UNIT-1
MOLE CONCEPT

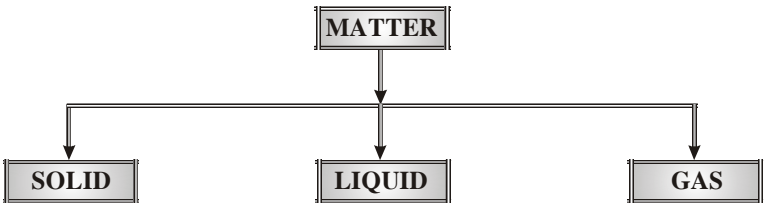
CLASSIFICATION OF MATTER (Read Yourself)

□ ***MATTER :*** Matter is anything that has mass and occupies space.

Two ways of classifying matter.

I. Physical classification II. Chemical classification

I. Physical classification :

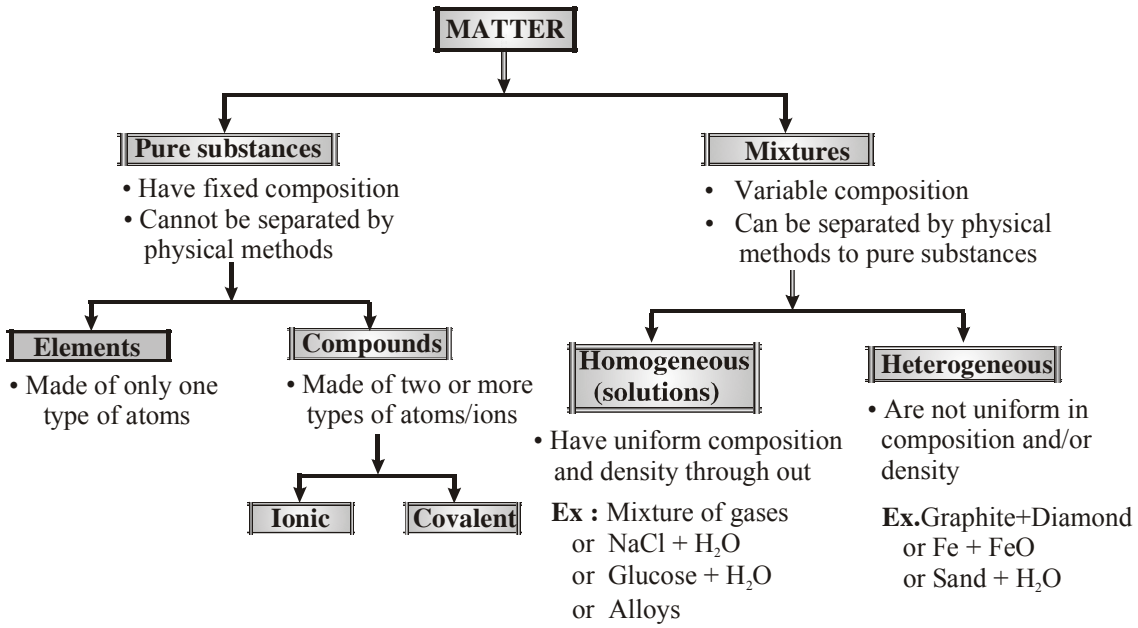


(i)	Particles held very closely packed in ordered manner.	Particles are less closely packed.	Particles are farthest apart
(ii)	No freedom of movement of particles	Particles can move around to some extent	Movement of particles is very easy and fast
(iii)	Definite shape and volume	Definite volume, indefinite shape	indefinite shape and volume
(iv)	Exists at low T and high P	Exists at intermediate P & T	Exists at high T and low P

Note : For same substance :

- Solid and Liquid co-exist at ***MELTING POINT***.
- Liquid and gas co-exist at ***BOILING POINT***.
- Solid and gas co-exist at ***SUBLIMATION POINT***.
- Solid, liquid and gas co-exist at ***TRIPLE POINT***.

II. Chemical classification :



Note
- **PHASE :** It is the state of matter uniform in density and composition.
- Homogeneous mixtures have single phase while heterogeneous mixtures are multi-phase.
  Ex : $NaCl + H_2O$ mixture has one phase
  Ex : Graphite + Diamond mixture has 2 phases.

SOME SPECIFIC PROPERTIES OF SUBSTANCES

❖ **Deliquescence :**

The property of certain compounds of taking up the moisture present in atmosphere and becoming wet when exposed, is known as deliquescence. These compounds are known as deliquescent. Sodium hydroxide, potassium hydroxide, anhydrous calcium chloride, anhydrous magnesium chloride, anhydrous feric chloride, etc., are the examples of deliquescent compounds.

❖ **Hygroscopicity :**

Certain compounds combine with the moisture of atmosphere and are converted into hydroxides or hydrates. Such substances are called hygroscopic. Anhydrous copper sulphate, quick lime (CaO), anhydrous sodium carbonate, etc., are of hygroscopic nature.

❖ **Efflorescence :**

The property of some crystalline substances of losing their water of crystallisation on exposure and becoming powdery on the surface is called efflorescence and such salts are know as efflorescent. The examples are : Ferrous sulphate ($FeSO_4.7H_2O$), sodium carbonate ($Na_2CO_3.10H_2O$), sodium sulphate ($Na_2SO_4.10H_2O$), potash alum [$K_2SO_4.Al_2(SO_4)_3.24H_2O$], etc.

❖ **Malleability :**

This property is shown by metals. When metallic solid is being beaten, it does not break but is converted into thin sheet. It is said to possess the property of malleability. Copper, gold, silver, aluminium, lead, etc., can be easily hammered into sheets. Gold is the most malleable metal.

❖ **Ductility :**

The property of metal to be drawn into wires is termed ductility. Copper, silver, gold, aluminium, iron, etc., are ductile in nature. Platinum is the most ductile metal.

❖ **Brittleness :**

The solid materials which break into small pieces on hammering are called brittle. The solids of non-metals are generally brittle in nature.

Ex: Ice, Diamond etc.

THE LAW OF CHEMICAL COMBINATION

Atoine Lavoisier, John Dalton and other scientists formulated certain laws concerning the composition of matter and chemical reactions. These laws are known as the law of chemical combination.

I. Law of indestructibility of matter or conservation of Mass :

Antoine Lavosier
(1743-1794)

Antoine-Laurent de Lavosier, the"father of modern chemistry," wasa French nobleman prominent in the histories of chemistry and biology. He named both oxygen and hydrogen and predicted silicon.

- This law was proposed by ***Lavoisier in 1774.***
- The experimental certification was given by Landolt.
- According to this law in all physical or chemical changes the total mass of the system remains constant or in a physical or chemical change, mass is neither created nor destroyed. Thus, in a chemical change-

Total mass of reactant reacted = Total mass of products formed

Ex. $H_2O(s) \longrightarrow H_2O(\ell)$

Above reaction shows the physical change and the wt. of H_2O (s) = wt. of $(H_2O)(\ell)$

In case the reacting materials are not completely consumed, the relationship will be.

Total masses of reactants = Total masses of product + masses of unreacted reactants

- In nuclear reactions (Mass + energy) is conserved not the mass seperately.

Ex.1 ***When 4.2 g $NaHCO_3$ is added to a solution of CH_3COOH weighing 10.0 g, it is observed that 2.2 g CO_2 is released into atmosphere. The residue is found to weigh 12.0 g. Show that these observations are in agreement with the law of conservation of mass.***

Sol. $NaHCO_3 + CH_3COOH \longrightarrow CH_3COONa + H_2O + CO_2$

Initial mass = 4.2 + 10 = 14.2

Final mass = 12 + 2.2 = 14.2

Thus, during the course of reaction law of conservation of mass is obeyed.

II. Law of constant or definite proportion :

Joseph Proust
(1754 - 1826)

Proust was born the son of anapothecary at Angers in north-west France. He studied in Paris.He lived in poverty for some years before being awarded a pension by Louis XVIII.

- This law was given by ***Joseph Louis Proust. in 1799***.
- Chemical compostion of a compound remains constant whether it is obtained by any method or any source.

◆ **Example :**

In water (H_2O), Hydrogen and Oxygen combine in 1 : 8 mass ratio, the ratio remains constant whether it is tap water, river water or sea water or produced by any chemical reaction.

Ex.2 ***1.80 g of a certain metal burnt in oxygen gave 3.0 g of its oxide. 1.50 g of the same metal heated in steam gave 2.50 g of its oxide. Show that these results illustrate the law of constant proportion.***

Sol. In the first sample of the oxide,

wt. of metal = 1.80 g , wt. of oxygen = (3.0 – 1.80) g = 1.2 g

$$\therefore \quad \frac{\text{wt.of metal}}{\text{wt.of oxygen}} = \frac{1.80\text{g}}{1.2\text{g}} = 1.5$$

In the second sample of the oxide,

wt. of metal = 1.50 g , wt. of oxygen = (2.50 – 1.50) g = 1 g

$$\therefore \quad \frac{\text{wt.of metal}}{\text{wt.of oxygen}} = \frac{1.50\text{g}}{1\text{g}} = 1.5$$

Thus, in both samples of the oxide the proportions of the weights of the metal and oxygen are fixed. Hence, the results follows the law of constant proportion.

Note: This law is not applicable in case of isotopes.

III. The law of multiple proportion :

- This law was given by Dalton in 1804.
- If two elements combine to form more than one compound, then the different masses of one element which combine with a fixed mass of the other element, bear a simple ratio to one another.

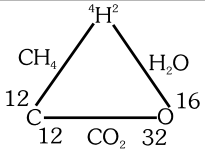
Ex. Nitrogen and oxygen combine to form five stable oxides –

N_2O	Nitrogen 28 parts	Oxygen 16 parts
N_2O_2	Nitrogen 28 parts	Oxygen 32 parts
N_2O_3	Nitrogen 28 parts	Oxygen 48 parts
N_2O_4	Nitrogen 28 parts	Oxygen 64 parts
N_2O_5	Nitrogen 28 parts	Oxygen 80 parts

The masses of oxygen which combine with same mass of nitrogen in the five compounds bear a ratio 16 : 32 : 48 : 64 : 80 or 1 : 2 : 3 : 4 : 5.

Note: This law is not applicable in case of isotopes.

IV. Law of reciprocal proportion (or law of equivalent wt.) :



This law was put forward by ***Richter in 1792***. It states as follows : The ratio of the weights of two elements A and B which combine separately with a fixed weight of the third element C is either the same or some simple multiple of the ratio of the weights in which A and B combine directly with each other. This law may be illustrated with the help of the following example.

The elements C and O combine separately with the third element H to form CH_4 and H_2O and they combine directly with each other to form CO_2, as shown in fig.

In CH_4, 12 parts by weight of carbon combine with 4 parts by weight of hydrogen. In H_2O, 2 parts by weight of hydrogen combine with 16 parts by weights of oxygen. Thus the weight of C and O which combine with fixed weight of hydrogen (say 4 parts by weight) are 12 and 32 i.e. they are in the ratio 12 : 32 or 3 : 8.

Now in CO_2, 12 parts by weight of carbon combine directly with 32 parts by weight of oxygen i.e. they combine directly in the ratio 12 : 32 or 3 : 8 which is the same as the first ratio.

- **Special Note :** This law is also called as law of equivalent wt. due to each element combined in their equivalent wt. ratio.

Ex.3 ***Ammonia contains 82.35% of nitrogen and 17.65% of hydrogen. Water contains 88.90% of oxygen and 11.10% of hydrogen. Nitrogen trioxide contains 63.15% of oxygen and 36.85% of nitrogen. Show that these data illustrate the law of reciprocal proportions.***

Sol. In NH_3, 17.65g of H combine with N = 82.35g

$\therefore$ 1 g of H combine with N = $\frac{82.35}{17.65}$ g = 4.66 g

In H_2O, 11.10 g of H combine with O = 88.90 g

$\therefore$ 1 g of H combine with O = $\frac{88.90}{11.10}$ g = 8.00 g

$\therefore$ Ratio of the weights of N and O which combine with fixed weight (=1g) of H = 4.66 : 8.00 = 1 : 1.7

In N_2O_3, ratio of weights of N and O which combine with each other = 36.85 : 63.15 = 1 : 1.7

Thus the two ratios are the same. Hence it illustrates the law of reciprocal proportions.

V. Law of Gaseous volumes :

- This law was given by ***Gay-Lussac.*** in 1808.
- According to this law, gases react with each other in the simple ratio of their volumes. If products are also gases then they are also in simple ratio of volume provided that all volumes are measure at same temp. & pressure.

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

1 vol. 3vol. 2vol.

Joseph Louis
Gay Lussac
(1778 - 1850)

Joseph Louis Gay-Lussac also ; 6 December 1778 – 9 May1850)was a French chemist and physicist. He is known mostly for two laws related to gases, and for his work on alcohol-water mixtures , which led to the degrees Gay-Lussac used to measure alcoholic beverages in many countries.

Ex.4 ***For the gaseous reaction , $H_2 + Cl_2 \longrightarrow 2HCl$. If 40 ml of hydrogen completely reacts with chlorine then find out the required volume of chlorine and volume of produced HCl ?***

Sol. According to Gay Lussac's Law :

$H_2 + Cl_2 \longrightarrow 2HCl$

$\because$ 1 ml of H_2 will react will 1 ml of Cl_2 and 2 ml of HCl will produce.

$\therefore$ 40 ml of H_2 will react with 40 ml of Cl_2 and 80 ml of HCl will produce.

required vol. of Cl_2 = 40 ml, produced vol. of HCl = 80 ml

VI. Berzelius Hypothesis and Avogadro's Hypothesis :

(A) Berzelius Hypothesis : Equal volumes of all gases under similar conditions of temperature and pressure contain equal number of atoms.

The above statements was incorrect and later it was modified by Avogadro.

(B) Avogadro's Hypothesis : Equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules.

Berzelius
(1779 - 1848)

Jöns Jacob Berzelius was a Swedish chemist. He worked out the modern technique of chemical formula notation, and is together with John Dalton, Antoine Lavoisier, and Robert Boyle considered a father of modern chemistry.

Lorenzo Romano Amedeo Carlo Avogadro di Quareqa edi Carreto (1776-1856)

Italian mathematical physicist. He practiced law for many years before he became interested in science. His most famous work, now known as Avogadro's law, was largely ignored during his lifetime, although it became the basis for determining atomic masses in the late nineteenth century.

❖ Application of Avogadro's hypothesis :

(a) In finding the atomicity

(b) Relation between molecular weight and vapour density

- Vapour density $= \frac{\text{density of gas}}{\text{density of } H_2}$ (constant P, T)
- Molecular weight $= 2 \times$ vapour density
- Density of $H_2 = 0.000089$ gm/cm^3

≈ 0.00009 gm/ml ≈ 0.089 gm/lit.

(c) Relation between molecular weight and volume.

- 1 molecular weight = 22.7 lit. volume of gas at STP
- Weight of 1 mole gas = weight of 22.7 lit gas at STP
- Gram molecular volume or Molar volume = 22.7 litre at STP

(d) Finding the molecular formula of gas.

DALTON'S ATOMIC THEORY

Ancient Indian and Greek philosophers have always wondered about the unknown and unseen form of matter. The idea of divisibility of matter was considered long back in India, around 500 BC. An Indian philosopher ***Maharishi Kanad***, postulated that if we go on dividing matter (padarth), we shall get smaller and smaller particles. Ultimately, a time will come when we shall come across the smallest particle beyond which further division will not be possible. He named these particles Parmanu. Another Indian philosopher, Pakudha Katyayama, elaborated this doctrine and said that these particles normally exist in a combined form which gives us various forms of matter. Around the same era, the Greek philosopher Democritus
expressed the belief that all matter consists of very small, indivisible particles, which he named ***atomos*** (meaning uncuttable or indivisible).

John Dalton (1766 - 1844) , an Englishman, began teaching at a Quaker school when he was 12. His fascination with science included an intense interest in meterology (he kept careful daily weather records for 46 years), which led to an interest in the gases of the air and their ultimate components, atom. Dalton is best known for his atomic theory, in which he postulated that the fundamental differences among atoms are their masses. He was the first to prepare a table of relative atomicweight.

Although Democritus' ideal was not accepted by many of his contemporaries (notably Plato and Aristotle), some how it endured. Experimental evidence from early scientific investigations provided support for the notion of "atomism" and gradually gave rise to the modern definitions of elements and compounds. It was in ***1808, John Dalton***, formulated a precise definition of the indivisible building blocks of matter that we call atoms. Dalton's work marked the beginning of the modern era of chemistry. The hypotheses about the nature of matter on which Dalton's atomic theory is based can be summarized as follows :

(i) Elements are composed of extremely small particles called atoms.

(ii) All atoms of a given element are identical, having the same size, mass and chemical properties. The atoms of one element are different from the atoms of all other elements.

(iii) Compounds are composed of atoms of more than one element. In any compound, the ratio of the numbers of atoms of any two of the elements present is either an integer or a simple fraction.

(iv) A chemical reaction involves only the separation, combination or rearrangement of atoms ; it does not result in their creation or destruction.

☐ *LIMITATIONS OF DALTON'S ATOMIC THEORY :*

According to Dalton's atomic theory, an atom is the ultimate, discrete and indivisible particle of matter. Later researches proved that Dalton's atomic theory was not wholly correct.

Dalton's atomic theory suffered from the following drawbacks :

(i) Atoms of the same or different types have a strong tendency to combine together to form a new 'group of atoms'. For example, hydrogen, nitrogen, oxygen gases exist in nature as 'group of two atoms'. This indicates that the smallest unit capable of independent existence is not an atom, but a 'group of atoms'.

(ii) With the discovery of sub-atomic particles, e.g., electrons, neutrons and protons, the atom can no longer be considered indivisible.

(iii) Discovery of isotopes indicated that all atoms of the same element are not perfectly identical. At least, they differ in their masses. Atoms of the same element having different masses are called isotopes.
Dalton's atomic theory could not explain why certain substances, all containing atoms of the same element, should differ in their properties. For example, charcoal, graphite and diamond all are made up of only Carbon-atoms, but still their properties are quite different.

ATOMIC AND MOLECULAR MASSES

☐ DIFFERENT TYPES OF ATOMIC MASSES :

The mass of an atom depends on the number of electrons, protons, and neutrons it contains. Knowledge of an atom's mass is important in laboratory work. But atoms are extremely small particles - even the smallest speck of dust that our unaided eyes can detect contains as many as 1×10^{16} atoms ! Clearly we cannot weigh a single atom, but it is possible to determine the mass of one atom relative to another experimentally. The first step is to assign a value to the mass of one atom of a given element so that it can be used as a standard.

❖ RELATIVE ATOMIC MASS :

Hydrogen, being lightest atom was arbitrarily assigned a mass of 1 (without any units) and other elements were assigned masses relative to it. However, the present system of atomic masses is based on carbon - 12 as the standard and has been agreed upon in 1961. Here, Carbon - 12 is one of the isotopes of carbon and can be represented as ^{12}C. In this system, ^{12}C is assigned a mass of exactly 12 atomic mass unit **(amu)** and masses of all other atoms are given relative to this standard. **Relative Atomic Mass is defined as the number which indicates how many times the mass of one atom of an element is heavier in comparison to 1/12th part of the mass of one atom of C-12.**

$$\text{Relative atomic mass of an element} = \frac{\text{mass of one atom of an element}}{\frac{1}{12}[\text{mass of one C-12 atom}]}$$

$$= \frac{\text{Mass of one atom of an element}}{1\,\text{amu}}$$

❖ ***ATOMIC MASS UNIT (a.m.u. or u) :*** The quantity $1/12^{th}$ mass of an atom of C^{12} is known as atomic mass unit.

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

$\therefore$ $1/12^{th}$ part of the mass of 1 atom $= \frac{1.9924 \times 10^{-23} g}{12} = 1.67 \times 10^{-24}$ g $= \frac{1}{6.022 \times 10^{23}}$ g

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

❖ ***GRAM ATOMIC MASS OR MASS OF 1 GRAM ATOM :***

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

gram atomic mass (GAM) = mass of 1 **gram atom** = mass of 1 **mole atoms**

= mass of N_A atoms = mass of 6.022×10^{23} atoms.

Ex. GAM of oxygen= mass of 1 **g atom** of oxygen = mass of 1 **mol atoms** of oxygen.

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

Ex. Mass of one atom of Oxygen = 16 amu or $16 \times 1.67 \times 10^{-24}$ g

Mass of N_A atoms of Oxygen = $16 \times 1.67 \times 10^{-24} \times 6.022 \times 10^{-23}$ g = 16 g

Now see the table given below and understand the definition given before.

Element	R.A.M. (Relative Atomic Mass)	Atomic mass (mass of one atom)	Gram Atomic mass or weight
N	14	14 amu	14 gm
He	4	4 amu	4 gm
C	12	12 amu	12 gm

❖ **AVERAGE ATOMIC MASS :**

If an element exists in different isotopic forms (or allotropic forms) having relative abundance X_1%, X_2% ..... X_n%, with relative atomic masses M_1, M_2 ..... M_n respectively then ,

Avg. Atomic mass of element $= \frac{X_1}{100}(M_1) + \frac{X_2}{100}(M_2) + ... + \frac{X_n}{100}(M_n) = \sum_{i=1 \text{ to } n} \frac{X_i}{100}(M_i)$

Ex.5 ***The atomic mass of an element is 50***

(i) Calculate the mass of one atom, in amu

(ii) Calculate the mass of 6.022×10^{23} atoms, in gm

(iii) Calculate the number of atoms in its 10 gm

(iv) What mass of the element contains 3.011×10^{20} atoms

Sol. (i) 50 amu (ii) 50 gm

(iii) $\because$ 50 gm of element contains 6.022×10^{23} atoms

$\therefore$ 10 gm of element will contain $\frac{6.022 \times 10^{23}}{50} \times 10 = 1.2044 \times 10^{22}$ atoms

(iv) $\because 6.022 \times 10^{23}$ atoms weighs 50 gm

$\therefore 3.011 \times 10^{20}$ atoms weighs $\frac{50}{6.022\times10^{23}} \times 3.011 \times 10^{20} = 0.025$ gm

Ex.6 ***An element exist in nature in two isotopic forms : X^{30} (90%) and X^{32}(10%). What is the average atomic mass of element ?***

Sol. Av. atomic mass $= \frac{\Sigma(\%\text{abundance} \times \text{atomic mass})}{100} = \frac{90\times30+10\times32}{100} = 30.2$

□ RELATIVE MOLECULAR MASS :

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

OR

The molecular mass of a substance 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.

Ex. molecular mass of oxygen (O_2) = 32

molecular mass of (O_3) = 48

molecular mass of HCl = 1 + 35.5 = 36.5

molecular mass of H_2SO_4 = 2 + 32 + 64 = 98

❖ GRAM MOLECULAR MASS (MASS OF 1 GRAM MOLECULE) :

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

gram molecular mass (GMM) = mass of 1 **gram molecule** = mass of 1 **mole molecules**

= mass of N_A molecules = mass of 6.022×10^{23} molecules

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

= mass of 1 **mole molecules** of H_2SO_4

= mass of N_A molecules of H_2SO_4

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

Ex. Molecular Mass of N_2 = 28 amu = $28 \times 1.67 \times 10^{-28}$ g

Mass of N_A molecules of N_2 = $28 \times 1.67 \times 10^{-24} \times 6.022 \times 10^{23}$ g = 28 g

❖ **AVERAGE MOLECULAR MASS OF NON-REACTING GAS MIXTURE :**

$$M_{avg.} = \frac{\text{Total mass of mixture}}{\text{Total mole}}$$

Ex.7 ***The molecular mass of a compound is 75***

(i) Calculate the mass of 100 molecules, in amu.

(ii) Calculate the mass of 5000 molecules, in gm.

(iii) What is the mass of 6.022×10^{20} molecules, in gm

(iv) How many molecules are in its 2.5 mg

Sol. (i) mass of 1 molecules = 75 amu

$\therefore$ mass of 100 molecules = 7500 amu

(ii) Mass of 5000 molecules = 5000×75 amu

$= 5000 \times 75 \times 1.67 \times 10^{-24} = 6.26\ 25 \times 10^{-19}$ gm

(iii) $\therefore$ 6.022×10^{23} molecules weighs 75 gm

$\because$ 6.022×10^{20} molecules weighs $\frac{75}{6.022 \times 10^{23}} \times 6.022 \times 10^{20} = 0.075$ gm

(iv) $\therefore$ 75 gm compound contains 6.022×10^{23} molecules

$\because$ 2.5×10^{-3} gm will contain $\frac{6.022 \times 10^{23}}{75} \times 2.5 \times 10^{-3} = 2.007 \times 10^{19}$ molecules.

Ex.8 ***A gaseous mixture contains 40% H_2 and 60% He, by volume. What is the average molecular mass of mixture ?***

Sol. $M_{av} = \frac{\Sigma(\%\text{ by vol.} \times \text{molecular mass})}{100} = \frac{40 \times 2 + 60 \times 4}{100} = 3.20$

INTRODUCTION TO MOLE

Atoms and molecules are extremely small in size and their numbers in even a small amount of any substance is really very large. To handle such large numbers, a unit of similar magnitude is required. The 14th Geneva conference on weight and measures adopted mole as a ***seventh basic SI unit of the amount of a substance.*** Mole concept is essential tool for the fundamental study of chemical calculations. This concept is simple but its application requires a thorough practice. There are many ways of measuring the amount of substance, weight and volume being the most common, but basic unit of chemistry is the atom or a molecule and measuring the number of molecule is more important.

□ **DEFINITION OF MOLE AND MOLAR MASS :**

- A mole is the amount of a substance that contains as many entities (Atoms, Molecules, Ions or any other particles) as there are atoms in exactly 12 g of C-12 isotope.
- A mole of a substance contains Avogadro's number (6.022×10^{23}) of particles.

The term mole, like a dozen or a gross, thus refers to a particular number of things. A dozen eggs equals 12 eggs, a gross of pencils equals 144 pencils, and a mole of ethanol equal 6.022×10^{23} ethanol molecules.

- The ***molar mass*** of a substance is the mass of one mole of the substance. Carbon-12 has a molar mass of exactly 12 g/mol , by definition.
- 1 gram-atom = 1 mole atoms = N_A atoms
- 1 gram-molecule = 1 mole molecules = N_A molecules
- 1 gram-ion = 1 mole ions = N_A ions

❖ **Methods to calculate moles :**

(i) If number of particles (molecules or atoms) is given then,

$$\text{mole} = \frac{\text{Given number of molecule / atom}}{N_A}$$

(ii) If mass is given then, number of mole = $\frac{\text{Given mass of substance (in gm)}}{\text{GAM / GMM}}$

(iii) If volume of gas is given then, mole

$$= \frac{\text{Volume of gas at STP}}{22.7\text{ L}} = \frac{\text{Volume of gas at } 0^\circ\text{C and 1 atm}}{22.4\text{L}}$$

(Standard molar volume is the volume occupies by 1 mole of any gas at NTP or STP, which is equal to 22.7 L)

(iv) Under any condition of temperature and pressure, moles of gases may be calculated using IDEAL GAS EQUATION : PV = nRT,

where, R = Universal Gas Constant

= 0.0821 L-atm/K-mol

= 8.314 J/K-mol

$\approx$ 2 cal/K-mol

Units of pressure and their relation:

1 atm = 76 cm Hg

= 760 mm Hg

= 760 torr (1 torr = 1 mm Hg)

= 1.01325×10^6 dyne/cm^2

= 1.01325×10^5 N/m^2 or Pa

= 1.01325 bar (1 bar = 10^5 Pa)

1 bar = 75 cm Hg

Units of Volume and their relation:

1 ml = 1 cm^3 = 1 c.c.

1 Litre = 1000 ml = 1 dm^3

1 m^3 = 1000 L

Units of Temperature and their relation:

T = 273 + t

where, T = Absolute temperature (in Kelvin) and t = temperature in ^oC

(v) Sometimes gas is collected over water. In this case, the measured pressure is sum of pressure of gas and the vapour pressure of water (also called Aqueous Tension). In order to calculate moles of gas, the vapour pressure of water should be deducted from the measured pressure.

Ex.9 ***Calculate the number of g-molecules (mole of molecules) in the following : (i) 3.2 gm CH_4 (ii) 70 gm nitrogen (iii) 4.5 × 10^{24} molecules of ozone (iv) 2.4 × 10^{21} atoms of hydrogen (v) 11.2 L ideal gas at 0°C and 1 atm (vi) 4.54 ml SO_3 gas at STP (vii) 8.21 L C_2H_6 gas at 400K and 2 atm (viii) 164.2 ml He gas at 27°C and 570 torr [N_A = 6 × 10^{23}]***

Sol. (i) 3.2 gram CH_4

$$\text{number of moles } (CH_4) = \frac{w}{M} = \frac{3.2}{16} = 0.2 \text{ moles}$$

(ii) 70 gram N_2

$$\text{Number of moles} = \frac{w}{M} = \frac{70}{28} = 2.5$$

(iii) 4.5 × 10^{24} molecules of O_3

$$\text{Number of moles} = \frac{\text{no. of molecules}}{N_A} = \frac{4.5\times10^{24}}{6\times10^{23}} = 7.5$$

(iv) 2.4 × 10^{21} atoms of hydrogen

$$\text{Number of gram molecules of } H_2 = \frac{\text{no. of molecules}}{N_A} = \frac{2.4\times10^{21}}{2\times6\times10^{23}} = 0.002$$

(v) 11.2 litre ideal gas at 0°C and 1 atm

$$\text{Number of moles} = \frac{\text{Volume at } 0^oC \& 1 \text{ atm}}{22.4 \text{ litre}} = \frac{11.2}{22.4} = 0.5$$

(vi) 4.54 ml SO_3 gas at STP

$$\text{Number of moles} = \frac{V_{STP}(ml)}{22700 ml} = \frac{4.54}{22700} = 2 \times 10^{-4}$$

(vii) 8.21 litre C_2H_6 at 400 K and 2 litre

$$n = \frac{PV}{R.T} = \frac{2\times8.21}{0.0821\times400} = 0.5$$

(viii) 164.2 × ml He gas at 27ºC and 570 torr

$$n = \frac{PV}{RT} = \left(\frac{570}{760}atm\right) \times \frac{164.2 \times 10^{-3} litre}{0.0821 \times 300} = 0.005$$

Ex.10 ***Find no. of protons in 180 ml H_2O. Density of water = 1 gm/ml.***

Sol. Mass of water = density × volume = 180 g

$$\text{Moles of water} = \frac{180}{18} = 10$$

1 mol water has 10 mol protons
10 mol water has 100 mol protons
10 mol water has 100 N_A protons
10 mol water has 6.023×10^{25} protons

Ex.11 ***What mass of $Na_2SO_4.7H_2O$ contains exactly 6.023×10^{22} atoms of oxygen ?***

Sol. Molar mass of $Na_2SO_4.7H_2O$ = 275 gm.
1 mole $Na_2SO_4.7H_2O$ has 11 mol O-atoms.
$\Rightarrow$ 11 N_A O – atoms are in 275 g $Na_2SO_4 . 7H_2O$

$$\Rightarrow 6.023 \times 10^{22} \text{ O – atoms are in} = \frac{275}{11 \times 6.023 \times 10^{23}} \times 6.023 \times 10^{22} \text{ g} = 2.5 \text{ g}$$

Ex.12 ***What is number of atoms and molecules in 112 L of $O_3(g)$ at 0ºC and 1atm ?***

Sol. $$\text{Moles of molecules} = \frac{112}{22.4} = 5$$

Moles of atoms = 5 × 3 = 15
No. of molecules = 5 N_A
No. of atoms = 15N_A.

CLASS ILLUSTRATIONS (1)

BASIC

1. Find :

(i) No. of moles of Cu atom in 10^{20} atoms of Cu.

(ii) Mass of 200 $^{16}_{8}O$ atoms in amu

(iii) Mass of 100 atoms of $^{14}_{7}N$ in gm.

(iv) No. of molecules & atoms in 54 gm H_2O.

(v) No. of atoms in 88 gm CO_2.

2. Calculate mass of O atoms in 6 gm CH_3COOH ?

3. Calculate mass of water present in 499 gm $CuSO_4.5H_2O$?
(Atomic mass : Cu = 63.5, S = 32, O = 16, H = 1)

4. What mass of $Na_2SO_4.7H_2O$ contains exactly 6.022×10^{22} atoms of oxygen ?

5. The weight (in **gram**) of pure potash Alum ($K_2SO_4.Al_2(SO_4)_3.24H_2O$) which contains 0.64 kg oxygen is. (Atomic weight of K = 39, S = 32, Al = 27)

6. The Kohinoor diamond was the largest diamond ever found. How many moles of carbon atom were peresent in it, if it is weigh 3300 carat. **[Given: 1 carat = 200 mg]**

AVERAGE MOLAR MASS

7. The percentage by mole of NO_2 in a mixture of NO_2(g) and NO(g) having average molecular mass 34 is :

(A) 25% (B) 20% (C) 40% (D) 75%

8. The average atomic mass of a mixture containing 79 mole % of ^{24}Mg and remaining 21 mole % of ^{25}Mg and ^{26}Mg , is 24.31. **% mole** of ^{26}Mg is

(A) 5 (B) 20 (C) 10 (D) 15

□ **DENSITY :**

It is of two types.

I. Absolute density

II. Relative density

❖ **For liquids and solids :**

$$\text{Absolute density} = \frac{\text{mass}}{\text{volume}}$$

$$\textbf{Relative density or specific gravity} = \frac{\text{density of the substance}}{\text{density of water at } 4^\circ C\ (1\,gm\,ml^{-1})}$$

❖ **For gases :**

$$\text{Absolute density} = \frac{\text{mass}}{\text{volume}} = \frac{PM}{RT}$$

where P is pressure of gas, M = mol. wt. of gas, R is the gas constant, T is absolute temperature.

Vapour Density :

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

$$\text{Vapour density} = \frac{d_{gas}}{d_{H_2}} = \frac{PM_{gas}/RT}{PM_{H_2}/RT}$$

$$\text{V.D.} = \frac{M_{gas}}{M_{H_2}} = \frac{M_{gas}}{2} \Rightarrow \boxed{\mathbf{M_{gas} = 2 \times V.D.}}$$

Ex.13 ***A gaseous mixture of H_2 and NH_3 gas contains 68 mass % of NH_3. The vapour density of the mixture is –***

Sol. No. of moles of NH_3 in 100g mixture = $\frac{68}{17} = 4$

No. of moles of H_2 in 100g mixture = $\frac{32}{2} = 16$

$$M_{average} = \frac{\text{Total mass}}{\text{Total moles}} = \frac{100}{4+16} = 5$$

$$V.d = \frac{5}{2} = 2.5$$

STOICHIOMETRY

Stoichiometry is the calculation of amounts of reactants and products involved in a reaction. Stoichiometric calculations require a balanced chemical equation of the reaction.

• Remember a balanced chemical equation is one which contains an equal number of atoms of each element on both sides of equation.

□ SIGNIFICANCE OF STOICHIOMETRIC COEFFICIENTS :

Stoichiometric coefficients of chemical equation tells us about the ratio in which moles of reactants react and moles of products form.

Ex.	$2H_2(g)$ +	$O_2(g)$ ⟶	$2\ H_2O\ (g)$
1st interpretation	2 moles	1 mole	2 moles
2nd interpretation	$2\ N_A$ molecules	N_A molecules	$2\ N_A$ molecules
3rd interpretation	2 molecules	1 molecules	2 molecules

Ex.14 ***What mass of CaO is formed by heating 50 g $CaCO_3$ in air ?***

Sol. $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$

50 gm

$= \frac{50}{100}\text{mol}$

$= \frac{1}{2}\text{mol} \qquad \frac{1}{2}\text{mol}$

$= \frac{1}{2} \times 56 = 28\text{ gm}$

Ex.15 ***If 1 mole of ethanol (C_2H_5OH) completely burns to form carbon dioxide and water, mass of carbon dioxide formed is about***

Sol. $C_2H_5OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O$

1 3 2 2

2 mole of CO_2 are formed = 88g

Ex.16. ***What volume of CO_2 at 0ºC and 1 atm is formed by heating 200 g $CaCO_3$?***

Sol. $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$

200 gm

$= \frac{200}{100}\text{mol} \qquad = 2\text{ mol} \qquad 2\text{ mol}$

Volume of gas at 0ºC and 1 atm = No. of moles × 22.4 L = 2 × 22.4 = 44.8 L.

□ LIMITING REAGENT (L.R.) :

(i) The reactant which is completely consumed when a reaction goes to completion is called Limiting Reactant or Limiting reagent.

(ii) The reactant whose Stoichiometric amount is least, is limiting reactant.

$$\text{Where ; Stoichiometric amount} = \frac{\text{Given moles of reactant}}{\text{Stoichiometric coefficient of reactant in balance Reaction}}$$

(iii) For calculation of moles of product, LR should be used. When amounts of two or more than two reactants are given :

	aA	+	bB	$\longrightarrow$	cC	+	dD
Initial reacting mixture	n_A mol		n_B mol				
Stoichiometric amount	$\frac{n_A}{a}$		$\frac{n_B}{b}$				

If $\frac{n_A}{a} < \frac{n_B}{b} \Rightarrow$ A is limiting reagent.

If $\frac{n_A}{n_B} = \frac{a}{b}$ then reaction occurs to completion & no reactant is left at the end.

If $\frac{n_A}{a} > \frac{n_B}{b} \Rightarrow$ B is limiting reagent.

Ex.17. ***28 gm Lithium is mixed with 48 gm O_2 to reacts according to the following reaction.***
$4Li + O_2 \longrightarrow 2Li_2O$
The mass of Li_2O formed is

Sol.

	$4Li$	+	O_2	$\longrightarrow$	$2Li_2O$
moles taken	$\frac{28}{7}$		$\frac{48}{32}$		
	$=4$		$= 1.5$		
$\frac{\text{moles taken}}{\text{stoich. coeff.}}$	$\frac{4}{4} = 1$		$\frac{1.5}{1} = 1.5$		
	(L.R.)				

Moles of Li_2O formed $= \frac{2}{4} \times 4 = 2$

Mass of Li_2O formed $= 2 \times 30 = 60$ gm

Ex.18 ***Calculate the mass of sucrose $C_{12}H_{22}O_{11}$ (s) produced by mixing 78 g of C(s), 11 g of H_2(g) & 67.2 litre of O_2 (g) at 0°C and 1 atm according to given reaction (unbalanced) ?***

Sol. $12C(s) + 11\ H_2(g) + \frac{11}{2}O_2 \rightarrow C_{12}H_{22}O_{11}\ (s)$

Moles taken	$\frac{78}{12}$	$\frac{11}{2}$	$\frac{67.2}{22.4}$
	$= 6.5$	$= 5.5$	$= 3$
$\frac{\text{moles taken}}{\text{stoich. coeff.}}$	$\frac{6.5}{12}$	$\frac{5.5}{11}$	$\frac{3}{5.5}$
	$= 0.54$	$= 0.5$ (L.R.)	$= 0.545$

$\therefore$ Moles of $C_{12}H_{22}O_{11}$ formed $= \frac{5.5}{11} = 0.5$

Mass of sucrose obtained $= 0.5 \times 342 = 171$ grams.

CLASS ILLUSTRATION-(2)

STOICHOMETRY

9. How many gm of HCl is needed for complete reaction with 43.5 gm MnO_2 ? (Mn = 55)

$HCl + MnO_2 \longrightarrow MnCl_2 + H_2O + Cl_2$

10. Nitric acid is manufactured by the Ostwald process, in which nitrogen dioxide reacts with water.

$3\ NO_2\ (g) + H_2O\ (l) \rightarrow 2\ HNO_3\ (aq) + NO\ (g)$

How many grams of nitrogen dioxide are required in this reaction to produce 25.2 gm HNO_3 ?

LIMITING REAGENT

11 Carbon reacts with chlorine to form CCl_4. 36 gm of carbon was mixed with 142 g of Cl_2. Calculate mass of CCl_4 produced and the remaining mass of reactant.

12. A chemist wants to prepare diborane by the reaction

$6\ LiH + 8BF_3 \longrightarrow 6Li\ BF_4 + B_2H_6$

If he starts with 2.0 moles each of LiH & BF_3. How many moles of B_2H_6 can be prepared.

□ PROBLEMS BASED ON MIXTURE :

The composition of any mixture may be determined by reacting the mixture with some substance, by which either one or more component of mixture may react.

Ex.19 ***1.5 gm mixture of SiO_2 and Fe_2O_3 on very strong heating leave a residue weighting 1.46 gm. The reaction responsible for loss of weight is***

$$Fe_2O_3\ (s) \rightarrow Fe_3O_4\ (s) + O_2(g)$$

What is the percentage by mass of Fe_2O_3 in original sample.

Sol. $3Fe_2O_3\ (s) \rightarrow 2Fe_3O_4 + \frac{1}{2}O_2$

$3 \times 160 \qquad \frac{1}{2} \times 32$

$= 480\ gm \rightarrow \qquad = 16\ gm$

loss of 16 gm $\rightarrow$ 480 gm Fe_2O_3

loss of 0.04 gm $\rightarrow 0.04 \times \frac{480}{16} = 1.2$ gm Fe_2O_3

% by mass $= \frac{1.2}{1.5} \times 100 = 80\%$

CLASS ILLUSTRATION- (3)

MIXTURE

13. A sample of mixture of $CaCl_2$ and NaCl weighing 2.22 gm was treated to precipitate all the Ca as $CaCO_3$ which was then heated and quantitatively converted to 0.84 gm of CaO. Calculate the percentage (by mass) of $CaCl_2$ in the mixture.

14. When 4 gm of a mixture of $NaHCO_3$ and NaCl is heated, 0.66 gm CO_2 gas is evolved. Determine the percentage composition (by mass) of the original mixture.

□ PERCENTAGE 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 theorectical yield the percentage yield can be calculate as :

$$\boxed{\%\text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100}$$

The percentage yield of any product is always equal to the percentage extent of that reaction.

Ex.20 ***Aluminium reacts with sulphur to form aluminium sulphide. If 5.4 gm of Aluminium reacts with 12.8gm sulphure gives 12gm of aluminium sulphides, then the percent yield of the reaction is-***

Sol

	$2Al$	$+$	$3S \longrightarrow$	Al_2S_3
Mole taken	$\frac{5.4}{27}$ gm		$\frac{12.8}{32}$ gm	
	$= 0.2$		$= 0.4$	
$\frac{\text{moles taken}}{\text{stoich. coeff.}}$	$\frac{0.2}{2}$		$\frac{0.4}{3}$	
	$= 0.1$ (L.R.)		$= 0.133$	

$$\text{moles of } Al_2S_3 \text{ formed} = \frac{1}{2} \times 0.2 = 0.1$$

$$\text{mass} = 0.1 \times 150 = 15 \text{ gm}$$

$$\%\text{ yield} = \frac{\text{actual yield}}{\text{theoritical yield}} \times 100 = \frac{12}{15} \times 100 = 80\ \%$$

□ DEGREE OF DISSOCIATION, α :

It represents the mole of substance dissociated per mole of the substance taken.

$$A \rightarrow n \text{ particles; } \alpha = \frac{M_o - M}{(n-1).M}$$

where, n = number of product particles per particle of reactant

M_o = Molar mass of 'A'

M = Molar mass of final mixture

Dissociation decreases the average molar mass of system while association increases it.

Ex.21 ***For the reaction $2NH_3(g) \rightarrow N_2(g) + 3H_2(g)$***

Calculate degree of dissociation (α) if observed molar mass of mixture is 13.6

Sol. $\alpha = \frac{M_T - M_0}{(n-1)M_0} = \frac{17 - 13.6}{(2-1) \times 13.6} = 0.25$

▫ PERCENTAGE PURITY :

The percentage of a specified compound or element in an impure sample may be given as

$$\% \text{ purity} = \frac{\text{Actual mass of compound}}{\text{Total mass of sample}} \times 100$$

If impurity is unknown, it is always considered as inert (unreactive) material.

Ex.22 ***A chalk sample exactly requires 17.52 gram HCl for complete reaction with all $CaCO_3$ present in it. If the chalk sample is 72% pure, the mass of sample taken is***

Sol. $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

Moles of HCl = $\frac{17.52}{36.5}$

Moles of $CaCO_3$ = $\frac{1}{2} \times \frac{17.52}{36.5}$

Weight of $CaCO_3$ required = $\frac{1}{2} \times \frac{17.52}{36.5} \times 100$

Mass of sample taken :

$= \frac{1}{2} \times \frac{17.52}{36.5} \times \frac{100 \times 100}{72} = 33.33$ gm

CLASS ILLUSTRATION- (4)

PERCENTAGE YIELD , PERCENTAGE PURITY

15. A power company burns approximately 500 tons of coal per day to produce electricity. If the sulphur content of the coal is 1.20 % by weight, how many tons SO_2 are dumped into the atmosphere each day ?

16. Cyclohexanol is dehydrated to cyclohexene on heating with conc. H_2SO_4. If the yield of this reaction is 75%, how much cyclohexene will be obtained from 100 g of cyclohexanol ?

$$C_6H_{12}O \xrightarrow{\text{con. } H_2SO_4} C_6H_{10}$$

17. If the yield of chloroform obtainable from acetone and bleaching powder is 58%. What is the weight of acetone required for producing 23.9 gm of chloroform ?

$$2CH_3COCH_3 + 6CaOCl_2 \rightarrow Ca(CH_3COO)_2 + 2CHCl_3 + 3CaCl_2 + 2Ca(OH)_2$$

18. Calculate % yield of the reaction if 200g $KHCO_3$ produces 22g of CO_2 upon strong heating.

▫ PROBLEMS RELATED WITH SEQUENTIAL REACTION :

When one of products formed in previous reaction is consumed in the next one.

Ex.23 ***How many grams H_2SO_4 can be obtained from 1320 gm PbS as per reaction sequence ?***

$$2PbS + O_2 \longrightarrow 2PbO + 2SO_2$$
$$3SO_2 + 2HNO_3 + 2H_2O \longrightarrow 3H_2SO_4 + 2NO$$

[At. mass : Pb = 208, S = 32]

Sol. Moles of PbS = $\frac{1320}{240}$ = 5.5 mol

Moles of SO_2 = 5.5 mol = moles of H_2SO_4

Mass of H_2SO_4 = 5.5 × 98 = 539 gm

[When amount of only one reactant is given generally other is assumed in excess.]

▫ PROBLEM RELATED WITH PARALLEL REACTION :

When same two reactants form two or more products by independent reactions.

Ex.24 ***Carbon reacts with oxygen forming carbon monoxide and/or carbon dioxide depending an availablity of oxygen. Find moles of each product obtained when 160 gm oxygen reacts with (a) 12 g carbon (b) 120 g carbon (c) 72 g carbon.***

Sol. **(a)** $C + \frac{1}{2}O_2 \longrightarrow CO$ [initially use a reaction using lesser amount of oxygen]

t = 0	1mol	5mol	
t =∞	0	5–0.5 =	1mol
	(LR)	4.5mol	

Since CO & O_2 are left CO_2 is formed.

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

t = 0	1mol	4.5mol	0
t = ∞	0	4 mol	1 mol

At end 1 mole CO_2 & no CO present

(b) $C + \frac{1}{2}O_2 \longrightarrow CO$

t = 0	10mol	5mol	0
t =∞	0	0	10mol

At end only 10 mol CO present.

(c) $C + \frac{1}{2}O_2 \longrightarrow CO$

t = 0	6mol	5mol	0
t =∞	0	2mol	6mol
	[LR]		

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

t = 0	6mol	2mol	0
t =∞	2mol	0 [LR]	4mol

At end [2mol CO + 4mol CO_2] left.

Ex.25 ***25.4 gm of iodine and 14.2 gm of chlorine are made to react completely to yield mixture of ICl and ICl_3. Ratio of moles of ICl & ICl_3 formed is (Atomic mass : I = 127, Cl = 35.5)***

Sol.

I_2 + Cl_2 $\longrightarrow$ ICl + ICl_3

0.1mol 0.2 mol x mol y mol

$\Rightarrow$ 0.2=x+y $\Rightarrow$ x=0.1

0.4=x+3y $\Rightarrow$ y=0.1

Ans. 1 : 1

PRINCIPLE OF ATOM CONSERVATION (POAC)

POAC is nothing but the conservation of atoms of reactants and products involved in a chemical reaction. And if atoms are conserved, moles of atoms shall also be conserved. The principle is fruitful for the students when they don't get the idea of balanced chemical equation in the problem using POAC we do not need to balance a reaction and we can even add two or more reactions. This principle can be understood by the following example.

Consider the decomposition of $KClO_3$ (s) $\rightarrow$ KCl (s) + O_2 (g) (unbalanced chemical reaction)

Apply the principle of atom conservation (POAC) for K atoms.

or moles of K atoms in $KClO_3$ = moles of K atoms in KCl

Now, since 1 molecule of $KClO_3$ contains 1 atom of K

Thus, moles of K atoms in $KClO_3$ = 1 $\times$ moles of $KClO_3$

and moles of K atoms in KCl = 1 $\times$ moles of KCl

$\therefore$ moles of $KClO_3$ = moles of KCl or $\dfrac{\text{wt. of } KClO_3 \text{ in g}}{\text{mol. wt. of } KClO_3} = \dfrac{\text{wt. of KCl in g}}{\text{mol. wt. of KCl}}$

- The above equation gives the mass-mass relationship between $KClO_3$ and KCl which is important in stoichiometric calculations.Again, applying the principle of atom conservation for O atoms,

moles of O in $KClO_3$ = 3 $\times$ moles of $KClO_3$

moles of O in O_2 = 2 $\times$ moles of O_2

$\therefore$ 3 $\times$ moles of $KClO_3$ = 2 $\times$ moles of O_2

or $3 \times \dfrac{\text{wt. of } KClO_3}{\text{mol. wt. of } KClO_3} = 2 \times \dfrac{\text{vol. of } O_2 \text{ at 1atm and 0°C}}{\text{Molar vol. (22.4 lt)}}$

- The above equations thus gives the mass-volume relationship of reactants and products.

CLASS ILLUSTRATION- (5)

19. Calcium phosphide Ca_3P_2 formed by reacting magnesium with excess calcium orthophosphate $Ca_3(PO_4)_2$, was hydrolysed by excess water. The evolved phosphine PH_3 was burnt in air to yield phosphrous pentoxide (P_2O_5). How many gram of magnesium metaphosphate would be obtain if 192 gram Mg were used (Atomic weight of Mg = 24, P = 31)

$$Ca_3(PO_4)_2 + Mg \longrightarrow Ca_3P_2 + MgO$$

$$Ca_3P_2 + H_2O \longrightarrow Ca(OH)_2 + PH_3$$

$$PH_3 + O_2 \longrightarrow P_2O_5 + H_2O$$

$$MgO + P_2O_5 \longrightarrow Mg(PO_3)_2$$

20. 27.6 g 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.
[mol. wt. of K_2CO_3 = 138 and mol. wt. of $K_2Zn_3[Fe(CN)_6]_2$ = 698]

PERCENTAGE DETERMINATION OF ELEMENTS IN ORGANIC COMPOUNDS :

All these methods are applications of POAC

Do not remember the formulas, derive them using the concept, its easy.

(a) **Liebig's method :** (*for Carbon and hydrogen*)

$$(w)\text{ Organic Compound} \xrightarrow[CuO]{\Delta} (w_1)CO_2 + H_2O(w_2)$$

$$\% \text{ of C} = \frac{w_1}{44}\times\frac{12}{w}\times 100$$

$$\% \text{ of H} = \frac{w_2}{18}\times\frac{2}{w}\times 100$$

where w_1 = wt. of CO_2 produced, w_2 = wt. of H_2O produced,
w = wt. of organic compound taken

(b) **Duma's method :** (*for nitrogen*)

$$(w)\text{ Organic Compound} \xrightarrow[CuO]{\Delta} N_2 \rightarrow (P, V, T \text{ given})$$

use PV = nRT to calculate moles of N_2, n.

$$\therefore \% \text{ of N} = \frac{n\times 28}{w}\times 100$$

(c) **Kjeldahl's method :** (*for nitrogen*)

$$(w)O.C. + H_2SO_4 \rightarrow (NH_4)_2SO_4 \xrightarrow{NaOH} NH_3 + H_2SO_4 \rightarrow$$ (molarity and volume (V litre) consumed given)

$$\Rightarrow \% \text{ of N} = \frac{MV\times 2\times 14}{w}\times 100$$

where M = molarity of H_2SO_4. Some N containing compounds do not give the above set of reaction as in Kjeldahl's method.

(d) **Sulphur :**

$$(w)\text{ O.C.} + HNO_3 \rightarrow H_2SO_4 + BaCl_2 \rightarrow (w_1)\ BaSO_4$$

$$\Rightarrow \quad \% \text{ of S} = \frac{w_1}{233}\times\frac{1\times 32}{w}\times 100\%.$$

where w_1 = wt. of $BaSO_4$, w = wt. of organic compound

(e) **Phosphorus :**

$$O.C + HNO_3 \rightarrow H_3PO_4 + [NH_3 + \text{magnesia mixture ammonium molybdate}] \rightarrow$$

$$MgNH_4PO_4 \xrightarrow{\Delta} Mg_2P_2O_7$$

$$\% \text{ of P} = \frac{w_1}{222}\times\frac{2\times 31}{w}\times 100$$

(f) **Carius method :** (*Halogens*)

O.C. + HNO_3 + $AgNO_3$ $\rightarrow$ AgX

If X is Cl then colour = white

If X is Br then colour = dull yellow

If X is I then colour = bright yellow

Flourine can't be estimated by this

$$\% \text{ of X} = \frac{w_1}{(\text{M. weight of AgX})} \times \frac{1 \times (\text{At.wt.of X})}{w} \times 100$$

Ex.26 ***A sample of 0.5 gm of an organic compound was treated according to Kjeldahl's method. The ammonia evolved was absorbed by 2.45 gm of H_2SO_4. The residual acid required solution containing 0.6 gm. NaOH for neutralisation. Find the percentage composition of nitrogen in the compound ?***

Sol. 2 NH_3 + H_2SO_4 $\rightarrow$ $(NH_4)_2$ SO_4

H_2SO_4 + 2NaOH $\rightarrow$ Na_2SO_4 + $2H_2O$

m mol of H_2SO_4 used to react with NaOH $= \frac{0.6}{40} = 15$mmol.

Remaining mmol of H_2SO_4 = $\frac{2.45}{98} \times 10^3 - 15 = 10$

mmol of NH_3 used = 10 × 2 = 20

% N in sample = $\frac{20 \times 10^{-3} \times 14}{0.5} \times 100 = 56\%$

Ex.27 ***Calculate the molar mass of a compound in the Dumas method at 100ºC for which volume of experimental container was 452 ml and the pressure was 745.1 torr. The difference in mass between the empty container and the final measurement was 1.129 gm.***

Sol. n = $\frac{PV}{RT} = \frac{745.1}{760} \times \frac{452 \times 10}{0.0821 \times 373}$ = 0.01448 mol

molar mass (M) = $\frac{1.129}{0.01448}$ = 78.0 gm/mol.

EMPIRICAL AND MOLECULAR FORMULA

We have just seen that knowing the molecular formula of the compound we can calculate percentage composition of the elements. Conversely if we know the percentage composition of the elements initially, we can calculate the relative number of atoms of each element in the molecules of the compound. This gives us the empirical formula of the compound. Further if the molecular mass is known then the molecular formula can be easily determined.

Thus, the empirical formula of a compound is a chemical formula showing the relative number of atoms in the simplest ratio, the molecular formula gives the actual number of atoms of each element in a molecule.

i.e. **Empirical formula :** Formula depicting constituent atoms in their simplest ratio.

Molecular formula : Formula depicting actual number of atoms in one molecule of the compound.

The molecular formula is generally an integral multiple of the empirical formula.

i.e. molecular formula = empirical formula × n

where $n = \frac{\text{molecular formula mass}}{\text{empirical formula mass}}$

Example :

Molecular Formula	H_2O_2	C_6H_6	C_2H_6	$C_2H_4O_2$
	2 : 2	6 : 6	2 : 6	2 : 4 : 2
Simplest ratio	1 : 1	1 : 1	1 : 3	1 : 2 : 1
Empirical Formula	H O	C H	CH_3	CH_2O

□ DETERMINATION OF EMPIRICAL FORMULA :

Following steps are involved in determining the empirical formula of the compounds –

(i) First of all find the % by wt. of each element present in the compound.

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

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

(iv) If the value of simplest atomic ratio is fractional then raise the value to the nearest whole number or multiply with suitable coefficient to convert it into nearest whole number

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

□ DETERMINATION OF MOLECULAR FORMULA :

(i) Find out the empirical formula mass by adding the atomic masses of all the atoms present in the empirical formula of compound.

(ii) Divide the molecular mass (determined experimentally by some suitable method) by the empirical formula mass and find out the value of n.

(iii) Multiply the empirical formula of the compound with n so as to find out the molecular formula of the compound.

Ex.28. ***An organic compound contains 49.3% carbon, 6.84% hydrogen and its vapour density is 73. Molecular formula of compound is :-***

Sol. V.D. = 73 $\Rightarrow$ M = 2 × 73 = 146

$C = 146 \times \frac{49.3}{100} = 71.978 \text{ g} \simeq 6 \text{ mole}$

$H = 146 \times \frac{6.84}{100} = 9.9864 \text{ g} \simeq 10 \text{ mole}$

$O = 146 \times \frac{43.86}{100} = 64.86 \text{ g} \approx 4 \text{ mol}$

M.F. = $C_6H_{10}O_4$

Ex.29 ***The empirical formula of an organic compound containing carbon & hydrogen is CH_2. The mass of 1 litre of organic gas is exactly equal to mass of 1 litre N_2 therefore molecular formula of organic gas is.***

Sol. Empirical Mass of CH_2 = 12 + 2 = 14

$\because$ Mass of 1 litre of organic gas = Mass of 1 litre of N_2

Since V, P, T, n are same.

Therefore $PV = \frac{m}{M}RT$

implies that molar mass should also be same.

$\therefore$ Molecular mass of organic compound will be 28 g

$$n = \frac{\text{Molecular mass}}{\text{Empirical mass}} = \frac{28}{14} = 2$$

So molecular formula $= 2 \times CH_2 = C_2H_4$

CLASS ILLUSTRATION- (6)

PERCENTAGE COMPOSITION, EMPERICAL AND MOLECULAR FORMULA

21. A moth repellent has the composition 49% C, 2.7% H and 48.3% Cl. Its molecular weight is 147 gm. Determine its molecular formula

22. The empirical formula of a compounds is CH_2O. 0.25 mole of this compound contains 1 gm hydrogen. The molecular formula of compound is -

23. A compound has 62 % carbon, 10.4 % hydrogen and 27.5 % oxygen. If molar mass of compound is 58, find number of H-atoms per molecule of the compound.

EXPERIMENTAL METHODS TO DETERMINE ATOMIC & MOLECULAR MASSES

I. For determination of atomic mass :

Dulong's & Petit's law :

In case of solid elements, it is observed that product of atomic weight and specific heat capacity is almost constant.

Atomic weight of metal × specific heat capacity (cal/gm°C) ≈ 6.4.

It should be remembered that this law is an empirical observation and this gives an approximate value of atomic weight. This law gives better result for heavier elements, at high temperature conditions.

Ex.30 ***The product of atomic mass (gm/mol) and specfic heat (cal/K-gm) of elements is approximately 6.4, except***

(A) Pt (B) Au (C) Pb (D) Ne

Ans. (D)

II. Experimental methods for molecular mass determination.

(*a*) Victor Meyer's Method

(*b*) Silver Salt Method

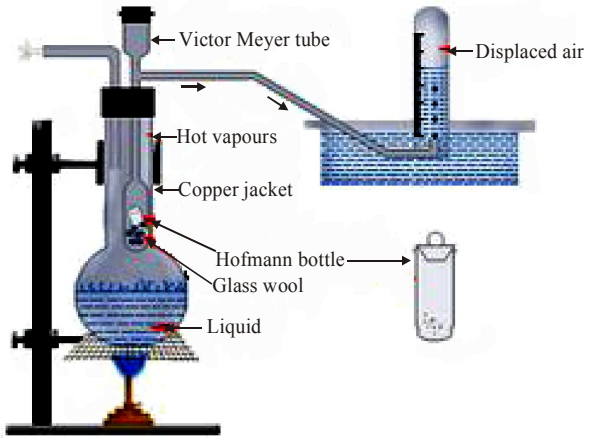
(*c*) Chloroplatinate Salt Method

(a) Victor Meyer's Method : (Applicable for volatile substance)

A known mass of the volatile substance taken in the Hoffmann's bottle and is vapourised by throwing the Hoffmann's bottle into the Vector Meyer's tube. The vapour displace an equal volume of the air. which is measured at the room temperature and atmospheric presssure. The barometric pressure and the room temperature is recorded. Following diagram gives the experimental set-up for the Victor-Meyer's process.

Calculation involved

Let the mass of the substance taken by = Wg
Volume of moist air collected = Vcm^3
Room temperature = TK
Barometric pressure = P mm Hg
Aqueous tension at TK = p mm Hg
Pressure of dry air = (P – p) mm Hg
Calculation of molecular mass (M)



$$\frac{(P-p)}{760}\times\frac{V}{1000}=\frac{w}{M}\times RT$$

$$\Rightarrow \quad M=\frac{w\times RT\times 760\times 1000}{(P-p)\times V}$$

Applying PV = nRT for the dry vapour and using n = w/M

Ex.31 ***0.15 g of substance displaced 58.9 cm^3 of air at 300 K and 746 mm pressure. Calculate the molecular mass. (Aq. Tension at 300K = 26.7 mm).***

Sol. Mass of substance = 0.15 g

Volume of Air displaced (V) = 58.9 cm^3

Temperature (T) = 300 K

Pressure (P) = 746 – 26.7 = 719.3 mm

$$\text{Molecular mass}=\frac{719.3}{760}\times\frac{58.9}{1000}=\frac{0.15}{M}\times.0821\times 300$$

$\therefore$ Molecular mass = 66.24 g/mol.

(b) Silver salt Method : (Used for organic acids)

A known mass of the acid is dissolved in water followed by the subsequent addition of silver nitrate solution till the preceipitation of silver salt is complete. The precipitate is separated. dried, weighed and ignited till decomposition is complete. The residue of pure silver left behind is weighed.

$$\text{Organic acid} \xrightarrow{AgNO_3} \text{Silver salt} \xrightarrow{\text{Ignite}} Ag$$

Calculations involved
Let the mass of the silver salt formed = Wgm
The mass of Ag formed = x gm
For polybasic acid of the type H_nX (n is basicity)

$$\underset{\text{Organic acid}}{H_nA} \xrightarrow{AgNO_3} \underset{\text{Silver salt (Wg)}}{Ag_nA} \xrightarrow{\text{Ignite}} \underset{\text{Silver (xg)}}{nAg}$$

Mass of the silver salt that gives x gm of Ag = W gm

$$\text{Mass of the silver salt that gives ng (108 g) of Ag} = \frac{108nW}{x}\,g$$

$$\text{Molar mass of salt} = \frac{108\times nW}{x}g$$

$$\textbf{Molar mass of acid} = \frac{108\times nW}{x}-n\times 108+n\times 1=n\left(\frac{108W}{x}-107\right)gmol^{-1}$$

Ex.32 ***0.41g of the silver salt of a dibasic organic acid left a residue to 0.216g of silver on ignition. Calculate the molecular mass of the acid.***

Sol. Mass of of the silver salt taken (W) = 0.41 g , Mass of Ag formed = 0.216g

$$H_2X \longrightarrow \underset{w=0.41g}{Ag_2X} \longrightarrow \underset{x=0.216g}{2Ag}$$

$$\text{Now molar mass acid} = n\left(\frac{108W}{x} - 107\right) gmol^{-1} = 2\left(\frac{108\times 0.41}{0.216} - 107\right) gmol^{-1} = 196 gmol^{-1}$$

Molar mass = 196 g/mol

(c) Platinic chloride Method : (Applicable for finding the molecular mass of organic bases).

A known mass of organic base is allowed to react with chloroplatinic acid (H_2PtCl_6) in conc. HCl to form insoluble platinic chloride. The precipitate of platinic chloride is separated, dried, weighed and is subsequently ignited till decomposition is complete. The residue left is platinum which is again weighed. The molecular mass is then calculated by knowing the mass of the platinic chloride salt and that of platinum left.

If B represents the molecule of monoacidic organic base, then the formula of platinic chloride salt is $B_2H_2PtCl_6$.

$$\underset{\text{Organic base}}{B} \xrightarrow[\text{conc. HCl}]{H_2PtCl_6} \underset{\text{Platinic chloride salt (Wg)}}{B_2H_2PtCl_6} \xrightarrow{\text{Ignite}} \underset{(xg)}{Pt}$$

It may be noted that salt formed with diacidic base would be $B_2(H_2PtCl_6)_2$: with triacidic base would be $B_2(H_2PtCl_6)_3$ and with polyacidic base would be $B_2(H_2PtCl_6)_n$.

Now from the formula $B_2(H_2PtCl_6)$

Molar mass of salt = (2 × molar mass of base) + (Molar mass of H_2PtCl_6)

Molar mass of base = $\frac{1}{2}$ (Molar mass of salt – Molar mass of H_2PtCl_6)

$$= \frac{1}{2}\left(\frac{W\times 195\times n}{x} - n\times 410\right) = \frac{n}{2}\left(\frac{w\times 195}{x} - 410\right) gmol^{-1}$$

Ex.33 ***0.30 gm chloroplatinate salt of a diacidic organic base exactly produce 0.09 gm platinum, on strong ignition. The molecular mass of organic base is (Pt = 195)***

Sol. Molar mass of base is

$$= \frac{n}{2}\left(\frac{w\times 195}{x} - 410\right)$$

$$= \frac{2}{2}\left(\frac{0.3\times 195}{0.09} - 410\right) = 240 \text{ gm/ mol}$$

ANSWEERS

CLASS ILLUSTRATION-(1)

1. (i) $\frac{10^{20}}{N_A}$ moles (ii) 3200 amu (iii) $14 \times 1.66 \times 10^{-24} \times 100$ g

   (iv) $3N_A$, $9N_A$ (v) $6N_A$

2. 3.2 g 3. 180 g 4. 2.436 g 5. Ans. (948)

6. Ans(55) 7. Ans.(A) 8. Ans.(C)

CLASS ILLUSTRATION-(2)

9. 73 gm 10. 27.6 gm 11 154 gm, 24 gm 12. 0.25 mole

CLASS ILLUSTRATION- (3)

13. 75% 14. 63 % , 37%

CLASS ILLUSTRATION- (4)

15 Ans.12 16. Ans. 61.5 gm 17 Ans.20 gm 18. Ans. (050)

CLASS ILLUSTRATION- (5)

19. Ans.1456 gm. 20. Ans. 11.6 g

CLASS ILLUSTRATION- (6)

21. $C_6H_4Cl_2$ 22. Ans. $(C_2H_4O_2)$ 23. Ans. (6)

UNIT-02

1. SOLUTIONS CONCENTRATION TERMS

A solution is a homogenous mixture of two or more pure substances whose composition may be altered within certain limits. Though the solution is homogenous in nature, yet it retains the properties of its constituents.

Generally solution is composed of two components, solute and solvent. Such type of solution is known as binary solutions.

Solvent is that component in solution whose physical state is the same as that of the resulting solution while other component is called as solute. If the physical state of both component is same, than the component in excess is known as solvent and other one is called as solute. Each component in a binary solution can be in any physical state such as liquid, solid and gaseous state.

Table 2.1: Types of Solutions

Type of Solutions	Solute	Solvent	Common Example
Gaseous Solutions	Gas	Gas	Mixture of oxygen and nitrogen gases
	Liquid	Gas	Chloroform mixed with nitrogen gas
	Solid	Gas	Camphor in nitrogen gas
Liquid Solutions	Gas	Liquid	Oxygen dissolved in water
	Liquid	Liquid	Ethanol dissolved in water
	Solid	Liquid	Glucose dissolved in water
Solid Solutions	Gas	Solid	Solution of hydrogen in palladium
	Liquid	Solid	Amalgam of mercury with sodium
	Solid	Solid	Copper dissolved in gold

2. CONCENTRATION TERMS :

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

2.1 Mass percentage :

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

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

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

2.2 Mass-volume percentage (W/V %) :

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

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

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

2.3 Volume Percent :

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

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

2.4 $\textbf{Mole \%} = \frac{\text{Moles of solute}}{\text{Total moles}} \times 100$

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

2.5 Mole Fraction (X) :

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

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

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

2.6 Molarity (M) :

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

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

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

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

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

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

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

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

Concentration Type	Mathematical Formula	Concept
Percentage by mass	$\%\left(\frac{w}{w}\right) = \frac{\text{Mass of solute} \times 100}{\text{Mass of solution}}$	Mass of solute present in 100 gm of solution.
Volume percentage	$\%\left(\frac{v}{v}\right) = \frac{\text{Volume of solute} \times 100}{\text{Volume of solution}}$	Volume of solute present in 100 cm^3 of solution.
Mass-volume percentage	$\%\left(\frac{w}{v}\right) = \frac{\text{Mass of solute} \times 100}{\text{Volume of solution}}$	Mass of solute present in 100 cm^3 of solution.
Parts per million	$\text{ppm} = \frac{\text{Mass of solute} \times 10^6}{\text{Mass of solution}}$	Parts by mass of solute per million parts by mass of the solution
Mole fraction	$X_A = \frac{\text{Mole of A}}{\text{Mole of A} + \text{Mole of B} + \text{Mole of C} + ...}$ $X_B = \frac{\text{Mole of B}}{\text{Mole of A} + \text{Mole of B} + \text{Mole of C} + ...}$	Ratio of number of moles of one component to the total number of moles.
Molarity	$M = \frac{\text{Mole of solute}}{\text{Volume of solution (in L)}}$	Moles of solute in one liter of solution.
Molality	$m = \frac{\text{Mass of solute} \times 1000}{\text{Molar mass of solute} \times \text{Mass of solvent(g)}}$	Moles of solute in one kg of solvent

Ex.1 ***Calculate the mole fractions of the components of the solution composed by 92 g glycerol and 90 g water ? (M (water) = 18 ; M (glycerol) = 92)***

Ans. Moles of water = 90 g / 18 g = 5 mol water
Moles of glycerol = 92 g / 92 g = 1 mol glycerol
Total moles in solution = 5 + 1 = 6 mol
Mole fraction of water = 5 mol / 6 mol = 0.833
Mole fraction of glycerol = 1 mol / 6 mol = 0.167

Ex.2 ***What will be the Molarity of solution when water is added to 10 g $CaCO_3$ to make 100 mL of solution?***

Ans. Mol of $CaCO_3$ = 10 / 100 = 0.1
Molarity = Mole of solute / Volume of solution (L) = 0.10 mol / 0.10 L
Therefore ; Molarity of given solution = 1.0 M

Ex.3 ***Calculate the molality of a solution containing 20 g of sodium hydroxide (NaOH) in 250 g of water?***

Ans. Moles of sodium hydroxide = 20 / 40 = 0.2 mol NaOH
250 gm = 0.25 kg of water
Hence molality of solution = Mole of solute / Mass of solvent (kg)= 0.2 mol / 0.25 kg
or Molality(m) = 0.8 mol / kg or 0.8 m

Ex.4 ***Calculate the grams of copper sulphate ($CuSO_4$) needed to prepare 250.0 mL of 1.00 M $CuSO_4$?***

Ans. Moles of $CuSO_4 = M \times V = 1 \times \frac{250}{1000}$
Molar mass of copper sulphate = 159.6 g/mol
Hence Mass of copper sulphate (gm) = Moles of $CuSO_4$ × Molar mass of copper sulphate.

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

= 39.9 gm of Copper sulphate

Ex.5 ***How many grams of H_2SO_4 are present in 500 ml of 0.2M H_2SO_4 solution ?***

Ans. $M = \frac{\text{moles}}{\text{vol.}} \Rightarrow \text{moles of } H_2SO_4 = M \times V = 0.2 \times \frac{500}{1000} L = 0.1$
Mass of $H_2SO_4 = 0.1 \times 98 = 9.8$ g

Ex.6 ***Calculate the ppm of mercury in water in given sample contain 30 mg of Hg in 500 ml of solution.***

Ans. $\text{Parts per million} = \frac{\text{Mass of solute} \times 10^6}{\text{Mass of solution}}$
Mass of Hg = 30 mg
Mass of water = 500/1 = 500g = 50×10^4 mg
(density = mass / volume ; density of water 1 g / ml) $w = \frac{v}{d}$
Therefore, ppm of mercury = $\frac{30 \times 10^6}{50 \times 10^4}$ = 60 ppm of mercury

CLASS ILLUSTRATION- (1)

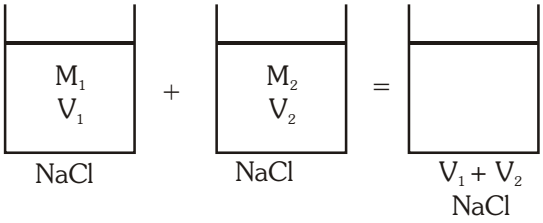
1. Calculate the molarity of the following solutions :
   (a) 4g of caustic soda is dissolved in 200 mL of the solution.
   (b) 5.3 g of anhydrous sodium carbonate is dissolved in 100 mL of solution.
   (c) 0.365 g of pure HCl gas is dissolved in 50 mL of solution.
2. The density of a solution containing 7.3% by mass of HCl is 1.2 g/mL. Calculate the molarity of the solution.
3. 15 g of methyl alcohol is present in 100 mL of solution. If density of solution is 0.90 g mL^{-1}. Calculate the mass percentage of methyl alcohol in solution
4. What is the concentration of chloride ion, in molarity, in a solution containing 10.56 gm $BaCl_2.8H_2O$ per litre of solution ? (Ba = 137)
5. The mole fraction of solute in aqueous urea solution is 0.2. Calculate the mass percent of solute ?
6. A solution has 80% $\frac{w}{w}$ NaOH with density 2gL^{-1}. Find (a) Molarity (b) Molality of solution.
7. 4.450 g 100 per cent sulphuric acid was added to 82.20 g water and the density of the solution was found to be 1.029 g/cc at 25°C and 1 atm pressure. Calculate (a) the weight percent, (b) the mole fraction , (c) the mole percent, (d) the molality , (e) the molarity of sulphuric acid in the solution under these conditions.

3. MIXING OF SOLUTIONS :

It is based on law of conservation of moles.

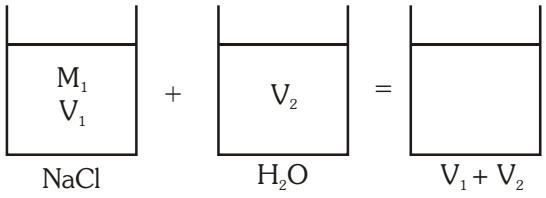
(i) Two solutions having same solute

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



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

$$\text{Final molarity} = \frac{M_1V_1}{V_1 + V_2}$$



n-fold or n-times dilution

⇒ **Final volume = $V_1 + V_2 = n(V_1)$**

Ex.7 ***50 ml 0.2 M H_2SO_4 is mixed with 50 ml 0.3M H_2SO_4. Find molarity of final solution.***

Ans. $M_f = \frac{\text{Total moles of } H_2SO_4}{\text{Total volume}} = \frac{50\times0.2\times10^{-3} + 50\times10^{-3}\times0.3}{(50+50)\times10^{-3}} = \boxed{0.25M}$

Ex.8 ***Find final molarity in each case :***

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

$$M_f = \frac{500\times0.1 + 500\times0.2}{500+500} = \boxed{0.15M}$$

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

$$M_f = \frac{50\times0.1 + 150\times0.3}{50+150+300} = \frac{50}{500} = 0.1\ M$$

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

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

Ex.9 ***Find number of Na^+ & PO_4^{-3} ions in 250 ml of 0.2M Na_3PO_4 solution.***

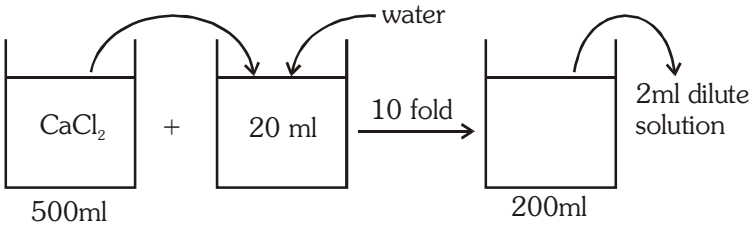
Ans. $Na_3PO_4 + aq. \longrightarrow 3Na^+(aq) + PO_4^{-3}(aq)$ [Ionic compound when added to water ionize completely].

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

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

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

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



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

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

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

Ex.11 ***What volumes of 1M & 2M H_2SO_4 solution are required to produce 2L of 1.75M H_2SO_4 solution?***

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

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

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

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

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

Ex.12 ***80g NaOH was added to 2L water. Find molality of solution if density of water = 1g/mL***

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

Ex.13 ***A 100g NaOH solution has 20g NaOH. Find molality.***

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

Ex.14 ***Find molality of aqueous solution of CH_3COOH whose molarity is 2M and density d = 1.2 g/mL.***

Hint : $\boxed{\mathbf{m = \frac{M}{d - MM_S} \times 1000}}$

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

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

CLASS ILLUSTRATION- (2)

MIXING / DILUTION / REACTION

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

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

10 Find out the volume of 98% w/w H_2SO_4 (density = 1.8 gm/ ml), must be diluted to prepare 12.6 litres of 2.0 M sulphuric acid solution.

11. When V ml of 2.2 M H_2SO_4 solution is mixed with 10 V ml of water, the volume contraction of 2% take place. Calculate the molarity of diluted solution ?

12. Calculate **molality (m)** of each ion present in the aqueous solution of **2M NH_4Cl** assuming 100% dissociation according to reaction.

$$NH_4Cl\,(aq) \longrightarrow NH_4^+\,(aq) + Cl^-\,(aq)$$

Given : Density of solution = 3.107 gm / ml.

4. SOME TYPICAL CONCENTRATION TERMS

4.1 *PERCENTAGE LABELLING OF OLEUM :*

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

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

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

Ex.15 Find the mass of free SO_3 present in 100 gm , 109 % oleum sample.

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

$$SO_3 + H_2O \longrightarrow H_2SO_4$$

9gm

1/2mole 1/2mole

40gm

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

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

II. VOLUME STRENGTH OF H_2O_2 SOLUTION :

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

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

Volume Strength of H_2O_2 Solution = 11.35 × molarity

Ex.16 ***Find the % w/v of "10 V" H_2O_2 solution-***

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

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

CLASS ILLUSTRATION- (3)

13. Find the % labelling of 100 gm oleum sample if it contains 20 gm SO_3.
14. A mixture is prepared by mixing 10 gm H_2SO_4 and 40 gm SO_3 calculate,
(a) mole fraction of H_2SO_4 , (b) % labelling of oleum
15. 500 ml of a H_2O_2 solution on complete decomposition produces 2 moles of H_2O. Calculate the volume strength of H_2O_2 solution?
16. $2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$

Under conditions where 1 mole of gas occupies 24 dm^3, X L of $\frac{1}{24}$ M solution of H_2O_2 produces 3 dm^3 of O_2. Thus X is :-

ANSWER-KEY

CLASS ILLUSTRATION-(1)

1. **Ans.(a) 0.5 M, (b) 0.5 M, (c) 0.2 M** **2.** **Ans.2.4M**

3. **Ans.16.66%** **4.** **Ans.0.06 M**

5. **Ans.45.45%**

6. **Ans.** **(a) =** $\boxed{0.04\,M}$ **(b) =** $\boxed{100\,mol\,kg^{-1}}$

7. **Ans.(a) 5.14 , (b) 0.0098, (c) 0.98 , (d) 0.552 (e) 0.539**

CLASS ILLUSTRATION-(2)

8. **Ans. 0.75 L**

9. **Ans.(1)** $[SO_4^{-2}]_f$ **= 1.95M**
(2) $[Ca^{+2}]_f$ **= 2M**
(3) $[Cl^-]_f$ **= 2.8M**

10. **1.4 litre** **11.** **0.204 M** **12.** **0.6667, 0.6667**

CLASS ILLUSTRATION- (3)

13. **Ans.104.5%** **14.** **(a) 0.169; (b) 118 %**

15. **Ans.45.4 V** **16.** **Ans.(6)**

UNIT-03

EUDIOMETRY

KEY CONCEPTS

EUDIOMETRY :

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

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

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

1 vol. 3 vol. 2 vol.

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

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

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

2 mole 1 mole 2 mole

2 vol. 1 vol. 0 vol.

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

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

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

$V = V_1 + V_2 +$.....

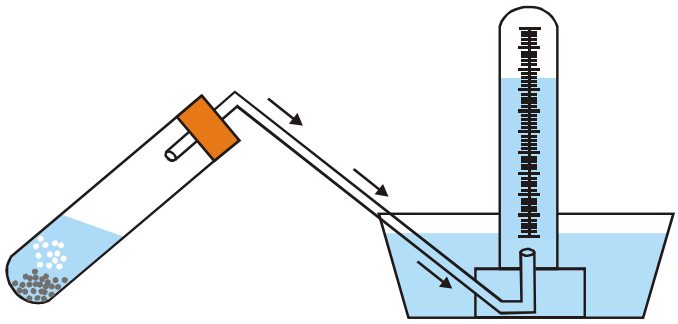
Partial volume of gas in a non-reacting gasesous mixture is its volume when the entire pressure of the mixture is supposed to be exerted only by that gas.

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

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

EUDIOMETER

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



Scheme of eudiometer

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

SOLVED EXAMPLE

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

Ex.1 **Ans.(30)**

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

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

Ex.2 ***A 3 L gas mixture of propane (C_3H_8) and butane (C_4H_{10}) on complete combustion at 25°C produced 10 L CO_2. Assuming constant P and T conditions what was volume of butane present in initial mixture ?***

Ex.2 **Ans.(1)**

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

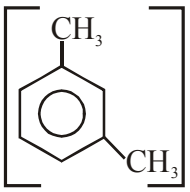
x L 3x L

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

(3-x) L 4(3 – x) L

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

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

Ex.3 ***100 ml gaseous meta Xylene***  ***undergoes combustion with excess of oxygen at room temperature and pressure. Volume contraction / expansion (in ml) during reaction is :***

Ex.3 **Ans. (350)**

$$C_8H_{10}(g) + \frac{21}{2}O_2(g) \longrightarrow 8CO_2(g) + 5H_2O(l)$$

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

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

Ex.4 ***An alkene upon combustion produces $CO_2(g)$ and $H_2O(g)$. In this combustion process if there is no volume change occurs then the no. of C atoms per molecule of alkene will be :***

Ex.4 **Ans.(2)**

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

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

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

Ex.5 ***A gaseous hydrocarbon (C_xH_y) requires 6 times of its own volume of O_2 for complete oxidation and produces 4 times of its volume of CO_2. Find out the volume of x + y.***

Ans. **(012)**

$$C_XH_y + (x + \frac{y}{4})\,O_2 \longrightarrow XCO_2 + \frac{y}{2}H_2O(l)$$

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

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

$ax = 4$ (a)

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

CLASS ILLUSTRATION (1)

1. 20 ml propane gas (C_3H_8) is burnt completely in excess of air. The volume of CO_2 gas formed is.
2. What volume of O_2(g) is needed for complete combustion of 40 ml ethane gas (C_2H_6) ?
3. 10 ml of CO is mixed with 25 ml air having 20% O_2 by volume. What would be the final volume if none of CO and O_2 is left after the reaction?
4. A gaseous alkane is exploded with O_2. The volume of O_2 required for complete combustion and the volume of CO_2 formed after combustion are in 7 : 4 ratio. What is the molecular formula of alkane ?
5. 10 ml of an oxide of nitrogen produce 20 ml NO_2 and 5 ml O_2 on complete decomposition. The oxide of nitrogen is-
6. 30 ml gaseous mixture of methane and ethylene in volume ratio X : Y requires 350 ml air containing 20% of O_2 by volume for complete combustion. If ratio of methane and ethylene changed to Y : X. What will be volume of air (in ml) required for complete reaction under similar condition of temperature and pressure.
7. On heating 60 ml mixture containing equal volume of chlorine gas and it's gaseous oxide, volume becomes 75 ml due complete decomposition of oxide. On treatment with KOH volume becomes 15 ml. What is the formula of oxide of chlorine ?
8. 5 L of A (g) & 3 L of B(g) measured at same T & P are mixed together which react as follows

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

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

ANSWER : CLASS ILLUSTRATION (1)

1.	Ans. 60 ml	2.	Ans. 140 ml.	3.	30 ml	4.	Ans. (C_2H_6)
5.	Ans.(N_2O_5)	6.	Ans. 400	7.	Ans.(Cl_2O)	8.	Ans.(3)

EXERCISE # O-I

UNIT-01 (MOLE CONCEPT)

Single Correct :

1. Which of the following contain largest number of carbon atoms?

(A) 15 gm ethane, C_2H_6 (B) 40.2 gm sodium oxalate, $Na_2C_2O_4$

(C) 72 gm glucose, $C_6H_{12}O_6$ (D) 35 gm pentene, C_5H_{10}

MC0083

2. An iodized salt contains 0.5 % of NaI. A person consumes 3 gm of salt everyday. The number of iodide ions going into his body everyday is (I = 127)

(A) 10^{-4} (B) 6.02×10^{-4} (C) 6.02×10^{19} (D) 6.02×10^{23}

MC0086

3. The percentage by mole of NO_2 in a mixture of NO_2(g) and NO(g) having average molecular mass 34 is :

(A) 25% (B) 20% (C) 40% (D) 75%

MC0087

4. The weight of 2.01×10^{23} molecules of CO is- **[AIEEE 2002]**

(A) 9.3 g (B) 7.2 g (C) 1.2 g (D) 3 g

MC0208

5. How many moles of magnesium phosphate, $Mg_3(PO_4)_2$ will contain 0.25 mole of oxygen atoms? **[AIEEE 2006]**

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

MC0209

6. In the reaction **[AIEEE-2007]**

$2Al_{(s)} + 6HCl_{(aq)} \rightarrow 2Al^{3+}_{(aq)} + 6Cl^{-}_{(aq)} + 3H_2(g)$.

(A) 6L $HCl_{(aq)}$ is consumed for every 3L H_2(g) produced

(B) 33.6 L $H_{2(g)}$ is produced regardless of temperature and pressure for every mole of Al that reacts

(C) 67.2 L $H_{2(g)}$ at STP is produced for every mole of Al that reacts

(D) 11.2 L $H_{2(g)}$ at STP is produced for every mole of of $HCl_{(aq)}$ consumed

MC0210

7. The number of carbon atoms present in a signature, if a signature written by carbon pencil, weighing 1.2×10^{-3} g is

(A) 12.04×10^{20} (B) 6.02×10^{19} (C) 3.01×10^{19} (D) 6.02×10^{20}

MC0088

8. The average atomic mass of a mixture containing 79 mole % of ^{24}Mg and remaining 21 mole % of ^{25}Mg and ^{26}Mg , is 24.31. **% mole** of ^{26}Mg is

(A) 5 (B) 20 (C) 10 (D) 15

MC0089

9. Volume of CO_2 obtained at STP by the complete decomposition of 9.85 g $BaCO_3$ is

(At. wt of Ba = 137)

(A) 2.24 lit (B) 1.12 lit (C) 1.135 lit (D) 2.27 lit

MC0092

10. The drain cleaner, **Drainex** contains small bits of aluminium (At. Wt. = 27) which reacts with caustic soda to produce dihydrogen. What is the volume (in ml) of dihydrogen at 27ºC and 1.013 bar that is produced when 0.27 gm of aluminium reacts :

$2Al + 2NaOH + 2H_2O \rightarrow 2NaAlO_2 + 3H_2$

(A) 0.3694 (B) 369.4 (C) 246.3 (D) 540.4

MC0093

11. Volume of O_2 obtained at 2 atm & 546K, by the complete decomposition of 8.5 g $NaNO_3$ is

$2NaNO_3 \rightarrow 2NaNO_2 + O_2$

(A) 2.24 lit (B) 1.12 lit (C) 0.84 lit (D) 0.56 lit

MC0094

12. Maximum mass of sucrose $C_{12}H_{22}O_{11}$ produced by mixing 84 gm of carbon, 12 gm of hydrogen and 56 lit. O_2 at 1 atm & 273 K according to given reaction, is

$C(s) + H_2(g) + O_2(g) \longrightarrow C_{12}H_{22}O_{11}(s)$

(A) 138.5 (B) 155.5 (C) 172.5 (D) 199.5

MC0211

13. The mass of CO_2 produced from 620 gm mixture of $C_2H_4O_2$ & O_2, prepared to produce maximum energy is (Combustion reaction is exothermic)

(A) 413.33 gm (B) 593.04 gm (C) 440 gm (D) 320 gm

MC0096

14. The mass of Mg_3N_2 produced if 48 gm of Mg metal is reacted with 34 gm NH_3 gas is

$Mg + NH_3 \longrightarrow Mg_3N_2 + H_2$

(A) $\frac{200}{3}$ gm (B) $\frac{100}{3}$ gm (C) $\frac{400}{3}$ gm (D) $\frac{150}{3}$ gm

MC0098

15. 90 gm mixture of H_2 and O_2 is taken in stoichiometric ratio and gives H_2O with 50% yield. The produced mass of H_2O (in gm) is :

(A) 45 gm (B) 36 gm (C) 20 gm (D) 90 gm

MC0103

16. An impure sample of $CaCO_3$ contains 38% of Ca. The percentage of impurity present in the sample is :

(A) 5% (B) 95% (C) 10% (D) 2.5%

MC0104

17. A sample of NH_3 gas is 20% dissociated into N_2 and H_2 gases. The mass ratio of N_2 and NH_3 gases in the final sample is -

(A) $\frac{7}{34}$ (B) $\frac{7}{17}$ (C) $\frac{14}{17}$ (D) $\frac{21}{17}$

MC0106

18. A compound contains 10^{-2} % of phosphorous. If atomic mass of phosphorus is 31, the molecular mass of the compound having one phosphorus atom per molecule is :–

(A) 31 (B) 3.1×10^3 (C) 3.1×10^5 (D) 3.1×10^4

MC0112

19. 13.4 gm of a sample of unstable hydrated salt : $Na_2SO_4.xH_2O$ was strongly heated. Weight loss on heating is found to be equal to 6.3 gm. Calculate the value of x.

(A) 6 (B) 5 (C) 7 (D) 8

CT0113

20. An organic compound contains 4% sulphur by mass. Its minimum molecular weight is :

(A) 200 (B) 400 (C) 800 (D) 1600

MC0114

21. In an organic compound of molar mass108 g mol^{-1} 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) $C_4H_{18}N_3$

MC0181

22. Which of the following series of compounds have same mass percentage of carbon ?

(A) CO_2, CO (B) CH_4, C_2H_6,C_2H_2

(C) C_2H_2, C_6H_6,$C_{10}H_8$ (D)HCHO, CH_3COOH,$C_6H_{12}O_6$

MC0116

23. A compound contains 69.5% oxygen and 30.5% nitrogen and its molecular weight is 92. The formula of that compound is :-

(A) N_2O (B) NO_2 (C) N_2O_4 (D) N_2O_5

MC0117

24. Which of the following compounds has same empirical formula as that of glucose:-

(A) CH_3CHO (B) CH_3COOH (C) CH_3OH (D) C_2H_6

MC119

25. A compound of X and Y has equal mass of them. If their atomic weights are 30 and 20 respectively. The molecular formula of compound is -

(A) X_2Y_2 (B) X_3Y_3 (C) X_2Y_3 (D) X_3Y_2

MC0121

26. 280 g of a mixture containing CH_4 and C_2H_6 in 5 : 2 molar ratio is burnt in presence of excess of oxygen. Calculate total moles of CO_2 produced.

(A) 9 (B) 18 (C) 7 (D) 12

MC0100

27. Mixture of $MgCO_3$ & $NaHCO_3$ on strong heating gives CO_2 & H_2O in 3 : 1 mole ratio. The weight % of $NaHCO_3$ present in the mixture is:

(A) 30% (B) 80% (C) 40% (D) 50%

MC0101

28. Calculate percentage change in M_{avg} of the mixture, if PCl_5 undergo 50% decomposition in a closed vessel.

$$PCl_5 \longrightarrow PCl_3 + Cl_2$$

(A) 50% (B) 66.66 % (C) 33.33 % (D) Zero

MC0182

29. A metal carbonate decomposes according to following reaction

$$M_2CO_3(s) \longrightarrow M_2O(s) + CO_2(g)$$

Percentage loss in mass on complete decomposition of $M_2CO_3(s)$. (Atomic mass of M = 102)

(A) $\frac{100}{3}$% (B) $\frac{50}{3}$% (C) $\frac{25}{3}$% (D) 15%

MC0102

30. Iodobenzene (C_6H_5I) is prepared from aniline ($C_6H_5NH_2$) in a two step process as shown below

$C_6H_5NH_2 + HNO_2 + HCl \rightarrow C_6H_5N_2^+Cl^- + 2H_2O$

$C_6H_5N_2^+Cl^- + KI \rightarrow C_6H_5I + N_2 + KCl$

In an actual preparation 9.30 g of aniline was converted to 16.32 g of iodobenzene. The percentage yield of iodobenzene is : (I = 127)

(A) 8 % (B) 50 % (C) 75 % (D) 80 %

MC0183

UNIT-02 (CONCENTRATION TERM)

31. 125 ml of 8% w/w NaOH solution (sp. gravity 1) is added to 125 ml of 10% w/v HCl solution. The nature of resultant solution would be _____.

(A) Acidic (B) Basic (C) Neutral (D) Can not be predicted

CT0032

32. 8 g NaOH is dissolved in one litre of solution, its molarity is :

(A) 0.8 M (B) 0.4 M (C) 0.2 M (D) 0.1 M

MC0184

33. The molarity of pure water is :

(A) 100 M (B) 55.6 M (C) 50 M (D) 18M

CT0033

34. Mole fraction of $C_3H_5(OH)_3$ (glycerine) in a solution of 36 g of water and 46 g of glycerine is :

(A) 0.46 (B) 0.36 (C) 0.20 (D) 0.40

CT0034

35. The mole fraction of oxygen in a mixture of 7g of nitrogen and 8g of oxygen is :

(A) $\frac{8}{15}$ (B) 0.5 (C) 0.25 (D) 1.0

MC0185

36. The molarity of a solution of sodium chloride (mole wt. = 58.5) in water containg 5.85 gm of sodium chloride in 500 ml of solution is :-

(A) 0.25 (B) 2.0 (C) 1.0 (D) 0.2

CT0036

37. The molarity of 98% by wt. H_2SO_4 (d = 1.8 g/ml) is

(A) 6 M (B) 18 M (B) 10 M (D) 4 M

CT0037

38. Which one of the following modes of expressing concentration of solution is independent of temperature -

(A) Molarity (B) Molality (C) % w/v (D) Grams per litre

CT0038

39. For preparing 0.1 M solution of H_2SO_4 in one litre, we need H_2SO_4 :

(A) 0.98 g (B) 4.9 g (C) 49.0 g (D) 9.8 g

MC0186

40. The relationship between mole fraction (X_A) of the solute & molality 'm' of its solution in ammonia would be

(A) $\frac{55.56(X_A)}{1-X_A}$ =m (B) $\frac{58.82(X_A)}{1-X_A}$ = m (C) $\frac{58.82(1-X_A)}{X_A}$ =m (D) $\frac{55.56(1-X_A)}{X_A}$ =m

CT0052

41. 3.0 molal NaOH solution has a density of 1.12 g/mL. The molarity of the solution is-
(A) 2.97 (B) 3 (C) 3.05 (D) 3.5

CT0053

42. H_2O_2 solution used for hair bleaching is sold as a solution of approximately 5.0 g H_2O_2 per 100 mL of the solution. The molecular mass of H_2O_2 is 34. The molarity of this solution is approximately:-
(A) 0.15 M (B) 1.5 M (C) 3.0 M (D) 3.4 M

MC0187

43. Molality of 20% (w/w) aglucose solution is :

(A) $\frac{25}{18}$m (B) $\frac{10}{9}$m (C) $\frac{25}{9}$m (D) $\frac{5}{18}$m

CT0042

44. Molarity of liquid HCl, if density is 1.17 g/cc. :
(A) 36.5 M (B) 18.25 M (C) 32.05 M (D) 42.10 M

CT0043

45. The molarity of a solution made by mixing 50 ml of conc. H_2SO_4 (18 M) with 50 ml of water, is:
(A) 36 M (B) 18 M (C) 9 M (D) 6M

CT0044

46. Molarity and Molality of a solute (M. wt = 50) in aqueous solution is 9 and 18 respectively. What is the density of solution.
(A) 1 g/cc (B) 0.95 g/cc (C) 1.05 g/cc (D) 0.662 g/cc

CT0051

47. 34 g of hydrogen peroxide is present in 1135 mL of solution. Volume strength of solution is:
(A) 10 V (B) 20 V (C) 30 V (D) 32 V

MC0188

48. Label an oleum sample which has mass fraction of SO_3 equal to 0.6 :
(A) 115 % (B) 109 % (C) 104.5 % (D) 113.5 %

MC0189

49. If 50 gm oleum sample rated as 118% is mixed with 18 gm water, then the correct option is
(A) The resulting solution contains 18 gm of water and 118 gm H_2SO_4
(B) The resulting solution contains 9 gm water and 59 gm H_2SO_4
(C) The resulting solution contains only 118 gm pure H_2SO_4
(D) The resulting solution contains 68 gm of pure H_2SO_4

CT0047

50. 12.5gm of fuming H_2SO_4 (labelled as 112%) is mixed with 100 lit water. Molar concentration of H^+ in resultant solution is :
[Note : Assume that H_2SO_4 dissociate completely and there is no change in volume on mixing]

(A) $\frac{2}{700}$ (B) $\frac{2}{350}$ (C) $\frac{3}{350}$ (D) $\frac{3}{700}$

CT0048

UNIT-03 (EUDIOMETRY)

51. 10 ml CH_4 gas is burnt completely in air (O_2 = 20%, by volume). The minimum volume of air needed is -

(A) 20 ml (B) 50 ml (C) 80 ml (D) 100 ml

MC0122

52. When 20 ml mixture of O_2 and O_3 is heated, the volume becomes 29 ml and disappears in alkaline pyragallol solution. What is the volume percent of O_2 in the original mixture?

(A) 90% (B) 10% (C) 18% (D) 2%

MC0190

53. One litre of CO_2 passed over hot coke the volume becomes 1.4 litres then the composition of products will not be (At NTP)

(A) $V_{CO_2} : V_{CO} = 3:4$ (B) $V_{CO_2} = 1.6$ ltr. (C) $n_{CO_2} : n_{CO} = 3:4$ (D) % V of CO $= \frac{400}{7}$

MC0191

54. 20 mL of a gaseous hydrocarbon was exploded with 120 mL of oxygen. A contraction of 60 mL was observed, and a further contraction of 60 mL took place when KOH was added. What is the formula of the hydrocarbon :

(A) C_3H_6 (B) C_3H_8 (C) C_2H_6 (D) C_4H_{10}

MC0192

55. Each volume of a gaseous organic compound containing C, H and S only produce 1 volume CO_2, 2 volume H_2O vapours and 1 volume SO_2 gases on complete combustion. The molecular formula of compound is -

(A) CH_2S (B) CH_4S (C) C_2H_4S (D) C_2H_6S

MC0126

56. How many litres of oxygen at 1atm & 273K will be required to burn completely 2.2 g of propane (C_3H_8)

(A) 11.2 L (B) 22.4 L (C) 5.6 L (D) 44.8 L

MC0090

57. An ideal gaseous mixture of ethane (C_2H_6) and ethene (C_2H_4) occupies 28 litre at 1atm, 0°C. The mixture reacts completely with 128 gm O_2 to produce CO_2 and H_2O. Mole fraction of C_2H_6 in the mixture is–

(A) 0.6 (B) 0.4 (C) 0.5 (D) 0.8

MC0099

EXERCISE # O-II

UNIT-01 (MOLE CONCEPT)

1. 12 g of Mg was burnt in a closed vessel containing 32 g oxygen. Which of the following is /are correct.
(A) 2 gm of Mg will be left unburnt.
(B) 0.75 gm-molecule of O_2 will be left unreacted.
(C) 20 gm of MgO will be formed.
(D) The mixture at the end will weight 44 g.

MC0139

2. 50 gm of $CaCO_3$ is allowed to react with 68.6 gm of H_3PO_4 then select the correct option(s)-

$$3CaCO_3 + 2H_3PO_4 \rightarrow Ca_3(PO_4)_2 + 3H_2O + 3CO_2$$

(A) 51.67 gm salt is formed
(B) Amount of unreacted reagent = 35.93 gm
(C) $n_{CO_2} = 0.5$ moles
(D) 0.7 mole CO_2 is evolved

MC0140

3. Select the correct statement(s) for $(NH_4)_3PO_4$.
(A) Ratio of number of oxygen atoms to number of hydrogen atoms is 1 : 3
(B) Ratio of number of cations to number of anions is 3 : 1
(C) Ratio of number of gm-atoms of nitrogen to gm-atoms of oxygen is 3 : 2
(D) Total number of atoms in one mole of $(NH_4)_3PO_4$ is 20.

MC0141

4. The recommended daily dose is 17.6 milligrams of vitamin C (ascorbic acid) having formula $C_6H_8O_6$. Match the following. Given : $N_A = 6 \times 10^{23}$

Column I		**Column II**	
(A)	O-atoms present in daily dose	(P)	10^{-4} mole
(B)	Moles of vitamin C in 1 gm of vitamin C	(Q)	5.68×10^{-3}
(C)	Moles of vitamin C that should be consumed daily	(R)	3.6×10^{20}

MC0153

Paragraph for 5 to 7

NaBr, used to produce AgBr for use in photography can be self prepared as follows :

$$Fe + Br_2 \longrightarrow FeBr_2 \quad ....(i)$$

$$FeBr_2 + Br_2 \longrightarrow Fe_3Br_8 \quad ....(ii) \quad \text{(not balanced)}$$

$$Fe_3Br_8 + Na_2CO_3 \longrightarrow NaBr + CO_2 + Fe_3O_4 \quad ....(iii) \quad \text{(not balanced)}$$

(At. mass : Fe = 56, Br = 80)

5. Mass of iron required to produce 2.06×10^3 kg NaBr
(A) 420 gm (B) 420 kg (C) 4.2×10^5 kg (D) 4.2×10^8 gm

MC0145

6. If the yield of (ii) is 60% & (iii) reaction is 70% then mass of iron required to produce 2.06×10^3 kg NaBr
(A) 10^5 kg (B) 10^5 gm (C) 10^3 kg (D) None

MC0146

7. If yield of (iii) reaction is 90% then mole of CO_2 formed when 2.06×10^3 gm NaBr is formed
(A) 20 (B) 10 (C) 9 (D) 440

MC0147

UNIT-02 (CONCENTRATION TERM)

8. **Statement -1** : Molality of pure ethanol is lesser than pure water.

Statement -2 : As density of ethanol is lesser than density of water.

[Given : $d_{ethanol}$ = 0.789 gm/ml; d_{water} = 1 gm/ml]

(A) Statement-1 is true, statement-2 is true and statement-2 is correct explanation for statement-1.

(B) Statement-1 is true, statement-2 is true and statement-2 is NOT the correct explanation for statement-1.

(C) Statement-1 is false, statement-2 is true.

(D) Statement-1 is true, statement-2 is false.

CT0054

9. Solution(s) containing 40 gm NaOH is/are

(A) 50 gm of 80% (w/w) NaOH

(B) 50 gm of 80% (w/v) NaOH [$d_{soln.}$ = 1.2 gm/ml]

(C) 50 gm of 20 M NaOH [$d_{soln.}$ = 1 gm/ml]

(D) 50 gm of 5m NaOH

CT0058

10. The **incorrect** statement(s) regarding 2M $MgCl_2$ aqueous solution is/are ($d_{solution}$ = 1.09 gm/ml)

(A) Molality of Cl^- is **4.44 m**

(B) Mole fraction of $MgCl_2$ is exactly **0.035**

(C) The conc. of $MgCl_2$ is **19% w/v**

(D) The conc. of $MgCl_2$ is **19×10^4 ppm**

CT0059

11. A sample of H_2O_2 solution labelled as 56.75 volume has density of 530 gm/L. Mark the correct option(s) representing concentration of same solution in other units. (Solution contains only H_2O and H_2O_2)

(A) $M_{H_2O_2} = 6$

(B) $\% \frac{w}{v} = 17$

(C) Mole fraction of $H_2O_2 = 0.25$

(D) $m_{H_2O_2} = \frac{1000}{72}$

CT0060

Comprehension Q.12 and Q.13 (2 questions)

2 litre of 9.8 % w/w H_2SO_4 (d = 1.5 gm/ml) solution is mixed with 3 litre of 1 M KOH solution.

12 The number of moles H_2SO_4 added are

(A) 1 (B) 2 (C) 3 (D) 0.5

CT0062

13. The concentration of H^+ if solution is acidic or concentration of OH^- if solution is basic in the final solution is

(A) 0 (B) $\frac{3}{10}$ (C) $\frac{3}{5}$ (D) $\frac{2}{5}$

CT0063

UNIT-03 (EUDIOMETRY)

14. 20 ml mixture of C_3H_8 and CO gas when burnt in excess of oxygen produce 40 ml CO_2 gas. Choose the correct statement(s) (Volume of gases measured under same T & P)
(A) Volume of C_3H_8 in the mixture is 15 ml
(B) Volume of CO in the mixture is 10 ml
(C) Total volume contraction due to combustion is 35 ml.
(D) The volume of oxygen used for combustion is 75 ml

MC0193

EXERCISE # J-MAINS

1. A transition metal M forms a volatile chloride which has a vapour density of 94.8. If it contains 74.75% of chlorine the formula of the metal chloride will be **[AIEEE 2012 (Online)]**

(A) MCl_2 (B) MCl_4 (C) MCl_5 (D) MCl_3

MC0176

2. The ratio of number of oxygen atoms (O) in 16.0 g ozone (O_3), 28.0 g carbon monoxide (CO) and 16.0 g oxygen (O_2) is :-

(Atomic mass : C = 12, O = 16 and Avogadro's constant $N_A = 6.0 \times 10^{23}$ mol^{-1})

[AIEEE 2012 (Online)]

(A) 3 : 1 : 1 (B) 1 : 1 : 2 (C) 3 : 1 : 2 (D) 1 : 1 : 1

MC0177

3. 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)-2013]**

(A) C_2H_4 (B) C_3H_4 (C) C_6H_5 (D) C_7H_8

MC0175

4. The ratio of masses of oxygen and nitrogen in a particular gaseous mixture is 1 : 4. The ratio of number of their molecule is : **[JEE(Main)-2014]**

(A) 1 : 8 (B) 3 : 16 (C) 1 : 4 (D) 7 : 32

MC0174

5. In Carius method of estimation of halogens, 250 mg of an organic compound gave 141 mg of AgBr. The percentage of bromine in the compound is :

(Atomic mass Ag = 108; Br = 80) **[JEE(Main)-2015]**

(A) 48 (B) 60 (C) 24 (D) 36

MC0173

6. The most abundant elements by mass in the body of a healthy human adult are : Oxygen (61.4%) ; Carbon (22.9%), Hydrogen (10.0%) ; and Nitrogen (2.6%). The weight which a 75 kg person would gain if all ^1H atoms are replaced by ^2H atoms is **[JEE(Main)-2017]**

(A) 15 kg (B) 37.5 kg (C) 7.5 kg (D) 10 kg

MC0171

7. 1 gram of a carbonate (M_2CO_3) on treatment with excess HCl produces 0.01186 mole of CO_2. the molar mass of M_2CO_3 in g mol^{-1} is :- **[JEE(Main)-2017]**

(A) 1186 (B) 84.3 (C)118.6 (D) 11.86

MC0172

8. The ratio of mass percent of C and H of an organic compound ($C_XH_YO_Z$) is 6 : 1. If one molecule of the above compound ($C_XH_YO_Z$) contains half as much oxygen as required to burn one molecule of compound C_XH_Y completely to CO_2 and H_2O. The empirical formula of compound $C_XH_YO_Z$ is

[JEE(Main)-2018 (offline)]

(A) C_2H_4O (B) $C_3H_4O_2$ (C) $C_2H_4O_3$ (D) $C_3H_6O_3$

MC0168

9. For per gram of reactant, the maximum quantity of N_2 gas is produced in which of the following thermal decomposition reactions ? **[JEE(Main)-2018 (online)]**

(Given : Atomic wt. – Cr = 52u, Ba = 137u)

(A) $2NH_4NO_3(s) \rightarrow 2N_2(g) + 4H_2O(g) + O_2(g)$

(B) $Ba(N_3)_2(s) \rightarrow Ba(s) + 3N_2(g)$

(C) $(NH_4)_2Cr_2O_7(s) \rightarrow N_2(g) + 4H_2O(g)$

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

MC0169

10. An unknown chlorohydrocarbon has 3.55% of chlorine. If each molecule of the hydrocarbon has one chlorine atom only; chlorine atoms present in 1 g of chlorohydrocarbon are :

(Atomic wt. of Cl = 35.5 u; Avogadro constant = 6.023×10^{23} mol^{-1})**[JEE(Main)-2018 (online)]**

(A) 6.023×10^{21} (B) 6.023×10^{23} (C) 6.023×10^{20} (D) 6.023×10^{9}

MC0170

11. A solution of two components containing n_1 moles of the 1st component and n_2 moles of the 2nd component is prepared. M_1 and M_2 are the molecular weights of component 1 and 2 respectively. If d is the density of the solution in g mL^{-1}, C_2 is the molarity and x_2 is the mole fraction of the 2nd component, then C_2 can be expressed as : **[JEE(Main)-2020 (online)]**

(1) $C_2 = \frac{1000x_2}{M_1 + x_2(M_2 - M_1)}$ (2) $C_2 = \frac{dx_2}{M_2 + x_2(M_2 - M_1)}$

(3) $C_2 = \frac{dx_1}{M_2 + x_2(M_2 - M_1)}$ (4) $C_2 = \frac{1000dx_2}{M_1 + x_2(M_2 - M_1)}$

MC0194

12. The strengths of 5.6 volume hydrogen peroxide (of density 1 g/mL) in terms of mass percentage and molarity (M), respectively, are: **[JEE(Main)-2020 (online)]**

(Take molar mass of hydrogen peroxide as 34 g/mol)

(1) 1.7 and 0.25 (2) 1.7 and 0.5

(3) 0.85 and 0.5 (4) 0.85 and 0.25

MC0195

13. 6.023×10^{22} molecules are present in 10 g of a substance 'x'. The molarity of a solution containing 5 g of substance 'x' in 2 L solution is _____ $\times 10^{-3}$. **[JEE(Main)-2020 (online)]**

MC0196

14. The mass of ammonia in grams produced when 2.8 kg of dinitrogen quantitatively reacts with 1 kg of dihydrogen is _____. **[JEE(Main)-2020 (online)]**

MC0197

15. The minimum number of moles of O_2 required for complete combustion of 1 mole of propane and 2 moles of butane is _____. **[JEE(Main)-2020 (online)]**

MC0198

16. The average molar mass of chlorine is 35.5 g mol^{-1}. The ratio of ^{35}Cl to ^{37}Cl in naturally occurring chlorine is close to : **[JEE(Main)-2020 (online)]**

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

MC0199

17. In an estimation of bromine by Carius method, 1.6 g of an organic compound gave 1.88 g of AgBr. The mass percentage of bromine in the compound is _____ **[JEE(Main)-2020 (online)]**

(Atomic mass, Ag=108, Br = 80 g mol^{-1})

MC0200

18. The ratio of the mass percentages of 'C & H' and 'C & O' of a saturated acyclic organic compound 'X' are 4 : 1 and 3 : 4 respectively. Then, the moles of oxygen gas required for complete combustion of two moles of organic compound 'X' is _____ . **[JEE(Main)-2020 (online)]**

MC0201

19. Complex A has a composition of $H_{12}O_6Cl_3Cr$. If the complex on treatment with conc. H_2SO_4 loses 13.5% of its original mass, the correct molecular formula of A is : **[JEE(Main)-2020 (online)]**

[Given : atomic mass of Cr = 52 amu and Cl = 35 amu]

(1) $[Cr(H_2O)_5Cl]Cl_2 \cdot H_2O$ (2) $[Cr(H_2O)_3Cl_3] \cdot 3H_2O$

(3) $[Cr(H_2O)_4Cl_2]Cl \cdot 2H_2O$ (4) $[Cr(H_2O)_6]Cl_3$

MC0202

EXERCISE # J-ADVANCED

1. How many moles of e– weight one Kg : **[JEE '2002 (Scr), 1]**

(A) 6.023×10^{23} (B) $\frac{1}{9.108} \times 10^{31}$ (C) $\frac{6.023}{9.108} \times 10^{54}$ (D) $\frac{1}{9.108 \times 6.023} \times 10^{8}$

MC0204

2. Calculate the amount of Calcium oxide required when it reacts with 852 g of P_4O_{10}. **[JEE 2005]**

$6CaO + P_4O_{10} \longrightarrow 2\ Ca_3(PO_4)_2$ [Ca = 40, P = 31]

MC0205

3. Given that the abundances of isotopes ^{54}Fe, ^{56}Fe and ^{57}Fe are 5%, 90% and 5%, respectively, the atomic mass of Fe is : **[JEE 2009]**

(A) 55.85 (B) 55.95 (C) 55.75 (D) 56.05

MC0206

4. The ammonia prepared by treating ammonium sulphate with calcium hydroxide is completely used by $NiCl_2.6H_2O$ to form a stable coordination compound. Assume that both the reactions are 100% complete. If 1584 g of ammonium sulphate and 952g of $NiCl_2.6H_2O$ are used in the preparation, the combined weight (in grams) of gypsum and the nickel-ammonia coordination compound thus produced is___.

[JEE 2018]

(Atomic weights in g mol^{-1}: H = 1, N = 14, O = 16, S = 32, Cl = 35.5, Ca = 40, Ni = 59)

MC0178

5. Galena (an ore) is partially oxidized by passing air through it at high temperature. After some time, the passage of air is stopped, but the heating is continued in a closed furnance such that the contents undergo self-reduction. The weight (in kg) of Pb produced per kg of O_2 consumed is _____ .

(Atomic weights in g mol^{-1} : O = 16, S = 32, Pb = 207) **[JEE 2018]**

MC0179

6. Aluminium reacts with sulfuric acid to form aluminium sulfate and hydrogen. What is the volume of hydrogen gas in liters (L) produced at 300 K and 1.0 atm pressure, when 5.4 g of aluminium and 50.0 mL of 5.0 M sulfuric acid are combined for the reaction ? **[JEE 2020]**

(Use molar mass of aluminium as 27.0 g mol^{-1} , R = 0.082 atm L mol^{-1} K^{-1})

MC0207

ANSWER KEY

EXERCISE # O-I

UNIT-01 (MOLE CONCEPT)

1.	Ans.(D)	2.	Ans.(C)	3	Ans.(A)	4.	Ans.(A)	5.	Ans.(A)
6.	Ans.(D)	7.	Ans.(B)	8	Ans.(C)	9.	Ans.(C)	10.	Ans.(B)
11.	Ans (B)	12.	Ans.(B)	13.	Ans.(C)	14.	Ans.(A)	15.	Ans.(A)
16.	Ans.(A)	17.	Ans.(A)	18.	Ans.(C)	19.	Ans.(C)	20.	Ans.(C)
21.	Ans.(A)	22.	Ans.(D)	23.	Ans.(C)	24.	Ans.(B)	25.	Ans.(C)
26.	Ans.(B)	27.	Ans.(D)	28	Ans.(C)	29.	Ans. (B)	30.	Ans.(D)

UNIT-02 (CONCENTRATION TERM)

31.	Ans.(A)	32.	Ans. (C)	33.	Ans. (B)	34.	Ans. (C)	35.	Ans. (B)
36.	Ans.(D)	37.	Ans.(B)	38.	Ans. (B)	39.	Ans. (D)	40.	Ans. (B)
41.	Ans. (B)	42.	Ans. (B)	43.	Ans. (A)	44.	Ans. (C)	45.	Ans. (C)
46.	Ans. (B)	47.	Ans.(A)	48.	Ans.(D)	49.	Ans. (B)	50.	Ans. (A)

UNIT-03 (EUDIOMETRY)

51.	Ans.(D)	52.	Ans.(B)	53.	Ans.(B)	54.	Ans.(B)	55.	Ans.(B)
56.	Ans. (C)	57.	Ans (B)						

EXERCISE # O-II

UNIT-01 (MOLE CONCEPT)

1.	Ans.(B,C,D)	2.	Ans.(A,B,C)	3.	Ans. (A, B)	4.	Ans. (A) R, (B) Q, (C) P
5.	Ans.(B)	6.	Ans.(C)	7.	Ans.(B)		

UNIT-02 (CONCENTRATION TERM)

8.	Ans.(B)	9.	Ans. (A, C)	10.	Ans. (B, D)	11.	Ans. (B, D)
12	Ans.(C)	13	Ans.(C)	14.	Ans.(B,C)		

EXERCISE # J-MAINS

1.	Ans. (B)	2.	Ans. (D)	3.	Ans. (D)	4.	Ans. (D)
5.	Ans. (C)	6.	Ans.(C)	7.	Ans.(B)	8.	Ans.(C)
9.	Ans.(D)	10.	Ans.(C)	11.	Ans.(4)	12.	Ans.(2)
13.	Ans.(25)	14.	Ans.(3400)	15.	Ans.(18)	16.	Ans.(4)
17.	Ans.(50.00)	18.	Ans.(5.00)	19.	Ans.(3)		

EXERCISE # J-ADVANCED

1.	Ans. (D)	2.	Ans. 1008 g	3.	Ans.(B)	4.	Ans. (2992)
5.	Ans.(6.47)	6.	Ans.(6.15)				