NURTURE COURSE

MOLE CONCEPT

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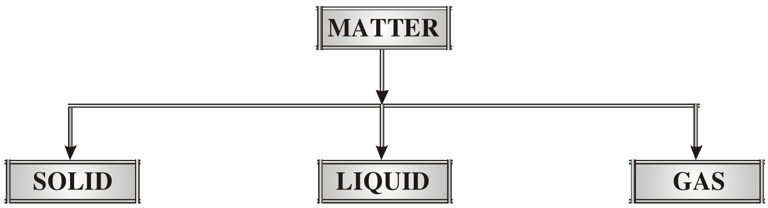
CLASSIFICATION OF MATTER

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

  Two ways of classifying matter.

  I. Physical classification II. Chemical classification

I. Physical classification :

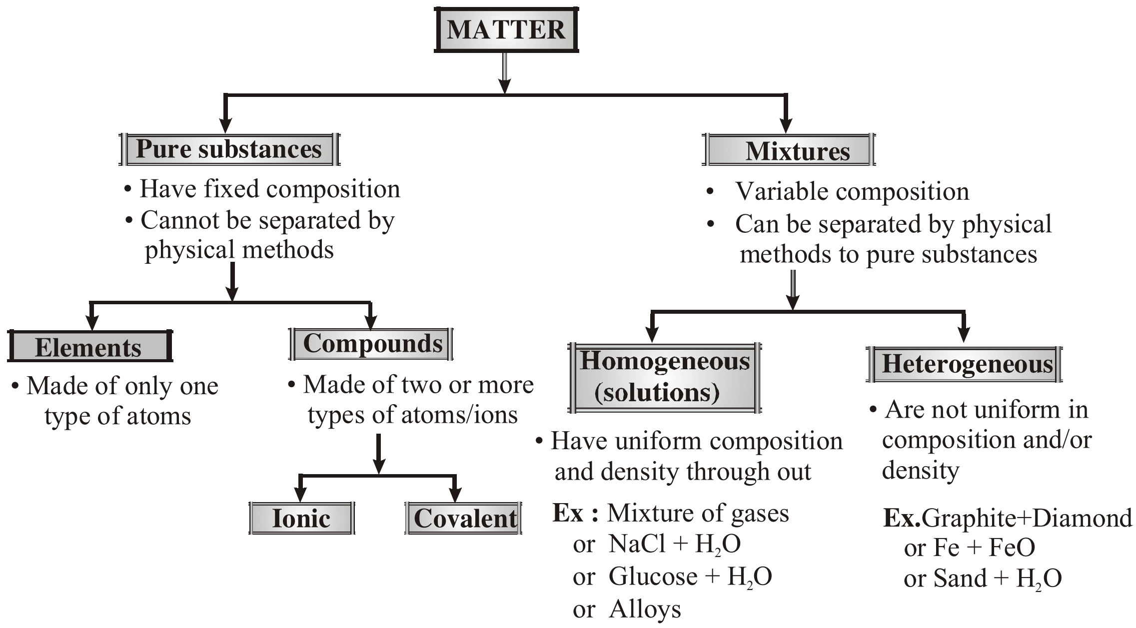


	SOLID	LIQUID	GAS
(i)	Particles held very closely packed in ordered manner.	Particles are less closely packed.	Particles are farthest apart
(ii)	No freedom of movement of particles	Particles can move around to some extent	Movement of particles is very easy and fast
(iii)	Definite shape and volume	Definite 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 :

- 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

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.

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 :

- 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.

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.

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.

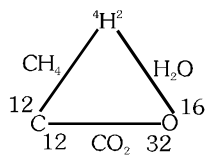
Ex. Nitrogen and oxygen combine to form five stable oxides –

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 = $\dfrac{82.35}{17.65}$ 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 = $\dfrac{88.90}{11.10}$ 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.

❖ **Application of Avogadro's hypothesis :**

(a) In finding the atomicity

(b) Relation between molecular weight and vapour density

• Vapour density $= \dfrac{\text{density of gas}}{\text{density of } H_2}$ (constant P, T)

• Molecular weight = 2 × vapour density

• Density of H_2 = 0.000089 gm/cm^3

$\approx$ 0.00009 gm/ml $\approx$ 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.

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.

DALTON'S ATOMIC THEORY

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.

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).

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 1/12th mass of an atom of C^{12} is known as atomic mass unit.

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

$\therefore$ 1/12th 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.

= mass of N_A atoms of oxygen = $\left(\frac{16}{N_A} g\right) \times N_A$ = 16 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) $\because$ 6.022×10^{23} atoms weighs 50 gm

$\therefore$ 3.011×10^{20} atoms weighs $\dfrac{50}{6.022\times10^{23}}\times3.011\times10^{20} = 0.025$ 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 = $\dfrac{\Sigma(\%\text{abundance}\times\text{atomic mass})}{100} = \dfrac{90\times30+10\times32}{100} = 30.2$

◻ 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

= $\left(\dfrac{98}{N_A}g\right)\times N_A = 98$ 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 = $\dfrac{\text{Given mass of substance (in gm)}}{\text{GAM / GMM}}$

(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 °C

(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 0ºC and 1 atm (vi) 4.54 ml SO_3 gas at STP (vii) 8.21 L C_2H_6 gas at 400K and 2atm (viii) 164.2 ml He gas at 27ºC and 570 torr [$N_A = 6 \times 10^{23}$]***

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

$$\text{Number of moles } = \frac{\text{no. of molecules}}{N_A} = \frac{4.5 \times 10^{24}}{6 \times 10^{23}} = 7.5$$

(iv) 2.4×10^{21} atoms of hydrogen

$$\text{Number of gram molecules of } H_2 = \frac{\text{no. of molecules}}{N_A} = \frac{2.4 \times 10^{21}}{2 \times 6 \times 10^{23}} = 0.002$$

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

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

(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

$$n = \frac{PV}{R.T} = \frac{2 \times 8.21}{0.0821 \times 400} = 0.5$$

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

$$n = \frac{PV}{RT} = \left(\frac{570}{760}\text{atm}\right)\times\frac{164.2\times10^{-3}\text{litre}}{0.0821\times300} = 0.005$$

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

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 $Na_2SO_4 . 7H_2O$

$\Rightarrow 6.023 \times 10^{22}$ O – atoms are in $= \frac{275}{11\times6.023\times10^{23}}\times6.023\times10^{22}$ g = 2.5 g

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

Sol. 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.

DENSITY :

It is of two types.

I. Absolute density

II. Relative density

❖ For liquids and solids :

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

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

❖ For gases :

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.

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

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

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

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

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

$= \frac{200}{100}$ mol = 2 mol 2 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.

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 $\frac{28}{7} = 4$ $\frac{48}{32} = 1.5$

$\frac{\text{moles taken}}{\text{stoich. coeff.}}$ $\frac{4}{4} = 1$ $\frac{1.5}{1} = 1.5$

(L.R.)

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

Mass of Li_2O formed = 2 × 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. $$12C(s) + 11\ H_2(g) + \frac{11}{2}O_2 \rightarrow C_{12}H_{22}O_{11}\ (s)$$

	C	H_2	O_2
Moles taken	$\frac{78}{12} = 6.5$	$\frac{11}{2} = 5.5$	$\frac{67.2}{22.4} = 3$
$\frac{\text{moles taken}}{\text{stoich.coeff.}}$	$\frac{6.5}{12} = 0.54$	$\frac{5.5}{11} = 0.5$ (L.R.)	$\frac{3}{5.5} = 0.545$

$\therefore$ 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.

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

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

$= 480$ 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 $= \frac{1.2}{1.5}\times 100 = 80\%$

◻ 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 $= 0.2$ $\qquad$ $\frac{12.8}{32}$ gm $= 0.4$

$\frac{\text{moles taken}}{\text{stoich. coeff.}}$ $\quad \frac{0.2}{2} = 0.1$ (L.R.) $\qquad \frac{0.4}{3} = 0.133$

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

mass $= 0.1 \times 150 = 15$ 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.

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

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

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

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 :

$= \frac{1}{2}\times\frac{17.52}{36.5}\times\frac{100\times100}{72} = 33.33$ gm

---

◻ 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]

	C	$\frac{1}{2}O_2$	CO
t = 0	1mol	5mol	
t = ∞	0 (LR)	5–0.5 = 4.5mol	1mol

Since CO & O_2 are left CO_2 is formed.

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

	CO	$\frac{1}{2}O_2$	CO_2
t = 0	1mol	4.5mol	0
t = ∞	0	4 mol	1 mol

At end 1 mole CO_2 & no CO present

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

	C	$\frac{1}{2}O_2$	CO
t = 0	10mol	5mol	0
t =∞	0	0	10mol

At end only 10 mol CO present.

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

	C	$\frac{1}{2}O_2$	CO
t = 0	6mol	5mol	0
t =∞	0 [LR]	2mol	6mol

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

	CO	$\frac{1}{2}O_2$	CO_2
t = 0	6mol	2mol	0
t =∞	2mol	0 [LR]	4 mol

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

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

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

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) → 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 × moles of $KClO_3$

and moles of K atoms in KCl = 1 × moles of KCl

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

- 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.

Ex.26 ***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)***

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

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

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

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

magnesium metaphosphate.

Sol. POAC on Mg

$1 \times n_{Mg} = 1 \times n_{Mg(PO_3)_2}$

$\dfrac{192}{24} = n_{Mg(PO_3)_2}$

$W_{Mg(PO_3)_2} = \dfrac{192}{24} \times 182\text{gm} = 1456$ gm.

Ex.27 ***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 $K_2Zn_3[Fe(CN)_6]_2 = 698$]***

Sol. Here we have no knowledge about series of chemical reactions but we know about initial reactant and final product, accordingly.

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

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

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

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

($\because$ 1 mole of K_2CO_3 contains 1 moles of C)

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

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

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*)

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

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

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

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*)

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

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*)

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

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

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 :**

$\text{O.C} + HNO_3 \rightarrow H_3PO_4 + [NH_3 + \text{magnesia mixture ammonium molybdate}] \rightarrow MgNH_4PO_4 \xrightarrow{\Delta} Mg_2P_2O_7$

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

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

$\text{O.C.} + HNO_3 + AgNO_3 \rightarrow AgX$

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.28 ***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 = $\frac{0.6}{40} = 15\text{mmol.}$

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

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

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 2 : 2	C_6H_6 6 : 6	C_2H_6 2 : 6	$C_2H_4O_2$ 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.30. ***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

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

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

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

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

Ex.31 ***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 × CH_2 = C_2H_4

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.32 ***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

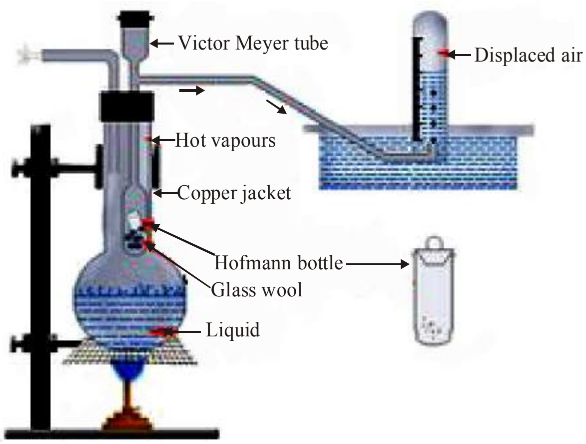
Room temperature = TK

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)



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

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

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

Ex.33 ***0.15 g of substance displaced 58.9 cm³ 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

Molecular mass = $\frac{719.3}{760}\times\frac{58.9}{1000}=\frac{0.15}{M}\times.0821\times300$

$\therefore$ 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.

Organic acid $\xrightarrow{AgNO_3}$ Silver salt $\xrightarrow{Ignite}$ 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)

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

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

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

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

Molar mass of acid = $\frac{108\times nW}{x}-n\times108+n\times1=n\left(\frac{108W}{x}-107\right)gmol^{-1}$

Ex.34 ***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

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

Now molar mass acid $=n\left(\frac{108W}{x}-107\right)gmol^{-1}=2\left(\frac{108\times0.41}{0.216}-107\right)gmol^{-1}=196gmol^{-1}$

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)

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

Ex.35 ***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

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

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

EXERCISE # S-I

PROBLEMS RELATED WITH DIFFERENT TYPES OF ATOMIC MASSES & BASIC CONCEPT OF MOLE

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. What will be the mass of one ^{12}C atom in g ?

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

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

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

6. 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)

7. 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]

8. Calculate volume of H_2 gas kept at STP if it contains as many H atoms as in 98 gm H_3PO_4. [Atomic mass of P = 31]

9. 80gm of SO_x gas occupies 14 litre at 2atm & 273K. The value of x is :

10. 40 mg of gaseous substance (X_2) occupies 4.8 mL of volume at 1 atm and 27°C. Atomic mass of element X is : (R : 0.08 atm L/mole-K)

STOICHIOMETRY

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

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

12. 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 ?

13. Flourine reacts with uranium to produce uranium hexafluoride, UF_6, as represented by this equation

$$U(s) + 3F_2(g) \rightarrow UF_6(g)$$

How many fluorine molecules are required to produce 7.04 mg of uranium hexafluoride, UF_6, from an excess of uranium ? The molar mass of UF_6 is 352 gm/mol. ($N_A = 6 \times 10^{23}$)

14. What total volume, in litre at 627°C and 0.821 atm, could be formed by the decomposition of 16 gm of NH_4NO_3 ? Reaction : $2\,NH_4NO_3 \rightarrow 2N_2 + O_2 + 4H_2O_{(g)}$.

15. Calculate mass of phosphoric acid required to obtain 53.4g pyrophosphoric acid.

$$2H_3PO_4 \rightarrow H_4P_2O_7 + H_2O$$

LIMITING REACTANT

16. 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.

17. Potassium superoxide, KO_2, is used in rebreathing gas masks to generate oxygen :

$$KO_2(s) + H_2O(l) \rightarrow KOH(s) + O_2(g)$$

If a reaction vessel contains 0.158 mol KO_2 and 0.10 mol H_2O, how many moles of O_2 can be produced ?

18. 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.

19. Sulphuric acid is produced when sulphur dioxide reacts with oxygen and water in the presence of a catalyst : $2SO_2(g) + O_2(g) + 2H_2O(l) \rightarrow 2H_2SO_4$. If 5.6 mol of SO_2 reacts with 4.8 mol of O_2 and a large excess of water, what is the maximum number of moles of H_2SO_4 that can be obtained ?

20. Titanium, which is used to make air plane engines and frames, can be obtained from titanium tetrachloride, which in turn is obtained from titanium oxide by the following process :

$$3\ TiO_2(s) + 4C(s) + 6Cl_2(g) \longrightarrow 3TiCl_4(g) + 2CO_2(g) + 2CO(g)$$

A vessel contains 4.32 g TiO_2, 5.76 g C and; 7.1 g Cl_2, suppose the reaction goes to completion as written, how many gram of $TiCl_4$ can be produced ? (Ti = 48)

PROBLEMS RELATED WITH MIXTURE

21. One gram of an alloy of aluminium and magnesium when heated with excess of dil. HCl forms magnesium chloride, aluminium chloride and hydrogen. The evolved hydrogen collected at 0°C has a volume of 1.12 litres at 1 atm pressure. Calculate the composition of (% by mass) of the alloy.

22. A sample containing only $CaCO_3$ and $MgCO_3$ is ignited to CaO and MgO. The mixture of oxides produced weight exactly half as much as the original sample. Calculate the percentages of $CaCO_3$ and $MgCO_3$ (by mass) in the sample.

23. Determine the percentage composition (by mass) of a mixture of anhydrous sodium carbonate and sodium bicarbonate from the following data:

wt. of the mixture taken = 2g

Loss in weight on heating = 0.124 gm.

Fill your answer as sum of digits (excluding decimal places) till you get the single digit answer.

24. 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.

25. 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 , PERCENTAGE PURITY , DEGREE OF DISSOCIATION

26. 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 ?

27. Calculate the percent loss in weight after complete decomposition of a pure sample of potassium chlorate.

$$KClO_3(s) \longrightarrow KCl(s) + O_2(g)$$

28. A sample of calcium carbonate is 80% pure, 25 gm of this sample is treated with excess of HCl.How much volume of CO_2 will be obtained at 1 atm & 273 K?

29. 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 ?

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

30. 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 ?

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

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

32. The vapour density of a sample of N_2O_4 gas is 35. What percent of N_2O_4 molecules are dissociated into NO_2 ? .

33. If a sample of pure SO_3 gas is heated to 600ºC, it dissociates into SO_2 and O_2 gases upto 50%. What is the average molar mass of the final sample.

34. When silent electric discharge is passed through O_2 gas, it converts into O_3. If the density of final sample is 20 times the density of hydrogen gas under similar conditions, calculate the mass percent of O_2 in the final sample.

35. When acetylene (C_2H_2) gas is passed through red hot iron tube, it trimerises into benzene (C_6H_6) vapours. If the average molar mass of vapours coming out through the tube is 50, calculate the degree of trimerisation of acetylene.

SEQUENTIAL & PARALLEL REACTIONS

36. $Br_2(l)$ reacts with $Cl_2(g)$ to form BrCl and $BrCl_3$, simultaneously. How many moles of $Cl_2(g)$ reacts completely with 3 moles of $Br_2(l)$ to form BrCl and $BrCl_3$ in 5 : 1 molar ratio

37. When 80 gm CH_4 is burnt completely, CO and CO_2 gases are formed in 1 : 4 mole ratio. What is the mass of O_2 gas used in combustion.

38. Sulphur trioxide may be prepared by the following two reactions :

$$S_8 + 8O_2(g) \rightarrow 8SO_2(g)$$
$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

How many grams of SO_3 will be produced from 1 mol of S_8?

39. Potassium superoxide, KO_2, is utilised in closed system breathing apparatus. Exhaled air contains CO_2 and H_2O, both of which are removed and the removal of water generates oxygen for breathing by the reaction

$4KO_2(s) + 2H_2O(l) \rightarrow 3O_2(g) + 4KOH(s)$

The potassium hydroxide removes CO_2 from the apparatus by the reaction :

$KOH\ (s) + CO_2\ (g) \rightarrow KHCO_3(s)$

(a) What mass of KO_2 generates 24 gm of oxygen ?

(b) What mass of CO_2 can be removed from the apparatus by 100 gm of KO_2 ?

40. In a determination of P an aqueous solution of NaH_2PO_4 is treated with a mixture of ammonium and magnesium ions to precipitate magnesium ammonium phosphate $Mg(NH_4)PO_4.\ 6H_2O$. This is heated and decomposed to magnesium pyrophosphate, $Mg_2P_2O_7$ which is weighed. A solution of NaH_2PO_4 yielded 1.11 g of $Mg_2P_2O_7$. What weight of NaH_2PO_4 was present originally ? (P = 31)

PERCENTAGE COMPOSITION, EMPERICAL AND MOLECULAR FORMULA

41. Haemoglobin contains 0.25% iron by mass. The molecular mass of of Haemoglobin is 89600 then the number of iron atoms per molecule of Haemoglobin (Atomic mass of Fe = 56) -

42. 1.6 g of an organic compound containing sulphur, when treated with series of reagents, produces H_2SO_4 which on reaction with $BaCl_2$ produces 0.233 g of $BaSO_4$. Calculate % by mass of S in the organic compound. **(Given : Atomic weight of Ba = 137)**

43. 90 g of a silver coin was dissolved in strong nitric acid, and excess of sodium chloride solution was added. The silver chloride precipitate was dried and weighed 71.75 g. Calculate the percentage of silver in the coin (Atomic mass of Ag = 108)

$Ag + 2HNO_3 \longrightarrow AgNO_3 + NO_2 + H_2O$

$AgNO_3 + NaCl \longrightarrow AgCl + NaNO_3$

44. When 2.0 gm of an organic compound is burnt completely, 150 ml N_2 gas at 27ºC and 0.821 atm is obtained. The mass percent of nitrogen in the compound is

45. A polystyrene of formula $Br_3C_6H_2(C_8H_8)_n$ was prepared by heating styrene with tribromobenzyl peroxide in the absence of air. It was found to contain 10.46% bromine by weight. Find the value of n. (Br = 80).

46. 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

47. 0.5 g of NaOH is required by 0.4 gm of a polybasic acid H_nA (Molecular weight = 96gm) for complete neutralization. Value of 'n' would be : (Assume all H atom are replaced)–

48. 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 -

49. 6 gm nitrogen on successive reaction with different compounds gets finally converted into 30 gm $[Cr(NH_3)_xBr_2]$ Value of x is [Atomic mass of Cr = 52, Br = 80]

50. 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.

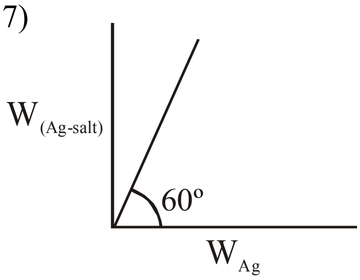
EXPERIMENTAL DETERMINATION OF ATOMIC & MOLECULAR MASSES

51. 0.80g of the chloroplatinate of a monoacidic acid base on ignition gave 0.24g of Pt. Calculate the molecular weight of the base.

53. In an experiment 3.0 g of pure $MOCl_3$ was allowed to undergo a set of reactions as a result of which all the Cl was converted to AgCl. The weight of AgCl was 7.175 g. Find atomic weight of M.
(Atomic weight of Ag = 108)

54. 0.4 gm vapours of a volatile liquid exactly displaces 82.1 cm^3 of air measured over water at 785 torr and 27ºC. The aqueous tension at 27ºC is 25 torr. The molar mass of volatile liquid is :

55. For calculation of molecular weight of a dibasic acid using silver salt method, when a graph is plotted between weight of silver salt and weight of Ag, a straight line is obtained as shown. Find molecular weight of acid. (Take $\sqrt{3}$ = 1.7)



EXERCISE # S-II

1. Sodium chlorate, $NaClO_3$, can be prepared by the following series of reactions:

$2KMnO_4 + 16\ HCl \rightarrow 2\ KCl + 2\ MnCl_2 + 8H_2O + 5\ Cl_2$

$6Cl_2 + 6\ Ca(OH)_2 \rightarrow Ca(ClO_3)_2 + 5\ CaCl_2 + 6H_2O$

$Ca(ClO_3)_2 + Na_2SO_4 \rightarrow CaSO_4 + 2\ NaClO_3$

What mass of $NaClO_3$ can be prepared from 100 ml of concentrated HCl (density 1.28 gm/ml and 36.5% by mass)? Assume all other substances are present in excess amounts.

2. Two substance P_4 & O_2 are allowed to react completely to form mixture of P_4O_6 & P_4O_{10} leaving none of the reactants. Using this information calculate the moles of P_4O_6 and P_4O_{10} in the final mixture when the following amounts of P_4 & O_2 are taken.

$$P_4 + 3O_2 \longrightarrow P_4O_6$$
$$P_4 + 5O_2 \longrightarrow P_4O_{10}$$

(i) If 1 mole P_4 & 4 mole of O_2

(ii) If 3 mole P_4 & 11 mole of O_2

(iii) If 3 mole P_4 & 13 mole of O_2

3. By the reaction of carbon and oxygen, a mixture of CO and CO_2 is obtained. What is the composition (% by mass) of the mixture obtained when 20 grams of O_2 reacts with 12 grams of carbon ?

4. Nitrogen (N), phosporus (P), and potassium (K) are the main nutrients in plant fertilizers. According to an industry convention, the numbers on the label refer to the mass % of N, P_2O_5, and K_2O, in that order. Calculate the N : P : K ratio of a 30 : 10 : 10 fertilizer in terms of moles of each elements, and express it as x : y : 1.0. (P = 31, K = 39)

5. A 10 g sample of a mixture of calcium chloride and sodium chloride is treated with Na_2CO_3 to precipitate calcium as calcium carbonate. This $CaCO_3$ is heated to convert all the calcium to CaO and the final mass of CaO is 1.12gm. Calculate % by mass of NaCl in the original mixture.

6. A mixture of Ferric oxide (**Fe_2O_3**) and **Al** is used as a solid rocket fuel which reacts to give **Al_2O_3** and **Fe**. No other reactants and products are involved. On complete reaction of 1 mole of **Fe_2O_3**, 200 units of energy is released.

(a) Write a balance reaction representing the above change.

(b) What should be the ratio of masses of Fe_2O_3 and Al taken so that maximum energy per unit mass of fuel is released.

(c) What would be energy released if 16 kg of Fe_2O_3 reacts with 2.7 kg of Al.

7. 1 gm sample of $KClO_3$ was heated under such conditions that a part of it decomposed according to the equation (1) $2KClO_3 \longrightarrow 2\ KCl + 3O_2$

and remaining underwent change according to the equation.

(2) $4KClO_3 \longrightarrow 3\ KClO_4 + KCl$

If the amount of O_2 evolved was 112 ml at 1 atm and 273 K., calculate the % by weight of $KClO_4$ in the residue.

8. 5.33 mg of salt $[Cr(H_2O)_5Cl].Cl_2.\ H_2O$ is treated with excess of $AgNO_3$(aq.) then mass of AgCl precipitate obtained will be :

Given : [Cr = 52, Cl = 35.5, Ag = 108]

9. If mass % of oxygen in monovalent metal carbonate is 48%, then find the number of atoms of metal present in 5mg of this metal carbonate sample is ($N_A = 6.0 \times 10^{23}$)

10. To find formula of compound composed of A & B which is given by A_xB_y, it is strongly heated in oxygen as per reaction-

$A_xB_y + O_2 \rightarrow AO$ + Oxide of B

If 2.5gm of A_xB_y on oxidation gives 3gm oxide of A, Find empirical formula of A_xB_y,

[Atomic mass of A = 24 & B = 14]

11. Calculate maximum mass of $CaCl_2$ produced when 2.4×10^{24} atoms of calcium is taken with 96 litre of Cl_2 gas at 380 mm pressure and at 27ºC.

[R : 0.08 atm L/mole-K & $N_A = 6 \times 10^{23}$]

12. $P_4S_3 + 8O_2 \longrightarrow P_4O_{10} + 3SO_2$

Calculate mass of P_4S_3 is required to produce at least 9.6 gm of each product.

13. Consider the given reaction

$H_4P_2O_7 + 2NaOH \rightarrow Na_2H_2P_2O_7 + 2H_2O$

If 534 gm of $H_4P_2O_7$ is reacted with 3.0×10^{24} formula units of NaOH, then total number of moles of H_2O produced is ($N_A = 6 \times 10^{23}$)

Comprehension based on "Law of Conservation of Mass" (14 & 15)

It states that matter can neither be created nor destroyed.

This law was put forth by Antoine Lavoisier in 1789. He performed careful experimental studies for combustion reactions for reaching to the above conclusion. This law formed the basis for several later developments in chemistry. Infact, this was the result of exact measurement of masses of reactants and products, and carefully planned experiments performed by Lavoisier.

14. What weight of silver nitrate will react with 0.585g NaCl to produce 1.435 g AgCl and 0.85g $NaNO_3$

15. 6.3 g sodium bicarbonate is added to 15g acetic acid solution. CO_2 formed is allowed to escape. The weight of the solution left is 18 gram. What is the mass of CO_2 formed.

Comprehension based on "Law of Definite Proportions" (16 & 17)

This law was given by, a French chemist, Joseph Proust. He stated that a given compound always contains exactly the same proportion of elements by weight.

Proust worked with two samples of cupric carbonate —one of which was of natural origin and the other was synthetic one. He found that the composition of elements present in it was same for both the samples as shown below :

	% of copper	% of oxygen	% of carbon
Natural Sample	51.35	9.74	38.91
Synthetic Sample	51.35	9.74	38.91

Thus, irrespective of the source, a given compound always contains same elements in the same proportion. The validity of this law has been confirmed by various experiments. It is sometimes also referred to as **Law of constant composition**.

Limitation :

The law is not applicable if the compound is formed from different isotopes of an element. The two isotopes of carbon C-12 and C-14 form carbondioxide $C^{12}O_2$ and $C^{14}O_2$. The ratio of C : O is 12 : 32 and 14 : 32 respectively. It is not a constant ratio.

16. 0.5gram silver is dissolved in excess of nitric acid. This solution is treated with excess of NaCl solution when 0.66g AgCl is formed. One gram metallic silver wire is heated in dry Cl_2, 1.32 g AgCl is formed. Show that these data confirm the law of constant proportion.

17. 6.488g lead reacts with 1.002 g oxygen to form an oxide. This oxide is also obtained by heating $Pb(NO_3)_2$, It is found that % of lead in this oxide is 86.62. Show that these date illustrate the law of definite proportions.

Comprehension based on "Law of Multiple Proportions" (18 to 20)

This law was proposed by Dalton in 1803. According to this law, if two elements combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers.

For example, hydrogen combines with oxygen to form two compounds, namely, water and hydrogen peroxide.

$$\begin{array}{lcl} \text{Hydrogen} + \text{Oxygen} & \rightarrow & \text{Water} \\ \quad 2g \qquad\quad 16g & & 18g \\ \text{Hydrogen} + \text{Oxygen} & \rightarrow & \text{Hydrogen Peroxide} \\ \quad 2g \qquad\quad 32g & & 34g \end{array}$$

Here, the masses of oxygen (i.e. 16 g and 32 g) which combine with a fixed mass of hydrogen (2g) bear a simple ratio, i.e. 16:32 or 1: 2.

18. N_2 and O_2 combine to form a number of compounds such as NO,N_2O, N_2O_3 , N_2O_5. Show that these follow the law of multiple proportion.

19. An element forms two oxides. In one oxide, one gram of the oxide contains 0.5 g of the element. In another oxide, 4g of the oxide contains 0.8g of the element. Show that these data confirm the law of multiple proportion.

20. 0.11g of an oxide of nitrogen gives 56mL N_2 at 273K and 1atm. 0.15g of another oxide of nitrogen gives 56 mL N_2 at 1atm, 273K. Show that these data confirm the law of multiple proportion.

EXERCISE # O-I

Single Correct :

1. One atomic mass unit in kilogram is

(A) $1/N_A$ (B) $12 / N_A$ (C) $1/1000\ N_A$ (D) $1000 / N_A$

2. The number of electron in 3.1 mg NO_3^- is ($N_A = 6 \times 10^{23}$)

(A) 32 (B) 1.6×10^{-3} (C) 9.6×10^{20} (D) 9.6×10^{23}

3. A gaseous mixture contains CO_2 (g) and N_2O (g) in a 2 : 5 ratio by mass. The ratio of the number of molecules of CO_2 (g) and N_2O (g) is

(A) 5 : 2 (B) 2 : 5 (C) 1 : 2 (D) 5 : 4

4. 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}

5. The number of hydrogen atoms in 0.9 gm glucose, $C_6H_{12}O_6$, is same as

(A) 0.048 gm hydrazine, N_2H_4 (B) 0.17 gm ammonia, NH_3

(C) 0.30 gm ethane, C_2H_6 (D) 0.03 gm hydrogen, H_2

6. Ethanol, C_2H_5OH, is the substance commonly called alcohol. The density of liquid alcohol is 0.7893 g/ml at 293 K. If 1.2 mole of ethanol are needed for a particular experiment, what volume of ethanol should be measured out?

(A) 55 ml (B) 58 ml (C) 70 ml (D) 79 ml

7. 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}

8. 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%

9. 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}

10. 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

11. 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

12. If 1/2 moles of oxygen combine with aluminium to form Al_2O_3 then weight of Aluminium metal used in the reaction is (Al= 27)–
(A) 27 g (B) 18 g (C) 54 g (D) 40.5 g

13. 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

14. 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

15. 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

16. 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

17. The minimum mass of mixture of A_2 and B_4 required to produce at least 1 kg of each product is :
(Given At. mass of 'A' = 10 ; At. mass of 'B' = 120)
$$5A_2 + 2B_4 \longrightarrow 2AB_2 + 4A_2B$$
(A) 2120 gm (B) 1060 gm (C) 560 gm (D) 1660 gm

18. 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

19. The mass of P_4O_{10} produced if 440 gm of P_4S_3 is mixed with 384 gm of O_2 is
$P_4S_3 + O_2 \longrightarrow P_4O_{10} + SO_2$
(A) 568 gm (B) 426 gm (C) 284 gm (D) 396 gm

20. 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

21. 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

22. 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

23. 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%

24. A mixture of Li_2CO_3 and Na_2CO_3 is heated strongly in an open vessel. If the loss in mass of mixture is $\frac{220}{9}\%$. The molar ratio of Li_2CO_3 and Na_2CO_3 in the initial mixture is (Li = 7, Na = 23)

(A) 1 : 1 (B) 1 : 2 (C) 2 : 1 (D) 11 : 4

25. 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

26. 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%

27. 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

28. 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%

29. The vapour density of sample of partially decomposed cyclobutane (C_4H_8) gas is 20. The degree of dissociation of C_4H_8 into C_2H_4 gas is -

(A) 0.25 (B) 0.50 (C) 0.60 (D) 0.40

30. 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) $\frac{21}{17}$

31. The density of a sample of SO_3 gas is 2.5 g/L at 0°C and 1 atm. It's degree of dissociation into SO_2 and O_2 gases is -

(A) $\frac{6}{7}$ (B) $\frac{1}{7}$ (C) $\frac{3}{7}$ (D) $\frac{5}{7}$

32. 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 %

33. Polyethene can be prepared by CaC_2 by the following sequence of reactions.

$CaC_2 + H_2O \rightarrow CaO + C_2H_2$

$C_2H_2 + H_2 \rightarrow C_2H_4$

$nC_2H_4 \rightarrow \underset{\text{(Polyethene)}}{(C_2H_4)_n}$

The mass in kg of polythene that can be prepared by 20 kg CaC_2.

(A) 4.1 kg (B) 8.75 kg (C) 3.78 kg (D) 10 kg

34. 25.4 gm of iodine and 14.2 gm of chlorine are made to react completely to yield a mixture of ICl and ICl_3. Ratio of moles of ICl & ICl_3 formed is (Atomic mass : I = 127, Cl = 35.5)

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

35. One commercial system removes SO_2 emission from smoke at 95°C by the following set of reactions-

$SO_2(g) + Cl_2(g) \rightarrow SO_2Cl_2(g)$

$SO_2Cl_2 + 2H_2O \rightarrow H_2SO_4 + 2HCl$

$H_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4 + 2H_2O$

Assuming the process to be 95% efficient. How many moles of $CaSO_4$ may be produced from 128g SO_2. [Ca = 40, S- 32, O-16]

(A) 1.9 moles (B) 2 mol (C) 3.8 mol (D) 0.95 mol

36. Equal masses of $KClO_3$ undergoes different reaction in two different container :

(i) $2KClO_3 \longrightarrow 2KCl + 3O_2$

(ii) $4KClO_3 \longrightarrow KCl + 3KClO_4$

Mass ratio of the KCl produced in respective reaction is x : 1. Value of '**x**' will be.

(A) 4 (B) 2 (C) 0.25 (D) 3

37. 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

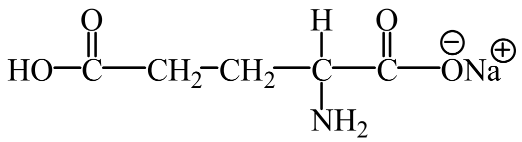
38. 13.4 gm of a sample of unstable hydrated salt : $Na_2SO_4.xH_2O$ was strongly heated. Weight loss on heating is found to be equal to 6.3 gm. Calculate the value of x.

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

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

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

40. Monosodium glutamate (MSG) is salt of one of the most abundant naturally occuring non-essential amino acid which is commonly used in food products like in **"MAGGI"** having structural formula as



Mass % of Na in MSG is-

(A) 14.8 (B) 15.1 (C) 13.6 (D) 16.5

41. 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,$C_6H_{12}O_6$

42. 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

43. 1 lt. of a hydrocarbon weighs as much as one litre of CO_2. The molecular formula of the hydrocarbon is -

(A) C_3H_8 (B) C_2H_6 (C) C_2H_4 (D) C_3H_6

44. 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

45. Two oxides of a metal contains 50% and 40% of a metal respectively. The formula of the first oxide is MO.Then the formula of the second oxide is -

(A) MO_2 (B) M_2O_3 (C) M_2O (D) M_2O_5

46. 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

47. The chloroplatinate of diacidic base contains 39% platinum. What is the molecular weight of the base. (Pt = 195) :

(A) 70 (B) 80 (C) 90 (D) 100

48. Vapour density of a volatile substance w.r.t CH_4 is 4 (CH_4 = 1). Its molecular weight would be –

(A) 8 (B) 32 (C) 64 (D) 128

49. 510 mg of a liquid on vapourisation in Victor Mayer's apparatus displaces 67.2 cm^3 of dry air at 0ºC & 1 atm . The molecular weight of liquid is -

(A) 130 (B) 17 (C) 1700 (D) 170

50. One gram of the silver salt of an organic dibasic acid yields, on strong heating, 0.6 g of silver aproximately. Determine the molecular formula of the acid. [Atomic weight of Ag = 108]

(A) $C_4H_6O_4$ (B) $C_4H_6O_6$ (C) $C_2H_6O_2$ (D) $C_5H_6O_5$

EXERCISE # O-II

1. A sample of iron ore, weighing 0.700g, is dissolved in nitric acid. The solution is then diluted with water, following with sufficient concentrated aqueous ammonia, to quantitative precipitation the iron as $Fe(OH)_3$. The precipitate is filtered, ignited and weighed as Fe_2O_3. If the mass of the ignited and dried preciipitate is 0.541g, what is the mass percent of iron in the original iron ore sample (Fe = 56)

(A) 27.0 % (B) 48.1 % (C) 54.1 % (D) 81.1 %

2. A sample of pure Cu (4.00g) heated in a stream of oxygen for some time, gains in weight with the formation of black oxide of copper (CuO). The final mass is 4.90 g. What percent of copper remains unoxidized (Cu = 64)

(A) 90 % (B) 10 % (C) 20 % (D) 80 %

Assertion Reason:

3. **Statement -1** : $2A + 3B \longrightarrow C$

4/3 moles of 'C' are always produced when 3 moles of 'A' & 4 moles of 'B' are added.

Statement -2 : 'B' is the limiting reactant for the given data.

(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.

4. **Assertion** : During a chemical reaction, total moles remains constant.

Reason : During a chemical reaction, total mass remains constant.

(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 true, statement-2 is false.

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

MULTIPLE CORRECT :

5. 40 gm of a carbonate of an **alkali metal** or **alkaline earth metal** containing some inert impurities was made to react with excess HCl solution. The liberated CO_2 occupied 12.315 lit. at 1 atm & 300 K. The correct option is

(A) Mass of impurity is 1 gm and metal is Be
(B) Mass of impurity is 3 gm and metal is Li
(C) Mass of impurity is 5 gm and metal is Be
(D) Mass of impurity is 2 gm and metal is Mg

6. 1 mole of H_2SO_4 will exactly neutralise :

(A) 2 mole of ammonia (B) 1 mole of $Ba(OH)_2$

(C) 0.5 mole of $Ca(OH)_2$ (D) 2 mole of KOH

7. 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.

8. 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

9. 'A' reacts by following two parallel reactions to give B & C If half of 'A' goes into reaction I and other half goes to reaction-II. Then , select the correct statement(s)

$A + N \xrightarrow{I} B + L$

$A + N \xrightarrow{II} \frac{1}{2}B + \frac{1}{2}(C) + L$

(A) B will be always greater than C

(B) If 2 mole of C are formed then total 2 mole of B are also formed

(C) If 2 mole of C are formed then total 4 mole of B are also formed

(D) If 2 mole of C are formed then total 6 mole of B are also formed

10. 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.

11. When N_2(g) and H_2(g) are mixed N_2H_4 (g) , NH_3(g) or both may form, depending upon the relative amount of N_2 and H_2 taken. If initial moles of N_2, H_2 are x, y and final moles of N_2H_4, NH_3 are z, v, then the correct options from the following in order of (x, y , z, v) is/are -

Match the column :

12. One type of artifical diamond (commonly called YAG for yttrium aluminium garnet) can be represented by the formula $Y_3Al_5O_{12}$.**[Y = 89, Al =27]**

Column I		Column II
Element		**Weight percentage**
(A) Y	(P)	22.73%
(B) Al	(Q)	32.32%
(C) O	(R)	44.95%

13. 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}

Matching list type :

14.

Column-I		Column-II (mass of product)
(P) $\underset{1g}{2H_2} + \underset{1g}{O_2} \rightarrow 2H_2O$	(1)	1.028 g
(Q) $\underset{1g}{3H_2} + \underset{1g}{N_2} \rightarrow 2NH_3$	(2)	1.333 g
(R) $\underset{1g}{H_2} + \underset{1g}{Cl_2} \rightarrow 2HCl$	(3)	1.125 g
(S) $\underset{1g}{2H_2} + \underset{1g}{C} \rightarrow CH_4$	(4)	1.214 g

Code :

	P	**Q**	**R**	**S**
(A)	3	4	1	2
(B)	2	4	1	3
(C)	4	3	1	2
(D)	2	3	1	4

Paragraph for Q.15 to Q.17

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)

15. 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

16. 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

17. 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

Paragraph for Q.18 to Q.19

Preparation of cobalt Metaborate involves the following steps of reactions:

(i) $Ca_2B_6O_{11} + Na_2CO_3\,(aq) \xrightarrow{\text{Boiled}} CaCO_3\,(\text{insoluble}) + Na_2B_4O_7 + NaBO_2$

(ii) $Na_2B_4O_7 \xrightarrow{\Delta} NaBO_2 + B_2O_3$

(iii) $CoO + B_2O_3 \xrightarrow{\Delta} Co(BO_2)_2.$

(Atomic weight : B = 11, Co = 59)

18. Mass of $Ca_2B_6O_{11}$ in **kg** required to produce 14.5 kg of $Co(BO_2)_2$, assuming 100% yield of each reaction is
(A) 32.2 (B) 100 (C) 28.2 (D) 30

19. If the yield of reaction (i), (ii) & (iii) is 60%, $\frac{200}{3}$ % & 32.2 % respectively, then mass of $Ca_2B_6O_{11}$ in **kg** required to produce 14.5 kg of $Co(BO_2)_2$ is
(A) 250 (B) 200 (C) 190 (D) 150

Paragraph for Q.20 to Q.22

Water is added to 3.52 gram of UF_6. The products are 3.08 gram of a solid [containing only U, O and F] and 0.8 gram of a gas only. The gas [containing fluorine and hydrogen only], contains 95 % by mass fluorine. (U = 238, F = 19). **[Assume that the empirical formula is same as molecular formula.]**

20. The empirical formula of the gas is
(A) HF_2 (B) H_2F (C) HF (D) HF_3

21. The empirical formula of the solid product is
(A) UF_2O_2 (B) UFO_2 (C) UF_2O (D) UFO

22. The percentage of fluorine of the original compound which is converted into gaseous compound is
(A) 66.66 % (B) 33.33 % (C) 50 % (D) 89.9 %

	Column-I		**Column-II**		**Column-III**
(P)	60 gram sample of hydro carbon that contain 20% H and rest C	(1)	%C = 40	(i)	No. of atoms of C and O = $8N_A$
(Q)	240 gram urea	(2)	$\%H = \frac{20}{3}$	(ii)	No. of C atoms = $4N_A$
(R)	120 gram acetic acid	(3)	$\%O = \frac{160}{3}$	(iii)	No. of total atoms = $16N_A$
(S)	120 gram glucose	(4)	% N = 46.7	(iv)	No. of total atoms is 2 times of no. of H atom

23. Out of below correct matching is -
(A) P – 1 – i (B) P – 1 – ii (C) Q – 2 – iii (D) S – 2 – iv

24. In which of following is incorrect -
(A) Q – 2 – iv (B) R – 3 – iv (C) P – 4 – iii (D) R – 1 – ii

25. Out of below correct matching is -
(A) S – 4 – iv (B) R – 1 – ii (C) P – 4 – iii (D) P – 2 – ii

EXERCISE # J-MAINS

1. The weight of 2.01×10^{23} molecules of CO is- **[AIEEE 2002]**
(1) 9.3 g (2) 7.2 g (3) 1.2 g (4) 3 g

2. In an organic compound of molar mass 108 g mol^{-1} C, H and N atoms are present in 9 : 1 : 3.5 by weight. Molecular formula can be : **[AIEEE 2002]**
(1) $C_6H_8N_2$ (2) $C_7H_{10}N$ (3) $C_5H_6N_3$ (4) $C_4H_{18}N_3$

3. If we consider that 1/6, in place of 1/12, mass of carbon atom is taken to be the relative atomic mass unit, the mass of one mole of the substance will :- **[AIEEE-2005]**
(1) be a function of the molecular mass of the substance
(2) remain unchanged
(3) increase two fold
(4) decrease twice

4. How many moles of magnesium phosphate, $Mg_3(PO_4)_2$ will contain 0.25 mole of oxygen atoms? **[AIEEE 2006]**
(1) 3.125×10^{-2} (2) 1.25×10^{-2} (3) 2.5×10^{-2} (4) 0.02

5. In the reaction **[AIEEE-2007]**
$2Al_{(s)} + 6HCl_{(aq)} \rightarrow 2Al^{3+}{}_{(aq)} + 6Cl^{-}{}_{(aq)} + 3H_2(g)$.
(1) 6L $HCl_{(aq)}$ is consumed for every 3L $H_2(g)$ produced
(2) 33.6 L $H_{2(g)}$ is produced regardless of temperature and pressure for every mole of Al that reacts
(3) 67.2 L $H_{2(g)}$ at STP is produced for every mole of Al that reacts
(4) 11.2 L $H_{2(g)}$ at STP is produced for every mole of of $HCl_{(aq)}$ consumed

6. 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)]**
(1) MCl_2 (2) MCl_4 (3) MCl_5 (4) MCl_3

7. 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)]
(1) 3 : 1 : 1 (2) 1 : 1 : 2 (3) 3 : 1 : 2 (4) 1 : 1 : 1

8. 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]**
(1) C_2H_4 (2) C_3H_4 (3) C_6H_5 (4) C_7H_8

9. 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]**
(1) 1 : 8 (2) 3 : 16 (3) 1 : 4 (4) 7 : 32

10. 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]**
(1) 48 (2) 60 (3) 24 (4) 36

11. 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]**

(1) 15 kg (2) 37.5 kg (3) 7.5 kg (4) 10 kg

12. 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]**

(1) 1186 (2) 84.3 (3)118.6 (4) 11.86

13. 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)]

(1) C_2H_4O (2) $C_3H_4O_2$ (3) $C_2H_4O_3$ (4) $C_3H_6O_3$

14. 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)

(1) $2NH_4NO_3(s) \rightarrow 2N_2(g) + 4H_2O(g) + O_2(g)$

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

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

(4) $2NH_3(g) \rightarrow N_2(g) + 3H_2(g)$

15. 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)]**

(1) 6.023×10^{21} (2) 6.023×10^{23} (3) 6.023×10^{20} (4) 6.023×10^{9}

EXERCISE # J-ADVANCE

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) $\frac{1}{9.108 \times 6.023} \times 10^{8}$

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]

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

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)

$$(NH_4)_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4.2H_2O + 2NH_3$$

$$NiCl_2 \cdot 6H_2O + 6NH_3 \rightarrow [Ni(NH_3)_6]Cl_2 + 6H_2O$$

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]**

$PbS + O_2 \longrightarrow Pb + SO_2$

ANSWER KEY

EXERCISE # S-I

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) $6\ N_A$
2.	1.992×10^{-23} g	3.	Ans.3.2 g
4.	180 g	5.	Ans.2.436 g
6.	Ans. (948)	7.	Ans.(55)
8.	Ans. (34.05L)	9.	Ans. (2)
10.	Ans.(100)	11.	Ans.73 gm
12.	Ans.27.6 gm	13	Ans.3.6 $\times 10^{19}$
14.	Ans.63 L	15.	Ans.58.8 g
16	Ans.154 gm, 24 gm	17.	Ans.0.1185
18.	Ans.0.25 mole	19.	Ans.5.6
20.	Ans.9.5	21.	Al = 60%; Mg = 40%
22.	Ans.$CaCO_3$ = 28.4%; $MgCO_3$ = 71.6%	23.	$NaHCO_3$ = 16.8 %; Na_2CO_3 = 83.2 %
24.	Ans.75%	25.	63 % , 37%
26	Ans.12	27.	39.18%
28	Ans.4.48 litre	29	61.5 gm
30	Ans.20 gm	31.	Ans. (050)
32.	Ans. (31.4%)	33.	Ans. (64 gm)
34.	Ans. (40%)	35.	Ans. (0.72)
36.	Ans.(4)	37.	Ans.(304 gm)
38	Ans. 640	39.	Ans.(a) 71 gm (b) 61.97 gm
40.	Ans.1.20 gm	41.	Ans. (4)
42.	Ans.(2)	43.	Ans. (60)
44.	Ans.(7)	45.	Ans.19.04 $\approx (19)$]
46.	$C_6H_4Cl_2$	47	Ans (3)
48.	Ans. ($C_2H_4O_2$)	49.	Ans. (4)
50.	Ans. (6)	51.	Ans. 120 gm/mole
52.	Ans. (96)	53.	Ans.(57.5)
54.	Ans.(120 gm/ mol)	55.	Ans. (153.2)

EXERCISE # S-II

1. 14.2 gm
2. (i) 0.5 , 0.5 ; (ii) 2, 1 (iii) 1, 2
3. % CO = 65.625 ; % CO_2 = 34.375
4. 10 : 0.66 : 1
5. %NaCl = 77.8%
6. (i) $Fe_2O_3 + 2\,Al \longrightarrow Al_2O_3 + 2Fe$; (ii) 80 : 27; (iii)10,000 units
7. 59.73%
8. Ans. 5.74 mg
9. Ans. 6×10^{19}
10. Ans.(A_3B_2)
11. Ans. (222 gm)
12. 11gm
13. Ans. (5)
14. Ans (1.70 gm)
15. Ans (3.3 gm)

EXERCISE # O-I

1.	Ans.(C)	2.	Ans.(C)	3.	Ans.(B)	4.	Ans.(D)
5.	Ans.(C)	6.	Ans.(C)	7.	Ans.(C)	8.	Ans.(A)
9.	Ans.(B)	10	Ans.(C)	11.	Ans. (C)	12.	Ans.(B)
13.	Ans.(C)	14.	Ans.(B)	15.	Ans (B)	16	Ans.(B)
17.	Ans.(A)	18.	Ans.(C)	19	Ans.(B)	20.	Ans.(A)
21.	Ans (B)	22.	Ans.(B)	23.	Ans.(D)	24.	Ans. (A)
25	Ans.(C)	26.	Ans.(B)	27.	Ans. (A)	28.	Ans.(A)
29.	Ans.(D)	30	Ans.(A)	31.	Ans.(A)	32.	Ans.(D)
33.	Ans.(B)	34.	Ans.(A)	35.	Ans. (A)	36.	Ans(A)
37.	Ans(C)	38.	Ans.(C)	39.	Ans.(C)	40.	Ans.(C)
41.	Ans.(D)	42.	Ans.(C)	43.	Ans.(A)	44.	Ans(B)
45.	Ans.(B)	46.	Ans.(C)	47.	Ans.(C)	48.	Ans.(C)
49.	Ans.(D)	50.	Ans.(D)				

EXERCISE O-II

1.	Ans.(C)	2.	Ans.(B)	3	Ans.(C)	4.	Ans.(D)
5	Ans.(B)	6.	Ans.(A,B,D)	7.	Ans (B,C,D)	8.	Ans.(A,B,C)
9.	Ans. (A,D)	10.	Ans. (A, B)	11.	Ans. (A,B,D)		
12.	Ans. (A) R, (B) P, (C) Q			13	Ans. (A) R, (B) Q, (C) P		
14	Ans.(A)	15.	Ans.(B)	16.	Ans.(C)	17.	Ans.(B)
18.	Ans.(A)	19.	Ans.(A)	20.	Ans.(C)	21.	Ans.(A)
22.	Ans.(A)	23.	Ans.(D)	24.	Ans.(C)	25.	Ans.(B)

EXERCISE # J-MAINS

1.	Ans.(1)	2.	Ans.(1)	3.	Ans.(1)	4.	Ans.(1)
5.	Ans.(4)	6.	Ans.(2)	7.	Ans.(4)	8.	Ans.(4)
9.	Ans.(4)	10.	Ans.(3)	11.	Ans.(3)	12.	Ans.(2)
13.	Ans.(3)	14.	Ans.(4)	15.	Ans.(3)		

EXERCISE # J-ADVANCE

1.	Ans.(D)	2.	Ans.1008 g	3.	Ans.(B)	4.	Ans.(2992)
5.	Ans. (6.47)						