N_A is Avogadro's number, then the number of valence electrons in 4.2 g of nitride ion (N^{3-}) is (Given One atom of N has 5 valence electrons)

(A) 2.4 N_A (B) 4.2 N_A (C) 1.6 N_A (D) 3.2 N_A

34. A person adds 3.42 of sucrose ($C_{12}H_{22}O_{11}$) in his cup of tea to sweeten it. How many atoms of carbon does he add?

(A) 132.44×10^{21} atoms (B) 66.22×10^{21} atoms

(C) 0.1 atoms (D) 72.27×10^{21} atoms

35. The total number of protons in 8.4 g of $MgCO_3$ is ($N_A = 6.02\times10^{23}$) :

(A) 2.52×10^{22} (B) 2.52×10^{24} (C) 3.01×10^{24} (D) 3.01×10^{22}

36. $4.4\,g$ of CO_2 and $2.24\,litre$ of H_2 at STP are mixed in a container. The total number of molecules present in the container will be

(A) 6.022×10^{23} (B) 1.2044×10^{23} (C) $2\,mole$ (D) 6.023×10^{24}

37. Which sample contains the largest number of atoms:

(A) $1mg$ of C_4H_{10} (B) $1mg$ of N_2 (C) $1mg$ of Na (D) $1mL$ of water

38. The atomic weight of a triatomic gas is a. The correct formula for the number of moles of gas in its wg is:

(A) $\frac{3w}{a}$ (B) $\frac{w}{3a}$ (C) $3wa$ (D) $\frac{a}{3w}$

39. Number of atoms in $558.5\,g\,Fe\,(at.wt\,55.85)$ is:

(A) Twice that in $60\,g$ carbon (B) 6.023×10^{22}

(C) Half in $8g$ He (D) $558.5\times6.023\times10^{23}$

40. How many moles of magnesium phosphate, $Mg_3(PO_4)_2$ will contain $0.25\,\text{mole}$ of oxygen atoms?

(A) 0.02 (B) 3.125×10^{-2} (C) 1.25×10^{-2} (D) 2.5×10^{-2}

ANSWER KEY

4. (A)	5. (D)	6. (C)	7. (A)	8. (A)	9. (A)
10. (A)	11. (D)	12. (C)	13. (D)	14. (D)	15. (A)
16. (C)	17. (B)	18. (B)	19. (C)	20. (A)	21. (B)
22. (A)	23. (D)	24. (C)	25. (B)	26. (D)	27. (C)
28. (D)	29. (C)	30. (C)	31. (B)	32. (C)	33. (A)
34. (D)	35. (B)	36. (B)	37. (D)	38. (B)	39. (A)
40. (B)					

6. DETERMINATION OF MOLECULAR & EMPERICAL FORMULAE

The molecular formula of a compound may be defined as the formula which gives the actual number of atoms of various elements present in the molecule of the compound.

For example, the molecular formula of the compound glucose can be represented as $C_6H_{12}O_6$. A molecule of glucose contains six atoms of carbon, twelve atoms of hydrogen and six atoms of oxygen.

In order to find out molecular formula of a compound, the first step is to determine its empirical formula from the percentage composition.

$$\boxed{\text{Mass \% of an element} = \frac{\text{Mass of that element}}{\text{Molar mass}} \times 100}$$

EMPIRICAL FORMULA:

The empirical formula of a compound may be defined as the formula which gives the simplest whole number ratio of atoms of the various elements present in the molecule of the compound. The empirical formula of the compound glucose ($C_6H_{12}O_6$) is CH_2O which shows that C,H and O are present in the simplest ratio of 1:2:1

Empirical formula mass of substance is equal to the sum of atomic masses of all the atoms in the empirical formula of the substance. Molecular formula is a whole number multiple of empirical formula.

Thus Molecular formula = (Empirical formula) $\times$ n where n = 1,2,3…

$$n = \frac{\text{Molecular Formula}}{\text{Empirical Formula}} = \frac{\text{Molecular Mass}}{\text{Empirical Mass}}$$

STEPS FOR WRITING THE EMPIRICAL FORMULA:

The percentage of the elements in the compound is determined by suitable methods and from the data collected; the empirical formula is determined by the following steps:

- Divide the percentage of each element by its atomic mass. This will give the relative number of moles of various elements present in the compound.
- Divide the quotients obtained in the above step by the smallest of them so as to get a simple ratio of moles of various elements.
- Multiply the figures, so obtained by a suitable integer if necessary in order to obtain a whole number ratio.
- Finally write down the symbols of the various elements side by side and put the above number as the subscripts to the lower right hand corner of each symbol. This will represent the empirical formula of the compound.

STEPS FOR WRITING THE MOLECULAR FORMULA:

- Calculate the empirical formula as described above.
- Find out the empirical formula mass by adding the atomic masses of all the atoms present in the empirical formula of the compound.
- Divide the molecular mass (determined experimentally by some suitable method) by the empirical formula mass and find out the value of n.

DENSITY:

Density is of two types, Absolute Density and Relative Density

For liquid and solids:

Absolute density = $\dfrac{mass}{volume}$

Relative density or specific gravity = $\dfrac{\text{density of the substance}}{\text{density of water at } 4^{o}\,\text{C}}$

For Gases :

Absolute density (mass / volume) = $\dfrac{Molar\ mass}{Molar\ volume} = \dfrac{PM}{RT}$

Where P is pressure of gas, M = mol. Wt. of gas , R is the gas constant , T is the temperature.

VAPOUR DENSITY:

Vapour density is defined as the density of the gas with respect to hydrogen gas at the same temperature and pressure.

Vapour density = $\frac{d_{gas}}{d_{H_2}} = \frac{\frac{PMgas}{RT}}{\frac{PM_{H_2}}{RT}}$ => V. D. = $\frac{M_{gas}}{M_{H_2}} = \frac{M_{gas}}{2}$

$M_{gas} = 2\ (V.D.)$

Relative density can be calculated w.r.t. to other gases also.

Ex. 14 : What is the V.D. of SO_2 with respect to CH_4

Solution: **V.D. =** $\frac{(MW)_{SO_2}}{(MW)_{CH_4}}$ **V.D. =** $\frac{64}{16} = 4$

Ex. 15 : 11.2 litre of the particular gas at S.T.P. weighs 16 gram. What is the V. D. of gas.

Solution: wt.of 11.2 litre = 16 gram.

Moles = $\frac{11.2}{22.4} = \frac{16}{M}$ $M = 32\,gm/mole$ **V.D.=** $\frac{32}{2} = 16$

Ex. 16: A substance, on analysis, gave the following percentage composition: Na = 43. 4%,

C = 11.3%, O = 45.3%. Calculate its empirical formula. {Na = 23, C = 12, O = 16]

Solution:

Element	SYMBOL	% age	Atomic Mass	Relative number of moles	Simple ratio of moles	Simplest whole no. ratio
Sodium	Na	43.4	23	$\frac{43.4}{23}=1.88$	$\frac{1.88}{0.94}=2$	2
Carbon	C	11.3	12	$\frac{11.3}{12}=0.94$	$\frac{0.94}{0.94}=1$	1
Oxygen	O	45.3	16	$\frac{45.3}{16}=2.83$	$\frac{2.83}{0.94}=3$	3

Therefore, the empirical formula is Na_2CO_3.

Ex. 17: A compound has the following composition: Mg = 9.76%, S = 13.01%, O = 26.01%,
H_2O = 51.22%. What is its empirical formula? [Mg = 24, S = 32, O = 16, H = 1]

Solution:

Element	Symbol	% age	Atomic Mass	Relative number of moles	Simple ratioof moles	Simplest Wholeno. ratio
Magnesium	Mg	9.76	24	$\frac{9.76}{24} = 0.406$	$\frac{0.406}{0.406} = 1$	1
Sulphur	S	13.01	32	$\frac{13.01}{32} = 0.406$	$\frac{0.406}{0.406} = 1$	1
Oxygen	O	26.01	16	$\frac{26.01}{16} = 1.625$	$\frac{1.625}{0.406} = 4$	4
Water	H_2O	51.22	18	$\frac{51.22}{18} = 2.846$	$\frac{2.846}{0.406} = 7$	7

Hence, the empirical formula is $MgSO_4 . 7H_2O$.

Ex. 18: What is the simplest formula of the compound which has the following percentage composition. Carbon 80%, Hydrogen 20%. If the molecular mass is 30, calculate its molecular formula.

Solution:

Element	% age	At. Mass	Relative number of moles	Simple ratio of moles	Simples whole no.ratio
C	80	12	$\frac{80}{12} = 6.66$	$\frac{6.66}{6.66} = 1$	1
H	20	1	$\frac{20}{1} = 20$	$\frac{20}{6.66} = 3$	3

$\therefore$ Empirical formula is CH_3

$\therefore$ Empirical mass = $12 \times 1 + 1 \times 3 = 15$ $n = \frac{\text{Molecular mass}}{\text{Emprical formula mass}} = \frac{30}{15} = 2$

Molecular formula = Empirical formula $\times$ 2 = $CH_3 \times 2 = C_2H_6$

Ex. 19: A compound on analysis gave the following percentage composition: C = 54.54%, H = 9.09%, O = 36.36%. The vapour density of the compound was found to be 44. Find out the molecular formula of the compound.

Solution: Calculation of empirical formula.

Element	% age	At. Mass	Relative number of moles	Simple ratio of mole	Simplest whole no. ratio
C	54.54	12	$\frac{54.54}{12} = 4.53$	$\frac{4.53}{2.27} = 2$	2
H	9.09	1	$\frac{9.09}{1} = 9.09$	$\frac{9.09}{2.27} = 4$	4
O	36.36	16	$\frac{36.36}{16} = 2.27$	$\frac{2.27}{2.27} = 1$	1

$\therefore$ Empirical formula is C_2H_4O. Calculation of molecular formula:

Empirical formula mass = $12 \times 2 + 1 \times 4 + 16 \times 1 = 44$

Molecular mass = $2 \times$ Vapour density = $2 \times 44 = 88$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{88}{44} = 2$$

Molecular formula = Empirical formula $\times$ n = $C_2H_4O \times 2 = C_4H_8O_2$.

Ex. 20: An organic compound on analysis gave the following data: C = 57.82%, H = 3.6%, and the rest is oxygen. Its vapour density is 83. Find its empirical and molecular formula.

Solution: Calculation of empirical formula:

Element	% age	At. Mass	Relative number of moles	Simple ratio of moles	Simplest whole no. ratio
C	57.82	12	$\frac{57.82}{12} = 4.80$	$\frac{4.8}{2.4} = 2$	4
H	3.60	1	$\frac{3.60}{1} = 3.60$	$\frac{3.6}{2.3} = 1.5$	3
O	38.58	16	$\frac{38.58}{16} = 2.40$	$\frac{2.4}{2.4} = 1$	2

$\therefore$ Empirical formula is $C_4H_3O_2$. Empirical formula mass = $12\times4 + 1\times3 + 2\times16 = 83$

Molecular mass = $2 \times$ V.D. = $2 \times 83 = 166$

$$n = \frac{\text{Molecular mass}}{\text{Emprical formula mass}} = \frac{166}{83} = 2$$

Molecular formula = Empirical formula $\times$ n = $C_4H_3O_2 \times 2 = C_8H_6O_4$

Ex. 21: 2.746 gm of a compound gave on analysis 1.94 gm of silver, 0.268 gm of sulphur and 0.538 gm of oxygen. Find the empirical formula of the compound. (At masses: Ag = 108, S = 32, O = 16)

Solution: To calculate percentage composition.

Percentage composition of the compound is calculated as under:

$$\text{Silver} = \frac{1.94}{2.746}\times 100 = 70.65\%$$

Sulphur = $\frac{0.268}{2.746}\times 100 = 9.75\%$ Oxygen = $\frac{0.538}{2.746}\times 100$ = 19.6%

To calculate empirical formula:

Element	% age	At. Mass	Relative number of moles	Simplest ratio of moles	Simplest whole no.ratio
Ag	70.65	108	$\frac{70.65}{108} = 0.654$	$\frac{0.654}{0.305} = 2$	2
S	9.75	32	$\frac{9.75}{32} = 0.305$	$\frac{0.305}{0.305} = 1$	1
O	19.6	16	$\frac{19.6}{16} = 1.22$	$\frac{1.22}{0.305} = 4$	4

$\therefore$ Empirical formula is Ag_2SO_4

EMPIRICAL AND MOLECULAR FORMULA

41. The simplest formula of a compound containing 50% by mass of element X (at. wt. 10) and 50% by mass of element Y (at. wt. 20) is ;

(A) XY (B) X_2Y (C) XY_2 (D) X_2Y_3

42. The hydrated salt $Na_2SO_4 \cdot 10H_2O$ undergoes X% loss in weight on heating and becomes anhydrous. The value of X will be

(A) 10 (B) 45 (C) 56 (D) 70

43. An oxide of iodine (I =127) contains 25.4 g of iodine and 8 g of oxygen. Its formula could be

(A) I_2O_3 (B) I_2O (C) I_2O_5 (D) I_2O_7

44. The chloride of a metal contains 71% chlorine by weight and the vapour density of it is 50. The atomic weight of the metal will be

(A) 29 (B) 58 (C) 35.5 (D) 71

ANSWER KEY

41. (B)   42. (C)   43. (C)   44. (A)

7. STOICHIOMETRY

The word 'Stoichiometry' is derived from two Greek words - Stoicheion (meaning element) and metron (meaning measure

- Write the balanced chemical equation.
- Write the atomic mass/molecular mass/moles/molar volumes of the species involved in calculations.
- Calculate the result by applying unitary method.

A balanced equation for this reaction is as given below:

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

The above balance reaction gives the following information:

- For every 1 mole of CH_4, 2 mole of O_2 will be required to produce 1 mole of CO_2 and 2 moles of H_2O. this signifies **Mole – Mole relation**

- For every 16 gms of CH_4 , 64 gms of O_2 will be required to produce 44gms of CO_2 and 36 gms of H_2O this signifies **Mass – Mass relation**

- Ratio of moles of CO_2 : H_2O at any time = 1 : 2

- There will be no change in total mass of all reactants and products at any time for any chemical reaction.

- For the above reaction only, there will be no change in total number of moles of all reactants and products.

7.1 INTERPRETATION OF BALANCED CHEMICAL EQUATIONS

Once we get a balanced chemical equation then we can interpret a chemical equation by following ways

1. Mass – mass analysis
2. Mass – volume analysis
3. Volume – volume analysis

MASS – MASS ANALYSIS:

Consider the reaction, $2KClO_3 \rightarrow 2KCl + 3O_2$

According to stoichiometry of the reaction

Mass – mass ratio: 2 x 122.5 : 2 x 74.5 : 3 x 32

Or $$\frac{Mass\ of\ KClO_3}{Mass\ of\ KCl} = \frac{2\times122.5}{2\times74.5}$$

$$\frac{Mass\ of\ KClO_3}{Mass\ of\ O_2} = \frac{2\times122.5}{3\times32}$$

Ex. 22: 367.5 gram $KClO_3$ (M = 122.5) is heated.

How many gram KCl and oxygen is produced.

Solution: Balanced chemical equation for heating of $KClO_3$ is

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

Mass – mass ratio: 2 x 122.5 gm 2 x 74.5 gm : 3 x 32 gm

$$\frac{Mass\ of\ KClO_3}{Mass\ of\ KCl} = \frac{2\times122.5}{2\times74.5} \quad \Rightarrow \quad \frac{367.5}{W} = \frac{122.5}{74.5}$$

W_{KCl} = 3 x 74.5 = 223.5 gm

$$\frac{Mass\ of\ KClO_3}{Mass\ of\ O_2} = \frac{2\times122.5}{3\times32} \quad \Rightarrow \quad \frac{367.5}{W} = \frac{2\times122.5}{3\times32}$$

W_{oxygen} = 144 gm

Ex. 23: How many grams of oxygen (O_2) are required to completely react with 0.200 g of hydrogen (H_2) to yield water (H_2O)? Also calculate the amount of water formed. (At. Mass H = 1; O = 16).

Solution: The balanced equation for the reaction is

$2H_2$	+	O_2	$\longrightarrow$	$2H_2O$
2 mol		1 mol		2 mol
4 g		32 g		36 g

Now, 4g of H_2 require oxygen = 32 g

0.200 g of H_2 require oxygen = $\frac{32}{4} \times 0.200 = 1.6$ g

Again, 4g of H_2 produce H_2O = 36 g

0.200 g of H_2 produce H_2O = $\frac{36}{4} \times 0.200 = 1.8$ g.

MASS – VOLUME ANALYSIS :

Now again consider decomposition of $KClO_3$

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

Mass Volume ratio, 2 x 122.5 gm : 2 x 74.5 gm : 3 x 22.4 lt at S.T.P.

we can use two relation for volume of oxygen.

$$\frac{Mass\ of\ KClO_3}{volume\ of\ O_2\ at\ STP} = \frac{2\times 122.5}{3\times 22.4\,lt} \quad \ldots(i)$$

And $$\frac{Mass\ of\ KCl}{volume\ of\ O_2\ at\ STP} = \frac{2\times 74.5}{3\times 22.4\,lt} \quad \ldots(ii)$$

Ex. 24: Calculate the volume of O_2 and volume of air needed for combustion of 1 kg carbon at STP.(Assume air contains 20% oxygen)

Solution: $C + O_2 \longrightarrow CO_2$

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

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

= 1866.67 litre O_2

$\therefore$ $V_{air} = 5 \times V_{O_2} \; = 5 \times 1866.67 = 9333.35$ litre

Ex. 25: What volume of oxygen at N.T.P. can be produced by 6.125 g of potassium chlorate according to the reaction $2KClO_3 \longrightarrow 2KCl + 3O_2$.

Solution: The given chemical equation is :

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

2 mol 2 mol 3 mol

2×122.5 g 3×22.4L at N.T.P

Now 245 g of $KClO_3$ produce oxygen at N.T.P. = 3×22.4 L

6.125 g of $KClO_3$ produce oxygen = $\frac{3 \times 22.4}{245} \times 6.125 = 1.68$ L at N.T.P.

VOLUME – VOLUME RELATIONSHIP:

It relates the volume of gaseous species (reactants or product) with the volume of another gaseous species (reactant or product) involved in a chemical reaction.

Ex. 26: What volume of oxygen gas at NTP is necessary for complete combustion of 20 litre of propane measured at 0^{o}C and 760 mm pressure.

Solution: The balanced equation is

$$C_3H_8 + 5O_2 = 3CO_2 + 4H_2O$$

$1 vol \quad 5 vol$

$1 litre \quad 5 litre$

1 litre of propane requires = 5 litre of oxygen

20 litre of propane will require = 5 x 20 = 100 litre of oxygen at 760 mm pressure and 0^{o}C.

Ex. 27: The percentage by volume of C_3H_8 in a mixture of C_3H_8, CH_4 and CO is 36.5. Calculate the volume of CO_2 produced when 100 mL of the mixture is burnt in excess of O_2.

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

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

$$CO + \frac{1}{2}O_2 \longrightarrow CO_2$$

Let a mL, b mL and c mL be volumes of C_3H_8, CH_4 and CO respectively in 100 mL given simple, then a + b + c = 100 and a = 36.5

Now CO_2 is formed as a result of combustion of mixture.

$\therefore$ Vol. of CO_2 formed = $3a + b + c \left(\begin{array}{l} \because \text{ 1 vol. } C_3H_8 \text{ gives 3 vol. } CO_2 \\ \text{1vol. } CH_4 \text{ gives 1 vol. } CO_2 \\ \text{1 vol. CO gives 1 vol. } CO_2 \end{array} \right)$

$= 3 \times 36.5 + (100 - 36.5)$ = 173 mL

7.2 LIMITING REAGENT

In many situations one of the reactants is present in excess therefore some of this reactant is left over on completion of the reaction.

For example, consider the combustion of hydrogen.

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

Suppose that 2 moles of H_2 and 2 moles of O_2 are available for reaction.

It follows from the equation that only 1 mole of O_2 is required for complete combustion of 2 moles of H_2 ; 1 mole of O_2 will, therefore, be left over on completion of the reaction.

The amount of the product obtained is determined by the amount of the reactant that is completely consumed in the reaction. This reactant is called the limiting reagent.

Thus, limiting reagent may be defined as the reactant which is completely consumed during the reaction.

In the above example H_2 is the limiting reagent. The amount of H_2O formed will, therefore, be determined by the amount of H_2. Since 2 moles of H_2 are taken, it will form 2 moles of H_2O on combustion.

- ❖ The best method to identify limiting reagent is by dividing given moles of each reactant by their stoichiometric coefficient, the one with least ratio is limiting reagent. It is particularly useful when number of reactants are more than two.

Ex. 28: How much magnesium sulphide can be obtained from 2.00 g of magnesium and 2.00 g of sulphur by the reaction $Mg + S \longrightarrow MgS$? Which is the limiting reagent? Calculate the amount of the reactants which remains unreacted.

Solution : First of all each of this masses are converted into moles:

2.00 g of Mg = $\frac{2.00}{24.3}$ = 0.0824 moles of Mg

2.00 g of S = $\frac{2.00}{32.1}$ = 0.0624 moles of S

From the equation, $Mg + S \longrightarrow MgS$,

it follows that one mole of Mg reacts with one mole of S. We are given more moles of Mg than of S.

Therefore, Mg is in excess and some of it will remain unreacted when the reaction is over.

S is the limiting reagent and will control the amount of product.

From the equation, $Mg + S \longrightarrow MgS$,

we note that one mole of S gives one mole of MgS, so 0.0624 mole of S will react with 0.0624 mole of Mg to form 0.0624 mole of MgS.

Molar mass of MgS	=	56.4 g
$\therefore$ Mass of MgS formed	=	0.0624×56.4 g = 3.52 g of MgS
Moles of Mg left unreacted	=	0.0824 – 0.0624 moles of Mg
		= 0.0200 moles of Mg
Mass of Mg left unreacted	=	moles of Mg $\times$ molar mass of Mg
	=	0.0200×24.3 g of Mg = 0.486 g of Mg

Ex. 29: 4 mole of $MgCO_3$ is reacted with 6 moles of HCl solution. Find the volume of CO_2 gas produced at STP. The reaction is

$$MgCO_3 + 2HCl \rightarrow MgCl_2 + CO_2 + H_2O$$

Solution: From the reaction, $MgCO_3 + 2HCl \rightarrow MgCl_2 + CO_2 + H_2O$

Given moles	4 mole		6 mole
Given mole ratio	2	:	3
Stoichiometric Coefficient ratio	1	:	2

There should be one limiting reagent.

To find the limiting reagent, divide the given moles by stoichiometric coefficient.

$MgCO_3$	HCl
$\frac{4}{1} = 4$	$\frac{6}{2} = 3$

HCl is limiting reagent.

$$\because \quad \frac{\text{moles of HCl}}{2} = \frac{\text{moles of CO}_2\text{ produced}}{1}$$

$\therefore$ moles of CO_2 produced = 3 moles

$\therefore$ volumes of CO_2 produced at S.T.P, = $3 \times 22.4 = 67.2L$

Ex. 30: What weight of AgCl will be precipitated when a solution containing 4.77 g NaCl is added to a solution of 5.77g of $AgNO_3$?

Solution:	$AgNO_3$	+ $NaCl$	$\longrightarrow$ $AgCl$	+ $NaNO_3$
mmoles. Mixed	$\frac{5.77}{170}\times 1000$	$\frac{4.77}{58.5}\times 1000$		
	= 33.94	= 81.54	0	0
mmoles. left	0	47.60	33.94	33.94

$\therefore$ mmoles of AgCl formed = 33.94

$$\frac{w}{143.5}\times 1000 = 33.94 \qquad W_{AgCl} = 4.87\,g$$

7.3 CALCULATION INVOLVING PERCENT YIELD

In general, when a reaction is carried out in the laboratory we do not obtain actually the theoretical amount of the product. The amount of the product that is actually obtained is called the actual yield. Knowing the actual yield and theoretical yield, the percentage yield can be calculated as

$$\% \text{ yield} = \frac{\text{Actual yeild}}{\text{Therotical yeild}} \times 100$$

- ❖ The actual amount of any limiting reagent consumed in such incomplete reactions is given by

  [% yield × given moles of limiting reagent] [For reversible reactions].

- ❖ For irreversible reaction with % yield less than 100, the reactants are converted to product (desired) and waste.

7.4 PRINCIPLE OF ATOM CONSERVATION (POAC)

In chemical reaction atoms are conserved, so moles of atoms shall also be conserved. This is known as principle of atomic conservation. This principle is helpful in solving problems of nearly all stoichiometric calculations e.g.

$$KClO_{3(s)} \longrightarrow KCl_{(s)} + O_{2(g)}$$

Applying POAC for K atoms

Moles of K atoms in $KClO_3$ = Moles of K atoms in KCl

Since one mole of $KClO_3$ contains 1 mol of K atom. Similarly 1 mol of KCl contains one mole of K atoms. $\therefore\ 1 \times n_{KClO_3} = 1 \times n_{KCl} \Rightarrow 1 \times \frac{W_{KClO_3}}{M_{KClO_3}} = \frac{W_{KCl}}{M_{KCl}}$ (Mass-mass relationship)

Applying POAC for ‘O’ atoms

Moles of O atom in $KClO_3$ = Moles of O atom in O_2

$$\therefore\ 3 \times n_{KClO_3} = 2 \times n_{O_2} \quad \Rightarrow \quad 3 \times \frac{W_{KClO_3}}{M_{KClO_3}} = 2 \times \frac{\text{Vol. of } O_2 \text{ at STP}}{\text{Standard Molar Volume}}$$

Ex. 31: A sample of $KClO_3$ on decomposition yielded 448 mL of oxygen gas at NTP Calculate :

(i) weight of oxygen produced ,

(ii) weight of $KClO_3$ originally taken

(iii) weight of KCl produced

(K = 39 , Cl = 35.5 and O = 16)

Solution: (i) Mole of oxygen = $\frac{448}{22400} = 0.02$

Wt. of oxygen = 0.02 x 32 = 0.64gm

(ii) $KClO_3 \rightarrow KCl + O_2$

Applying POAC for O atoms,

Moles of O atoms in $KClO_3$ = moles of O atoms in O_2

3 (moles of $KClO_3$)= 2 (moles of O_2)

(1 mole of $KClO_3$ contains 3 moles of O and 1 mole of O_2 contains 2 moles of O)

$$3 \times \frac{wt. of\ KClO_3}{mol.\ wt.\ of\ KClO_3} = 2 \times \frac{vol.\ at\ \ NTP(litre)}{22.4}$$

$$3\times\frac{wt.of\ KClO_3}{122.5}=2\times\frac{vol.\ at\ \ NTP\ (litre)}{22.4}$$

Wt. of $KClO_3$ = 1.634 g

(iii) Again applying POAC for K atoms,

Moles of K atoms in $KClO_3$ = 1 x moles of KCl

(1 mole of $KClO_3$ contains 1 mole of K and 1 mole of KCl contains 1 mole of K)

$$1\times\frac{wt.\ of\ \ KClO_3}{mol.\ wt.\ of\ \ KClO_3}=1\times\frac{wt.of\ KCl}{mol.\ wt.\ of\ \ KCl}$$

Wt. of KCl = 0.9937 g.

Ex. 32: 27.6 g of K_2CO_3 **was treated by a series of reagents so as to convert all of its carbon to $K_2Zn_3[Fe(CN)_6]_2$. Calculate the weight of the product.**

Solution: $K_2CO_3 \xrightarrow[\text{steps}]{\text{Several}} K_2Zn_3[Fe(CN)_6]_2$

Since C atoms are conserved, applying POAC for C atoms,

Moles of C in K_2CO_3 = moles of C in $K_2Zn_3[Fe(CN)_6]_2$

1 x mole of K_2CO_3 = 12 x moles of $K_2Zn_3[Fe(CN)_6]_2$

$$\left(\because 1 \text{ mole of } K_2CO_3 \text{ contains } 1 \text{ mole C } \& \ 1 \text{ mole of } K_2Zn_3[Fe(CN)_6]_2 \text{ contains } 12 \text{ mole of } C \right)$$

$$\frac{\text{wt. of } K_2CO_3}{\text{mol. wt. of } K_2CO_3} = 12 \times \frac{\text{wt. of the product}}{\text{mol. wt. of product}}$$

Wt. of $K_2Zn_3\left[Fe(CN)_6\right]_2 = \dfrac{27.6}{138} \times \dfrac{698}{12} = 11.6$ g.

[mol. wt. of $K_2CO_3 = 138$ and mol. wt. of $K_2Zn_3\left[Fe(CN)_6\right]_2 = 698$]

Ex. 33: 1.84 g of a mixture of $CaCO_3$ and $MgCO_3$ was heated to a constant weight. The constant weight of the residue was found to be 0.96 g. Calculate the percentage composition of the mixture. (Ca = 40, Mg = 24, C = 12, O = 16)

Solution : On heating $CaCO_3$ and $MgCO_3$, one of the products, CO_2, escapes out.

We have,

$$\underset{x\,g}{CaCO_3} + \underset{(1.84-x)g}{MgCO_3} \rightarrow \underset{y\,g}{CaO} + \underset{(0.96-y)g}{MgO} + CO_2\uparrow$$

Applying POAC for Ca atoms,

Moles of Ca atoms in $CaCO_3$ = moles of Ca atoms in CaO

$$1 \times \text{moles of } CaCO_3 = 1 \times \text{moles of } CaO$$

$$\frac{x}{100} = \frac{y}{56} \qquad \begin{bmatrix} CaCO_3 = 100 \\ CaO = 56 \end{bmatrix} \qquad \ldots \text{(i)}$$

Again applying POAC for Mg atoms,

Moles of Mg in $MgCO_3$ = moles of Mg in MgO

$1 \times \text{moles of } MgCO_3 = 1 \times \text{moles of } MgO$

$$\frac{1.84 - x}{84} = \frac{0.96 - y}{40} \begin{bmatrix} MgCO_3 = 84 \\ MgO = 40 \end{bmatrix} \qquad \text{.... (ii)}$$

From eqns. (i) and (ii), we get x = 1 g, y = 0.84 g

$\% \text{ of } CaCO_3 = \frac{1}{1.84} \times 100 = 54.34\ \%$

$\% \text{ of } MgCO_3 = 45.66\ \%$

Ex. 34: What mass of zinc is required to produce hydrogen by reaction with HCl which is enough to produce 4 mol of ammonia according to the reactions.

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

$$3H_2 + N_2 \longrightarrow 2NH_3$$

Solution: The given equations are

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

$$3H_2 + N_2 \longrightarrow 2NH_3$$

From the equations it is clear that

2 mol of NH_3 require = 3 mol of H_2 ; 3 mol of H_2 require = 3 mol of Zn

Thus, 2 mol of NH_3 require = 3 mol of Zn = 3×65 g of Zn

$\therefore$ 4 mol of NH_3 require $= \frac{3 \times 65}{2} \times 4 = 390$ g of Zn.

STOICHIOMETRY

45. What quantity of limestone ($CaCO_3$) on heating will give 56 Kg of CaO?

(A) 1000 Kg (B) 44 Kg (C) 56 Kg (D) 100 Kg

46. 2 mol of H_2S and 11.2 L SO_2 at N.T.P. reacts to form x mol of sulphur; x is

$SO_2 + 2H_2S \longrightarrow 3S + 2H_2O$

(A) 1.5 (B) 3 (C) 11.2 (D) 6

47. How many grams of phosphoric acid (H_3PO_4) would be needed to neutralise 100 g of magnesium hydroxide ($Mg(OH)_2$).

(A) 66.7 g (B) 252 (C) 112.6 g (D) 168 g

48. If two mole of methanol (CH_3OH) completely burns to carbon dioxide and water, the weight of water formed is about
(A) 22 g (B) 18 g (C) 36 g (D) 72 g

49. In an experiment, it is found that 2.0769 g of Pure X produces 3.6769 g of pure X_2O_5. The number of moles of X is

(A) 0.04 (B) 0.06 (C) 0.40 (D) 0.02

50. 2.7 g of Ag_2CO_3 on being heated strongly yields a residue weighing

(A) 2.11 g (B) 2.48 g (C) 2.32 g (D) 2.64 g

51. If 1 mole of ethanol (C_2H_5OH) completely burns to CO_2 and H_2O, the weight of CO_2 formed is about.

(A) 22 g (B) 45 g (C) 66 g (D) 88 g

52. The percent loss in weight after heating a pure sample of $KClO_3$ (molecular weight = 122.5) will be

(A) 12.25 (B) 24.50 (C) 39.18 (D) 49.0

53. Calculate the weight of iron which will be converted into its oxide by the action of 18g of steam on it. From the reaction $2Fe + 3H_2O \longrightarrow Fe_2O_3 + 3H_2$.

(A) 37.3 gm (B) 3.73 gm (C) 56 gm (D) 5.6 gm

54. A 10.0 g sample of a mixture of calcium chloride and sodium chloride is treated with Na_2CO_3 to precipitate the calcium as calcium carbonate. This $CaCO_3$ is heated to convert all the calcium to CaO and the final mass of CaO is 1.62 gms. The % by mass of $CaCl_2$ in the original mixture is :

(A) 15.2% (B) 32.1% (C) 21.8% (D) 11.07%

55. If 0.5 mol of $BaCl_2$ is mixed with 0.2 mol of Na_3PO_4, the maximum number of mol of $Ba_3(PO_4)_2$ that can be formed is

(A) 0.7 (B) 0.5 (C) 0.2 (D) 0.1

56. 0.5 mole of H_2SO_4 is mixed with 0.2 mole of $Ca(OH)_2$. The maximum number of mole of $CaSO_4$ formed is

(A) 0.2 (B) 0.5 (C) 0.4 (D) 1.5

57. For the reaction $A + 2B \longrightarrow C + 3D$, 5 moles of A and 8 moles of B will produce
(A) 5 moles of C (B) 4 moles of C (C) 8 moles of C (D) None of these

58. 2.4 kg of carbon is made to react with 1.35 kg of aluminium to form Al_4C_3. The maximum amount in kg of aluminium carbide formed is
(A) 5.4 (B) 3.75 (C) 1.05 (D) 1.8

59. Consider the reaction $2A \longrightarrow 2B$, $B \longrightarrow 2C$, $3C \longrightarrow 4D$. The no. of moles of D formed starting 4 moles of A, are
(A) 8 (B) 16 (C) 4 (D) 10.67

60. Vapour density of air is (considering air as 80% N_2 and 20% O_2)
(A) 0.001293 (B) 1.293 (C) 14.4 (D) 28.9

61. The density of chlorine relative to air is
(A) 2.44 (B) 3 (C) 71 (D) 4

ANSWER KEY

45. (D)	46. (A)	47. (C)	48. (D)	49. (A)	50. (A)
51. (D)	52. (C)	53. (A)	54. (B)	55. (D)	56. (A)
57. (B)	58. (D)	59. (D)	60. (C)	61. (A)	

8. CONCENTRATION TERMS

- In a solution the designation of solute and solvent is often a matter of convenience, however many a times the one present in smaller quantity is termed as solute.
- Also, a solution may have more than one solute but solvent cannot be more than one.

The following concentrations terms are used to expressed the concentration of a solution. These are

1. Molarity (M)
2. Molality (m)
3. Mole fraction (X)
4. % calculation
5. Normality (N)
6. Ppm

8.1 MOLARITY (M)

The number of moles of a solute dissolved in 1 L (1000 ml) of the solution is known as the molarity of the solution.

$$\boxed{\text{Molarity of solution} = \frac{\text{number of moles}}{\text{volume of solution in litre}}}$$

Let a solution is prepared by dissolving w gm of solute of mol. wt. M in V ml water.

$\therefore$ Number of moles of solute dissolved = $\frac{w}{M}$

$\therefore$ V ml water have $\frac{w}{M}$ mole of solute

$\therefore$ 1000 ml water have $\dfrac{w \times 1000}{M \times V_{\text{in ml}}}$

$\therefore$ Molarity (M) = $\dfrac{w \times 1000}{(\text{Mol. wt of solute}) \times V_{\text{in ml}}}$

Some other relations may also useful.

Number of millimoles = $\dfrac{\text{mass of solute}}{(\text{Mol. wt. of solute})} \times 1000 = (\text{Molarity of solution} \times V_{\text{in ml}})$

$$\boxed{\text{Molarity} = \frac{\text{Number of millimole of solute}}{\text{Total volume of solution in ml}}}$$

Molarity is a unit that depends upon temperature. It varies inversely with temperature. Mathematically : Molarity decreases as temperature increases.

$$\text{Molarity} \propto \frac{1}{\text{temperature}} \propto \frac{1}{\text{volume}}$$

- If a particular solution having volume V_1 and molarity = M_1 is diluted to V_2 mL then

$$M_1V_1 = M_2V_2 \qquad M_2 \text{ : Resultant molarity}$$

- If a solution having volume V_1 and molarity M_1 is mixed with another solution of same solute having volume V_2 mL & molarity M_2

$$M_R \text{ = Resultant molarity} = \frac{M_1V_1 + M_2V_2}{V_1 + V_2}$$

Ex. 35:

149 gm of potassium chloride (KCl) is dissolved in 10 Lt of an aqueous solution. Determine the molarity of the solution (K = 39, Cl = 35.5)

Solution: Molecular mass of KCl = 39 + 35.5 = 74.5 gm

$\therefore$ Moles of KCl = $\frac{149\text{gm}}{74.5\text{gm}} = 2$ $\therefore$ Molarity of the solution = $\frac{2}{10} = 0.2\text{M}$

Ex. 36: What volume of a 3.0 M HCl solution be mixed with 500 mL of a 7 M HCl solution to prepare a HCl solution whose molarity will be 4.0?

Solution: Let V mL of 3.0 M HCl solution is taken, then

$$3\,V + 500 \times 7 = 4 \times (500 + V) \Rightarrow V = 1500\ \text{mL}$$

Ex. 37:

Determine molarity of a solution obtained by mixing 50 mL of a 0.26 M H_2SO_4 solution with another 150 mL 0.48 M H_4SO_4 solution.

Solution: It is a case of mixing of two solutions of different molarities. Applying the mixing formula:

$$\Rightarrow \quad M_3 = \frac{M_1V_1 + M_2V_2}{V_3} = \frac{50 \times 0.26 + 150 \times 0.48}{50 + 150} = 0.425 \text{ M}$$

Ex. 38: What volume of a 5.00 M H_2SO_4 solution should be added to a 150 mL 1.0 M H_2SO_4 solution to obtain a solution of sulphuric acid of molarity 2.5?

Solution: It is again a case of mixing of two solutions. Let us assume that V mL of the stock solution of H_2SO_4 is added.

$\Rightarrow \quad 5V + 150 \times 1.0 = 2.5 \times (150 + V)$

$\Rightarrow \quad 2.5\ V = 225$

$\therefore \quad V = \frac{225}{2.5} = 90\,\text{mL}$

Ex. 39:

A 150 mL 0.25 M NaCl solution, 250 mL 0.45 M $CaCl_2$ solution and a 100 mL 0.60 M $AlCl_3$ solution are mixed together and diluted to a final volume of 750 mL by adding enough water. Determine molarity of chloride ion (Cl^-) in solution assuming that all three salts are completely soluble as well as completely dissociated.

Solution: First we need to calculate total mmoles of Cl^- from the three salt solutions:

mmoles of Cl^- ion from NaCl = 37.5

mmoles of Cl^- ion from $CaCl_2 = 2 \times \text{mmoles of } CaCl_2$

$$= 2 \times 112.5 = 225$$

mmoles of Cl^- ion from $AlCl_3 = 3 \times \text{mmoles of } AlCl_3$

$$= 3 \times 60 = 180$$

$\Rightarrow$ Total mmoles of Cl^- in final solution

$$= 37.5 + 225 + 180 = 442.5$$

$\Rightarrow$ Molarity of $Cl^- = \dfrac{\text{mmoles of } Cl^-}{\text{mL of solution}} = \dfrac{442.5}{750} = 0.59 \text{ M}$

8.2 MOLALITY (m)

The molality is the number of moles of solute present in one Kg of solvent

$$m = \frac{w_{solute} \times 1000}{M.Mass \times w_{solvent}\,(gm)}$$

❖ Molality is independent of temperature changes.

Ex. 40: 255 gm of an aqueous solution contains 5 gm of urea. What is the concentration of the solution in terms of molality? (Mol. wt. of urea = 60)

Solution: Mass of urea = 5 gm

Molecular mass of urea = 60

Number of moles of urea = $\frac{5}{60} = 0.083$

Mass of solvent = (255 – 5) = 250 gm

$\therefore$ Molality of the solution = $\frac{\text{Number of moles of solute}}{\text{Mass of solvent in gram}} \times 1000$

= $\frac{0.083}{250} \times 1000 = 0.333$

Ex. 41: The molarity and molality of a solution are M and m respectively. If the molecular weight of the solute is M', calculate the density of the solution in terms of M, m and M'.

Solution: Let weight of solute be w g and weight of solvent is W g and volume of solution is V mL.

$\therefore$ $$M = \frac{w \times 1000}{M' \times V} \quad \ldots (1) \qquad m = \frac{w \times 1000}{M' \times W} \quad \ldots (2)$$

$$D = \frac{w + W}{V} \quad \ldots (3)$$ By Eq. (1) $$w = \frac{MM'V}{1000} \quad \ldots (4)$$

By Eq. (2) $W = \frac{w \times 1000}{M' \times m} = \frac{MM'V \times 1000}{1000 \times M' \times m}$ by Eq. (4)

$$W = \frac{MV}{m} \quad \ldots (5)$$

$\therefore$ By Eq. (3) $$D = \frac{\frac{MV}{m} + \frac{MM'V}{1000}}{V} \qquad D = M\left[\frac{1}{m} + \frac{M'}{1000}\right]$$

MOLARITY(M) AND MOLALITY(M) FOR PURE SUBSTANCES:

1. Water :

Let the sample of water has 1000 ml

Mass of water = 1000 gm [density of water = 1gm/mL.]

Moles of water $= \frac{1000}{18} mol$

$\therefore Molarity = \frac{\left(\frac{1000}{18}\right)}{1} = 55.55M$ & molality = $\frac{\left(\frac{1000}{18}\right)mol}{1\ kg} = 55.55\ m$

2. Pure ethanol :

d gm/ml (density of ethanol)

(C_2H_5OH) let volume of ethanol taken be 1000 ml.

$\therefore wt\,of$ ethanol in 1000 ml = $1000 \times d\ gm$

Mol of ethanol $= \frac{1000d}{46}$ $\therefore$ Molarity $= \frac{1000d}{46}$

& molality of ethanol = $\frac{\left(\frac{1000d}{46}\right) mol}{\left(\frac{1000d}{1000}\right) kg} = \frac{1000}{46}$

- **Parts per million (ppm)** $\rightarrow$ Amount of solute (in g) with 10^6 g solvent

- **Parts per billion (ppb)** $\rightarrow$ Amount of solute (in g) with 10^9 g solvent

8.3 MOLE FRACTION (X)

The ratio of number of moles of the solute or solvent present in the solution and the total number of moles present in the solution is known as the mole fraction of substance concerned.

Let number of moles of solute in solution = n

Number of moles of solvent in solution = N

$\therefore$ Mole fraction of solute $(X_1) = \dfrac{n}{n+N}$

$\therefore$ Mole fraction of solvent $(X_2) = \dfrac{N}{n+N}$

Also $X_1 + X_2 = 1$

❖ Mole fraction is a pure number. It will remain independent of temperature changes.

8.4 PERCENTAGE CONCENTRATION

The concentration of a solution may also be expressed in terms of percentage in the following way.

1. % WEIGHT BY WEIGHT (W/W):

It is given as mass of solute present in per 100 gm of solution.

$$\% \text{ w/w} = \frac{\text{mass of solute in gm}}{\text{mass of solution in gm}} \times 100$$

2. % WEIGHT BY VOLUME (W/V) :

It is given as mass of solute present in per 100 ml of solution

$$\% \text{ w/v} = \frac{\text{mass of solute in gm}}{\text{volume of solution in ml}} \times 100$$

3. % VOLUME BY VOLUME (V/V) :

It is given as volume of solute present in per 100 ml solution.

$$\% \text{ V/V} = \frac{\text{Volume of solute in ml}}{\text{Volume of solution in ml}} \times 100$$

Ex. 42: 0.5 g of a substance is dissolved in 25 g of a solvent. Calculate the percentage amount of the substance in the solution.

Solution: Mass of substance = 0.5 g Mass of solvent = 25 g

$\therefore$ Percentage of the substance (w/w) $= \frac{0.5}{0.5+25} \times 100 = 1.96$

Ex. 43: 20cm^3 of an alcohol is dissolved in 80cm^3 of water. Calculate the percentage of alcohol in solution.

Solution: Volume of alcohol = $20cm^3$ Volume of water

= $80cm^3$

$\therefore$ Percentage of alcohol $= \frac{20}{20+80} \times 100 = 20$%

CONCENTRATION TERMS

63. An aqueous solution of urea containing 18 g urea in 1500 cc of solution has a density of 1.052 g/cc. If the mol.wt. of urea is 60, then the molality of solution is

(A) 0.2 (B) 0.192 (C) 0.064 (D) 1.2

64. Molarity of 1g H_2SO_4 solution in 1 lit. water is nearly

(A) 0.1 (B) 0.20 (C) 0.05 (D) 0.01

65. 20 ml of 0.2 M $Al_2(SO_4)_3$ is mixed with 20 ml of 0.6 M $BaCl_2$. Concentration of Al^{3+} ion in the solution will be

(A) 0.2 M (B) 10.3 M (C) 0.1 M (D) 0.25 M

66. 50 ml of 0.01 M $FeSO_4$ will react with what volume of 0.01 M $KMnO_4$ solution in acid medium?
(1 mole $KMnO_4$ requires 5 mole of $FeSO_4$ for complete reaction)

(A) 50 ml (B) 25 ml (C) 100 ml (D) 10 ml

67. The number of H^+ ions present in 100 ml of 0.001M H_2SO_4 solution will be

(A) 120.4×10^{19} (B) 1.20×10^{20} (C) 6.023×10^{20} (D) 6.023×10^{21}

68. 3.0 molal NaOH solution has a density of 1.11 g/ml. The molarity of the solution is

(A) 2.97 (B) 3.05 (C) 3.64 (D) 3.050

69. 250 ml of a sodium carbonate solution contains 2.65 grams of Na_2CO_3. If 10 ml of this solution is diluted to one litre, what is the concentration of the resultant solution? (mol wt. of Na_2CO_3 = 106)

(A) 0.1 M (B) 0.001M (C) 0.01 M (D) 10^{-4} M

70. The mole fraction of NaCl in a solution containing 1 mole of NaCl in 1000g of water is:
(A) 0.0177 (B) 0.001 (C) 0.5 (D) 1.5

ANSWER KEY

63. (B)	64. (D)	65. (A)	66. (D)
67. (B)	68. (A)	69. (B)	70. (A)

PYQ

1. Which of the following concentration factor is affected by change in temperature ?
[AIEEE 2002]
(A) Molarity (B) Molality (C) Mole fraction (D) Weight

3. Number of atoms in 560g of Fe (atomic mass 56g/mol) is :
[AIEEE 2002]
(A) Twice that of 70g N (B) Half that of 20g H
(C) Both (A) and (B) (D) None of these

4. In an organic compound of molar mass 108 g/mol C, H and N atoms are present in 9 : 1 : 3.5 by weight. Molecular formula can be :
[AIEEE 2002]
(A) $C_6H_8N_2$ (B) $C_7H_{10}N$ (C) $C_5H_6N_3$ (D) $C_4H_{18}N_3$

4. What volume of hydrogen gas at 273 K and 1 atm pressure will be consumed in obtaining 21.6 gm of elemental boron (atomic mass = 10.8) from the reduction of boron trichloride by hydrogen?
[AIEEE 2003]
(A) 44.8 lit. (B) 22.4 lit. (C) 89.6 lit. (D) 67.2 lit.

5. 6.02×10^{20} molecules of urea are present in 100 ml of its solution. The concentration of urea solution is

[AIEEE 2004]

(A) 0.001 M (B) 0.01 M (C) 0.02 M (D) 0.1 M

6. If we consider that 1/6, in place of 1/12, mass of carbon atom is taken to be the relative atomic mass unit, the mass of one mole of a substance will

[AIEEE 2005]

(A) decrease twice
(B) increase two fold
(C) remain unchanged
(D) be a function of the molecular mass of the substance

7. Density of a 2.05M solution of acetic acid in water is 1.02 g/ml. The molality of the solution is :

[AIEEE-2006]

(A) 1.14 mol/kg (B) 3.28 mol/kg (C) 2.28 mol/kg (D) 0.44 mol/kg

8. How many moles of magnesium phosphate , $Mg_3(PO_4)_2$ will contains 0.25 mole of oxygen atoms?

(A) 0.02 (B) 3.125×10^{-2} (C) 1.25×10^{-2} (D) 2.5×10^{-2} **[AIEEE-2006]**

9. The density (in g/mL) of a 3.60 M sulphuric acid solution that is 29% by mass will be **[AIEEE-2007]**

(A) 1.22 (B) 1.45 (C) 1.64 (D) 1.88

10. The reaction,

$$2Al(s) + 6HCl(aq) \rightarrow 2Al^{3+}(aq) + 6Cl^{-}(aq) + 3H_2(g)$$

(A) 33.6 L $H_2(g)$ is produced regardless of temperature and pressure for every mole of Al that reacts
(B) 67.2 L $H_2(g)$ at STP is produced for every mole of Al that reacts
(C) 11.2 L $H_2(g)$ at STP is produced for every mole of HCl (aq) consumed
(D) 6 L HCl (aq) is consumed for every 3L H_2 (g) produced **[AIEEE-2007]**

11. A 5.2 molal aqueous solution of methyl alcohol, CH_3OH is supplied. What is the mole fraction of methyl alcohol in the solution
(A) 0.86 (B) 0.086 (C) 0.043 (D) 1.0 **[AIEEE-2007]**

12. The density of a solution prepared by dissolving 120 g of urea (Mol.Mass = 60u) in 1000 g of water is 1.15 g/mL. The molarity of this solution is
(A) 1.02 M (B) 0.50 M (C) 2.05 M (D) 1.78 M **[AIEEE-2007]**

13. The Molarity of a solution obtained by mixing 750 mL of 0.5 (M) HCl with 250 mL of 2(M) HCl will be (A) 1.75 M (B) 0.975 M (C) 0.875 M (D) 1.78 M **[AIEEE-2013]**

14. Number of atoms in the following samples of substances is the largest in **[JEE Main Online 2013]**

(a) 4.0 g of hydrogen
(b) 70.0g of chlorine
(c) 127.0 g of iodine
(d) 48.0 g of magnesium

15. The number of protons, electrons and neutrons in a molecule of heavy water are respectively

[JEE Main Online 2013]

(a) 8, 10, 11
(b) 10, 10, 10
(c) 10, 11, 10
(d) 11, 10, 10

16. A gaseous hydrocarbon gives upon combustion 0.72 g of water and 3.08 g of CO_2. The empirical formula of the hydrocarbon is **[JEE Main Online 2013]**

(a) C_2H_4 (b) C_3H_4 (c) C_6H_5 (d) C_7H_8

17. The density of 3M solution of sodium chloride is 1.252 g mL^{-1}. The molality of the solution will be

(molar mass, NaCl = 58.5 g mol^{-1}) **[JEE Main Online 2013]**

(a) 2.60 m (b) 2.18 m (c) 2.79 m (d) 3.00 m

18. 10 mL of 2M NaOH solution is added to 200 mL of 0.5 LM of NaOH solution. What is the final concentration? **[JEE Main Online 2013]**

(a) 0.57 M (b) 5.7 M (c) 11.4 M (d) 1.14 M

PYQ

1. (A)	2. (C)	3. (A)	4. (D)	5. (B)	6. (C)
7. (C)	8. (B)	9. (A)	10. (C)	11. (B)	12. (C)
13. (C)	14. (A)	15. (B)	16. (D)	17. (C)	18. (A)