"BASIC CONCEPTS" CHAPTER NO. "1"

WHAT IS AN ATOM? WRITE BRIEF HISTORY OF ATOM.

DEFINATION:- Atom is defined as smallest particle of an element which may or may not exist independently. For example Noble gas atoms can exist independently. He, Ne, Ar etc. But hydrogen, oxygen atoms cannot exist independently. The modern researches have clearly shown that atom is further composed of subatomic particles like electrons, protons, neutrons, hyperon, neutrino, anti-neutrino etc. More than 100 particles are present in Atom.

HISTORY OF ATOM Greek philosophers thought that when matter is broken into smaller and smaller particles, finally a smallest particle is obtained which cannot be further subdivided. DEMOCRITUS called these particles ATOMS. This term has been derived from word "ATOMOS" which means INDIVISIBLE.

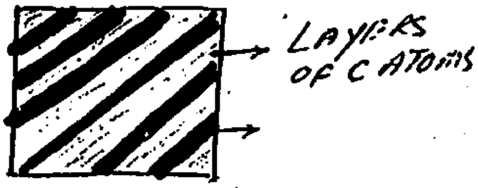
In late 17th century it was found that atoms of some elements are present in different substances. It was also discovered how atoms of different elements combine to form compounds and how to break a given compound into its elements.

ATOMIC THEORY OF MATTER:- Dalton showed that law of conservation of matter and law of constant proportions could be explained on the basis of Atoms. He developed an atomic Theory according to which all matter is composed of atoms of different elements. Atoms of different elements differ in their properties.

J. BERZELIUS determined atomic masses of elements. Berzelius also developed a system of giving symbols to each element.

EVIDENCE OF ATOMS:-

It is not possible actually to see atoms. An optical microscope can measure size of an object of size 500nm or above. The size of an atom is almost 0.2 nm. The objects of size of an atom could be observed in an electron microscope. It uses a beam of electrons instead of visible light. The wavelength of electron is much shorter than that of visible light. An electron microscope photograph of a piece of graphite is shown over here. The bright bands in fig are layers of c-atoms.



SIZE AND MASS OF ATOM:- In twentieth century X-Ray work has shown that the diameter of atoms are of the order of 2×10^{-10} meter or 0.2 nm. The Masses of atoms range from 10^{-25} kg to 10^{-27} kg. The masses of atoms are generally expressed in terms of amu

1 amu = 1.661×10^{-27} kg. A full stop can have two million atoms present in it. It gives an idea about v.small size of an atom.

WHAT IS MOLECULE? WRITE ITS ATOMICITY & CHARACTERISTICS

The smallest particle of a pure substance which can exist independently is called MOLECULE.

ATOMICITY:- The number of atoms present in a molecule is called its ATOMICITY. Thus a molecule can be

MONOATOMIC:- It contains only one atom like He, Ne, Ar

DIATOMIC:- It contains two atoms like H_2, O_2, N_2 etc

TRIATOMIC:- It contains three atoms like H_2O, CO_2 etc

POLYATOMIC:- It contains many atoms. Such molecules are also

MACROMOLECULE:- These are very large molecules containing very large number of atoms. For example HAEMOGLOBIN contains 10,000 atoms in one molecule and it is 68000 times heavier than H-Atom.

A molecule may contain atoms of same element or atoms of different elements

HOMOATOMIC MOLECULES:- H_2, O_2, P_4, S_8.

HETEROATOMIC MOLECULES:- HCl, NH_3, $C_6H_{12}O_6$, H_2SO_4.

PROPERTIES OF MOLECULES:-

1:- Molecules of same substance are similar in all respects.

2:- Molecules have empty spaces between them. These are maximum in gases and minimum in solids

3:- The molecules are in the state of constant motion. These movements are maximum in gases, lesser in liquids; solids have only vibrational motion.

4. Molecules have attractive forces for each other

5. The molecules have definite kinetic and potential energy.

6. For a chemical reaction molecules must collide with each other. The molecules exchange atoms as a result of collision and new compounds are formed

WHAT IS ION? WRITE ENERGY CHANGES DURING FORMATION OF A POSITIVE AND NEGATIVE ION

ION:- The species which carry positive or negative charge are called Ions. There are two type of ions

(1) POSITIVE ION (CATION) (2) NEGATIVE ION (ANION

POSITIVE ION Positive ion is formed when one or more electrons are removed from a neutral atom

$$A - e \longrightarrow A^{1+}$$

If an atom loses one electron, unipositive ion is formed

If an atom loses two electrons dipositive ion is formed

If an atom loses three electrons tripositive ion is formed

Energy is required to remove an electron from an atom to produce +ve ion. It is called Ionization energy. Thus formation of positive ion is an endothermic process. The most common positive ions are Na^{1+}, K^{1+}, Ca^{2+}, Mg^{2+}, Al^{3+}, Fe^{3+}, Sn^{4+} etc

NEGATIVE ION (ANION) Negative ion is formed when an atom gains one two or more electrons.

$$B + e^{\ominus} \longrightarrow B^{1-}$$

The formation of negative ion is exothermic process. The most common negative ions are $Cl^{1-}, Br^{1-}, F^{1-}, S^{2-}$ etc. Uninegative, dinegative and trinegative ions are formed due to addition of one, two or three electrons in an atom. Some negative ions consist of group of atoms For example OH^{1-}, SO_4^{2-}, PO_4^{3-}, MnO_4^{1-}, CrO_4^{2-} etc. Positive ions having group of atoms are less common e.g. NH_4^{1+} and some carbocations (carbon having +ve charge). Properties of ions are different from corresponding atoms.

WHAT IS MOLECULAR ION? WRITE ITS SIGNIFICANCE

When a molecule gains or loses electron, a molecular ion is formed. For example CH_4^{1+}, CO^+, N_2^{1+} etc. Cationic (+ve) molecular ions are more abundant than Anionic (−ve) molecular ions. This molecular ions are produced by passing high energy electron beam or α-particles through a gas.

SIGNIFICANCE:- The molecular ions are quite unstable. The breaking of molecular ion gives useful information about structure of natural products.

RELATIVE ATOMIC MASS:- The mass of an atom of an element as compared to mass of carbon atom taken as 12 is called RELATIVE ATOMIC MASS. For example on C-12 scale mass of $H = \frac{1}{12} \times 12 = 1.0078$ amu and ~~that~~ mass of carbon is 12.0000 amu. The masses of atoms are extremely small. We donot have a balance to weigh atoms

Thus we use relative atomic mass unit scale. The elements have fractional relative atomic masses due to their different Isotopic abundance.

WHAT ARE ISOTOPES? DISCUSS THEIR RELATIVE ABUNDANCE

ATOMS OF SAME ELEMENT HAVING SAME ATOMIC NUMBER BUT DIFFERENT ATOMIC WTS. ARE CALLED ISOTOPES. THIS PHENOMENON IS CALLED ISOTOPY. It was discovered by Soddy. Isotopes have same number of electrons, protons and same electronic configuration. They differ in number of NEUTRONS present in nucleus. The isotopes have same chemical properties and same position in periodic Table. Consider example of Hydrogen. It has three isotopes viz.

		PROTONS	ELECTRONS CONFIG.	NEUTRONS	Z	A
PROTIUM	^1_1H	1	1 ($1S^1$)	0	1	1
DEUTERIUM	^2_1H	1	1 ($1S^1$)	1	1	2
TRITIUM	^3_1H	1	1 ($1S^1$)	2	1	3

Thus we can define isotope as atom of same element which differs in number of Neutrons in Nucleus. Carbon has three isotopes $^{12}_6C$ $^{13}_6C$ $^{14}_6C$. Each one of them have 6-electrons and 6-protons but they have 6, 7, 8 neutron respectively.

The numbers of isotopes of some other elements are ($O = 3$, $Ni = 5$, $Ca = 6$, $Pd = 6$, $Cd = 9$, $S_n = 11$).

RELATIVE ABUNDANCE OF ISOTOPES:-

The isotopes of all elements have their own natural abundance. This relative abundance is determined from mass spectrometery. The properties of element which are written in books are that of most abundant isotope of an element.

PLEASE

There are 280 stable naturally occuring isotopes in nature. There are 40 radioactive isotopes. Besides 300 unstable radioactive isotopes have been produced by artificial radioactivity. The numbers of isotopes of an element is a complex property. Some general informations are given over here.

MONOISOTOPIC ELEMENTS:- These elements have only single isotope. For example GOLD, Iodine, flourine, Arsenic

ODD ATOMIC NUMBER:- Elements having odd atomic No. never possess more than two stable isotopes.

EVEN ATOMIC NUMBER:- Elements with even atomic No. have large number of isotopes.

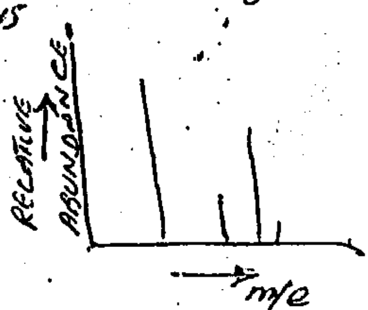
MASS NO. MULTIPLE OF FOUR:- The isotopes with mass No. which is multiple of four are quiet abundant. For example ^{16}O, ^{24}Mg, ^{32}Si, ^{40}Ca, ^{56}Fe are almost 50% of earths crust.

EVEN MASS NO. & ATOMIC NO.:- Out of 280 naturally occuring isotopes 154 have even mass Number and even atomic Number.

MASS SPECTROMETERY:- DETERMINATION OF RELATIVE ATOMIC MASSES OF ISOTOPES

Mass Spectrometer is an instrument which is used to measure the exact masses of different isotopes of an element.

BASIC PRINCIPLE:- The substance whose mass is to be determined is first changed into vapours. These vapours are then ionized in an ionization chamber, with the beam of high energy electrons. The positive ions are separated on the basis of their $\frac{m}{e}$ ratio in a magnetic analyzer. These ions are detected by ion collector. The results are represented in the form of a spectrum. m/e is plotted on X-axis and relative No. of ions along Y-axis



INSTRUMENT AND ITS WORKING:-

ASTON'S MASS SPECTROGRAPH was first of all used to identify different isotopes of an element. Dempster mass spectrometer was designed for identification of elements which are found in solid state.

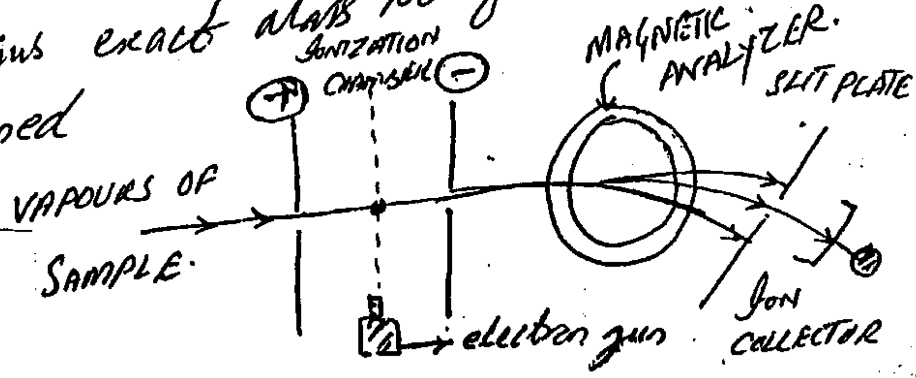
SAMPLE CHAMBER:- The sample is first changed into vapours. The pressure of these vapours is kept v. low (10^{-6} Torr – 10^{-7} Torr)

IONIZATION CHAMBER:- These vapours are allowed to enter the ionization chamber, these are then bombarded with beam of high energy electrons. The vapours are ionized. The positively charged ions have different masses depending upon nature of isotopes present in them.

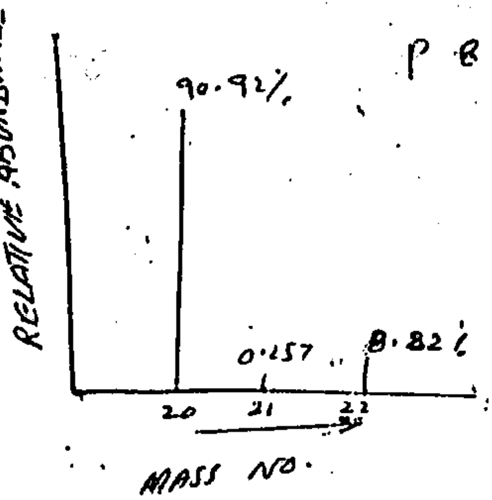
MAGNETIC ANALYZER:- These positive ions are then passed through a strong magnetic field. They are separated on the basis of their m/e value. The ions with definite $\frac{m}{e}$ value move ~~out~~ in the form of a group.

ELECTROMETER OR ION COLLECTOR:- The ions with definite $\frac{m}{e}$ value fall on electrometer one after another. The electrometer develops the electrical current. The strength of electrical current gives relative abundance of ions of definite $\frac{m}{e}$ value. Different isotopes have different $\frac{m}{e}$ value. They are allowed to fall on collector one after another and electrical current is measured. The current strength in each case gives relative abundance of each isotope.

The same experiment is performed with C-12 isotope and current produced is compared. Thus exact mass No. of each isotope can be determined



Computer plotted graph of Neon is shown over here. In modern spectrograph each ion strikes detector. The ionic current is amplified and fed to recorder. The recorder makes graph showing relative abundance of ions plotted against mass no.



SEE EXAMPLE No. 1 - (P-8) TEXT BOOK

Q: WHAT IS EMPIRICAL FORMULA AND MOLECULAR FORMULA OF COMPOUND? EXPLAIN STEPS INVOLVED IN DETERMINATION OF EMPIRICAL AND MOLECULAR FORMULA?

ANS:- The simplest formula which gives information about simple ratio of atoms of different elements present in a chemical compound is called Empirical formula.

For example empirical formula of a compound is A_xB_y then it means compound contain "Atoms of "A" and "B" elements in $x : y$

Empirical formula of Benzene = CH

" " Glucose = CH_2O

Determination of Empirical formula involves following steps

1) percentage composition of each element is determined

2) No. of gram atoms of each element are determined by dividing % composition by Atomic wt of element

3) The Atomic ratio is determined by dividing the No. of moles by smallest value

4) If atomic ratio is not in whole numbers it is converted into round numbers by multiplying with suitable integer. (See example no.3)

MOLECULER FORMULA / The formula of a compound based upon actual molecule is called molecular formula. It indicates actual No. of atoms of different elements present in 1 molecule of compound. For example molecular formula of Benzene is C_6H_6 and molecular formula of glucose is $C_6H_{12}O_6$.

Molecular formula = (Empirical formula) x n

It means molecular formula is simple multiple of empirical formula. Some compounds have same empirical formula and molecular formula e.g H_2O, CO_2, NH_3. For these compounds n = 1

$$n = \frac{\text{Molecular Wt}}{\text{Emp. Formula Wt}}$$

DIFFERENCES BETWEEN :-

EMPIRICAL FORMULA	MOLECULAR FORMULA
It is based upon formula unit	It is based upon actual molecule.
It indicates relative number of atoms of various element present in a compound	It gives exact number of atoms of various elements present in molecule of compound.
It may or may not exist independently	It represents compound in the form which can exist independently.
Ionic and covalent compounds, both have empirical formula.	Only covalent compounds have molecular formula.
The sum of atomic weights of atoms present in empirical formula is called Empirical formula weight.	The sum of atomic weights of atoms in 1 molecule is called molecular weight.

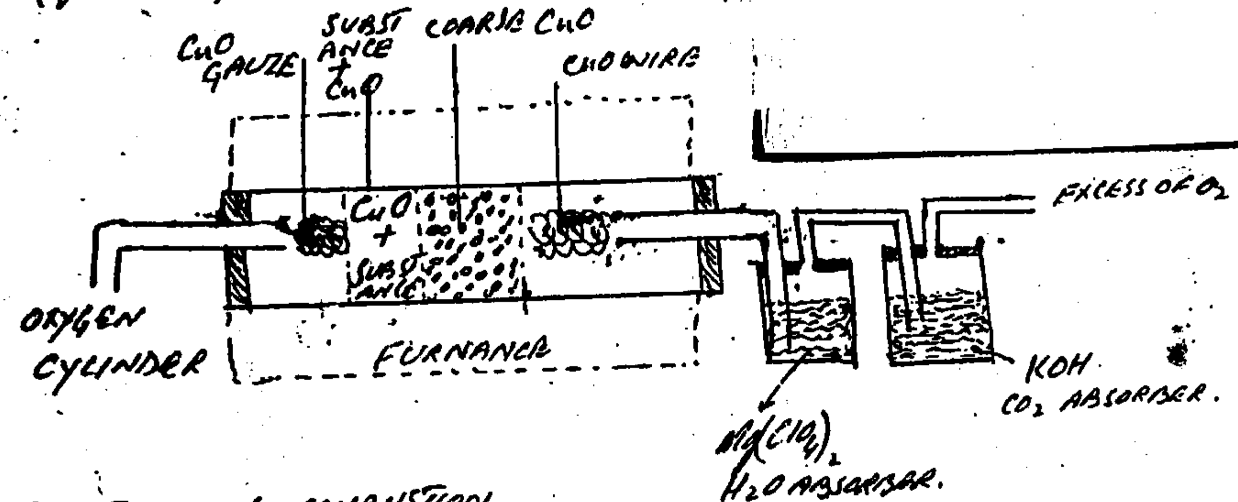
STEPS INVOLVED IN DETERMINATION OF MOLECULER FORMULA.

1) DETERMINATION OF % COMPOSITION BY CHEMICAL ANALYSIS
2) DETERMINATION OF EMPIRICAL FORMULA
3) DETERMINATION OF EMPIRICAL FORMULA WT & MOL WT
(4) MOLECULER FORMULA = (EMPIRICAL FORMULA) x n

1. DETERMINATION OF PERCENTAGE COMPOSITION:-

There are several methods of chemical analysis. Chemical analysis is experimental procedure by which amount of various elements in chemical compound is determined.

In most of organic analysis combustion process is used. In this process a weighed amount of organic compound is burned completely in stream of oxygen. The CuO placed at end of tube ensures complete combustion of organic compound to CO_2 and H_2O.



REACTION IN COMBUSTION

$$C, H \text{ IN SUBSTANCE} + O_2 \xrightarrow{\Delta} CO_2 + H_2O$$

The amount of CO_2 and water produced as a result of combustion is determined by passing these gases through weighed amount of KOH and $Mg(ClO_4)_2$ respectively. KOH is CO_2 absorber and $Mg(ClO_4)_2$ is H_2O absorber. The increase in weights of these 2 compounds after complete combustion is due to CO_2 and H_2O produced during reaction. Some other absorbers can be used to determine weights of other compounds produced as a result of combustion.

The % of "C", "H" and oxygen can be determined as follows.

$$\% \text{ of Carbon in O.C} = \frac{\text{wt. of } CO_2}{\text{wt of org. Compound}} \times \frac{12}{44} \times 100$$

$$\% \text{ of HYDROGEN in org. Compound} = \frac{\text{wt. of } H_2O}{\text{wt of org. Comp}} \times \frac{2}{18} \times 100$$

If organic compound contains only hydrogen, carbon and oxygen then % of oxygen is

% of oxygen = 100 — (% of Carbon + % of Hydrogen).

If organic compound contains some other elements as well then % of each element is determined, the % of oxygen is = 100 − Sum of % of all elements present in O.C. There is no direct method of determining % of oxygen.

2) DETERMINATION OF EMPIRICAL FORMULA

The determination of empirical formula from % composition involves following steps.

(1) MOLE RATIO:- The % of each element is divided by its atomic wt to get mole ratio of that element.

(2) ATOMIC RATIO:- The mole ratio is divided by smallest figure (among mole ratios) to get atomic ratio.

(3) WHOLE NUMBER:- The atomic ratio is converted into nearest whole number either by rounding it off to whole No. or by multiplying it by suitable Integer.

Thus empirical formula is obtained.

3) EMPIRICAL FORMULA WEIGHT

The empirical formula wt of compound is determined by adding atomic wts of all atoms present in empirical formula. For example empirical formula of glucose is CH_2O. Its empirical formula weight is = 12 + 2 + 16 = 30.

4) DETERMINATION OF MOLECULAR FORMULA (CHEMICAL FORMULA).

Molecular formula (Chemical formula) = (Empirical formula) × n

$$n = \frac{\text{Molecular wt}^*}{\text{Empirical formula wt}}$$

*{Molecular wt of compound can be determined by elevation of B.P. or depression in F.P. methods}

WHAT IS MOLE? HOW TO CONVERT A QUANTITY INTO MOLES:-

DEFINATION:- The atomic mass, molecular mass or formula mass or ionic mass of substance expressed in grams is called MOLE of that substance.

$$\text{MOLE} = \frac{\text{wt. of substance in grams}}{\text{At wt or M wt or formula wt of substance}}$$

MOLES OF ELEMENT:- If substance is an element which exists in atomic form then atomic wt of that element expressed in grams is equal to 1 mole

At wt of CARBON = 12 amu = 12 g Carbon = 1 mole

$$\text{MOLES OF ELEMENT} = \frac{\text{wt of element in grams}}{\text{Atomic wt of element}}$$

If element exists in molecules form then molecular wt expressed in grams is 1 mole

Oxygen has molecular wt = 32 = 32 g O_2 = 1 mole.

MOLES OF MOLECULAR COMPOUND The compounds which exist in molecules form, their molecules wt in grams is equal to 1 mole

For example M wt of water = 18

18 g of water is equal to 1 mole

$$\text{moles of molecules Substance} = \frac{\text{wt in gram's of molecules Sub.}}{\text{M wt of that substance}}$$

Substance	M wt	1 mole
H_2O	18	18 g
$C_6H_{12}O_6$	180	180 g
Sucrose	342	342 g.

Thus 1 mole of different substances have different wts.

IONIC SUBSTANCES The formula wt of ionic substances expressed in gram is also called 1 mole

$$\text{Mole of Ionic substance} = \frac{\text{wt of ionic substance in g}}{\text{Formula wt of Substance}}$$

Formula wt of NaCl = 58.5.

Thus 1 mole of NaCl = 58.5 g
1 mole of $AgNO_3$ = 170 g

MOLES OF IONS / The ionic mass of ionic species expressed in grams is also called 1 MOLE.

$$\text{Moles of ions} = \frac{\text{wt in grams of ionic species}}{\text{Formula wt of ionic species}}$$

For example 1 mole of OH^- ions = 17 g
1 mole of SO_4^{2-} ions = 96 g

Thus atomic wt, M.wt, ionic mass, formula mass expressed in grams is called 1-Mole
(Please see example No 6, 7 Text Book P-15, 16)

AVOGADRO'S NUMBER / The number of Atoms, ions or molecules present in 1 mole of substance is called AVOGADRO'S NUMBER

EXAMPLES:-
1 MOLE OF He = 4 g = 6.022×10^{23} ATOMS OF He
1 MOLE OF Na = 23 g = 6.022×10^{23} ATOMS OF Na.
1 mole of $C_6H_{12}O_6$ = 180 g = 6.022×10^{23} MOLECULES.

It must be noted that Avogadro's Number of atoms of different substances have different wts. Thus an atom of sodium is 23 times heavier than one atom of hydrogen. Magnesium atom has twice as heavy than Carbon atom. Thus 10 g of Mg and 5 g of Carbon will have equal Number of Atoms. Similarly

M.wt.	substance	one mole	Molecules
18	H_2O	18 g	6.022×10^{23} molecules
180	$C_6H_{12}O_6$	180 g	6.022×10^{23} molecules.
IONIC MASS	ION	ONE MOLE	NO. of ions
96	SO_4^{2-}	96 g	6.022×10^{23} IONS SO_4^{2-}
62	NO_3^{1-}	62 g	6.022×10^{23} IONS NO_3^{1-}

WHAT IS LIMITING REACTANT? HOW TO CALCULATE LIMITING REACTANT?

When reactants are not mixed in stoichiometric amounts one of the reactant is consumed earlier. This reactant is called LIMITING REACTANT. The other reactant is left unreacted after completion of reaction. This reactant is called EXCESSIVE REACTANT.

Since the limiting reactant is consumed earlier and reaction stops due to its consumption. Therefore amounts of product depend upon limiting reactant.

DETERMINATION: The limiting reactant is determined as follows

1) The amounts of reactants are converted into moles
2) The amount of products are calculated from molar amounts of reactants by using balanced chemical equation.
3) The reactant which produces least amount of products is called LIMITING REACTANT.

The amount of products depends upon limiting reactant.

EXAMPLE:- SODIUM THIOSULPHATE IS PRODUCED BY REACTION INDICATED BELOW

$$2Na_2S + Na_2CO_3 + 4SO_2 \longrightarrow 3Na_2S_2O_3 + CO_2$$

If 100g of each of reactant is mixed calculate amount of sodium thiosulphate produced.

SOLUTION: 1) Moles of each reactant $= \frac{\text{wt of substance}}{\text{M.wt of substance}}$

moles of $Na_2S = \frac{100}{78} = 1.28$ moles.

Moles of $Na_2CO_3 = \frac{100}{106} = 0.943$ moles

Moles of $SO_2 = \frac{100}{64} = 1.56$ moles.

[M.wt of $Na_2S = 78$, $Na_2CO_3 = 106$, $SO_2 = 64$]

2) AMOUNT OF PRODUCTS FROM EACH REACTANT

Given equation shows that

2 mole Na_2S produces $Na_2S_2O_3$ = 3 mole

1 mole Na_2S produces $Na_2S_2O_3 = \frac{3}{2}$ mole

1.28 mole Na_2S will produce $Na_2S_2O_3 = \frac{3}{2} \times 1.28$

Amount of $Na_2S_2O_3$ produced from Na_2S = 1.92 moles.

However their masses are not equal, due to different sizes and atomic masses of molecules. The sizes and masses of molecules donot effect total volume of gas because in gases empty spaces between molecules are almost 300 times greater than diameter of molecule.

(SEE EXAMPLE NO. 10)

WHAT IS STOICHIOMETERY? WRITE ITS ASSUMPTIONS AND TYPES OF RELATIONSHIPS INVOLVED?

The branch of chemistry which deals with quantitative relationship between reactants and products involved in a balanced chemical equation is called STOICHIOMETRY.

LIMITATIONS OF CHEMICAL EQUATION:-

1) The chemical equations do not tell about conditions and rate of chemical reactions.
2) A chemical equation can be written to describe a chemical change that dont occur.

ASSUMPTIONS OF STOICHIOMETRY:-

(1) All reactants are converted into the products.
2) No side reactions take place.
3) The law of conservation of mass and law of definite proportions is obeyed.

The following relationships can be studied with the help of stoichiometery.

MASS - MASS RELATIONSHIP:- If we are given mass of one substance we can calculate mass of other substance.

MASS -- MOLE AND MOLE-MASS RELATIONSHIP:- If we are given mass of one substance we can calculate mole of other, and vice versa.

MASS - VOLUME RELATIONSHIP:- If we are given the mass of one substance we can calculate volume of other substance and vice versa.

"SEE EXAMPLES NO. (11, 12)

(ii) 1 mole of Na_2CO_3 produces $Na_2S_2O_3$ = 3

0.94 moles of Na_2CO_3 produce $Na_2S_2O_3$ = 3×0.94

= **2.82 moles**

(iii) 4 mole of SO_2 produce $Na_2S_2O_3$ = 3

1 " " " " " = $\frac{3}{4}$

1.56 moles of SO_2 produce $Na_2S_2O_3$ = $\frac{3}{4}$×1.56

= **1.17 mole**

Since SO_2 produce least amount of products hence it is limiting reactant.

Thus $Na_2S_2O_3$ produced = 1.17 moles

= moles × M. wt of $Na_2S_2O_3$

weight of $Na_2S_2O_3$ = 1.17 × 158 = **184.86 g**

WHAT IS THEORETICAL YIELDS. WHY IS ACTUAL YIELD LESSER THAN THEORETICAL YIELD.

The amount of products expected to be formed from given amount of reactants, using balanced chemical equation is called Theoratical yield.

OR

"Amount of products calculated from stoichiometric calculations is called Theoratical yield."

However amount of products actually obtained after the completion of chemical reaction is less than theoratical yield. It may be due to following reasons.

(i) Formation of by-products due to side reactions.

(ii) The reaction may be reversible. Thus it does not go to completion.

(iii) Mechanical loss of products during filteration, crystallization etc.

So yield is generally expressed in terms of %.

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

(SEE EXAMPLE 14 P-27)

(FIRST YEAR)

EXERCISE CHAPTER NO. 1

Q1 Select the most suitable answer from the given ones in each question.

(i) The number of moles of CO_2 which contain 8.0 g of oxygen.

(a) 0.25
(b) 0.50
(c) 1.0
(d) 1.50

(ii) 27 g of Al will react completely with how much mass of O_2 to produce Al_2O_3.

(a) 8 g of oxygen
(b) 16 g of oxygen
(c) 32 g of oxygen
(d) 24 g of oxygen

(iii) One mole of SO_2 contains

(a) 6.02×10^{23} atoms of oxygen
(b) 18.1×10^{23} molecules of SO_2
(c) 6.023×10^{23} atoms of sulphur
(d) 4 gram atoms of SO_2

(iv) The largest number of molecules are present in

(a) 3.6 g of H_2O
(b) 4.8 g of C_2H_5OH
(c) 2.8 g of CO
(d) 5.4 g of N_2O_5

(v) The mass of one mole of electrons is

(a) 1.008 mg
(b) 0.184 mg
(c) 1.673 mg

(vi) Isotopes differ in

(a) properties which depend upon mass
(b) arrangement of electrons in orbitals
(c) chemical properties
(d) the extent to which they may be affected in electromagnetic field

(vii) The volume occupied by 1.4 g of N_2 at S.T.P is

(a) $2.24\ dm^3$
(b) $22.4\ dm^3$
(c) $1.12\ dm^3$
(d) $112\ cm^3$

(viii) Many elements have fractional atomic masses. This is because

(a) the mass of the atom is itself fractional.
(b) atomic masses are average masses of isobars.
(c) atomic masses are average masses of isotopes.
(d) atomic masses are average masses of isotopes proportional to their relative abundance.

(ix) A limiting reactant is the one which

(a) is taken in lesser quantity in grams as compared to other reactants.
(b) is taken in lesser quantity in volume as compared to the other reactants.
(c) gives the maximum amount of the product which is required.
(d) gives the minimum amount of the product under consideration.

(x) Which of the following statements is not true?

(a) isotopes with even atomic masses are comparatively abundant.
(b) isotopes with odd atomic masses are comparatively abundant
(c) isotopes with even atomic masses and even atomic numbers are comparatively abundant
(d) isotopes with even atomic masses and odd atomic numbers are comparatively abundant

Q2 Indicate true or false as the case may be:

i. Neon has three isotopes and the fourth one with atomic mass 20.18 amu F

ii. Empirical formula gives the information about the total number of atoms present in the molecule. F

iii. During combustion analysis $Mg(ClO_4)_2$ is employed to absorb water vapours T

iv. Molecular formula is the integral multiple of empirical formula and the integral multiple can never be unity. F

v. The number of electrons in the molecules of CO and N_2 are 14 each, so 1 mg of each gas will have same number of electrons. F

vi. Avogadro's hypothesis is applicable to all types of gases i.e. ideal and non-ideal. F

vii. Actual yield of a chemical reaction may be greater than the theoretical yield. F

viii. The number of atoms in 1.79 g of gold and 0.023 g of sodium are equal. F

Q3. Fill in the blanks

i. The unit of relative atomic mass is amu.

ii. The exact masses of isotopes can be determined by MASS spectrograph.

iii. The phenomenon of isotopy was first discovered by SODDY.

iv. One of the well-known elements TIN has eleven isotopes.

v. Fluorine and iodine both have ONE isotope each and are called MONOISOTOPIC

vi. Mass spectrum is a plot of data in such a way that m/e of the ions is plotted as ABSCISSA and relative number of ions as ORDINATE

vii. The m/e values of two isotopes of an element are DIFFEREN

viii. Stoichiometric calculations can be performed only when LAW OF CONS. OF MASS is obeyed.

Q.4 What are ions? Under what conditions are they produced? Can you explain the places of negative charge in PO_4^{3-}, MnO_4^-, $Cr_2O_7^{2-}$.

Q.5 (a) What are isotopes. How do you deduce the fractional atomic masses of elements from the relative isotopic abundance? Give two examples in support of your answer.

(b) How does a mass spectrograph show the relative abundance of isotopes of an element.

(c) What is the justification of two strong peaks in the mass spectrum for bromine; while for iodine only one peak at 127 amu, is indicated?

Q.6 Silver has atomic number 47 and has 16 known isotopes but two occur naturally i.e. Ag-107 and Ag-109. Given the following mass spectrometic data, calculate the average atomic mass of silver.

Isotopes	Mass (amu)	Percentage abundance
^{107}Ag	106.90509	51.84
^{109}Ag	108.90476	48.16

QNO. 6

SOLUTION :-

$$\text{AVERAGE ATOMIC MASS OF "Ag"} = \frac{106.90509 \times 51.84 + 108.90476 \times 48.16}{100}$$

$$= \frac{5541.95987 + 5244.85324}{100}$$

AVERAGE ATOMIC MASS OF "Ag" = 107.2681 amu

Q.7 Boron with atomic number 5 has two naturally occurring isotopes. Calculate the percentage abundance of ^{10}B and ^{11}B from the following information.

Average atomic mass of boron	=	10.81 amu
Isotopic mass of ^{10}B	=	10.0129 amu
Isotopic mass of ^{11}B	=	11.0093 amu

7: SOLUTION. SUPPOSE % OF ^{10}B IS "x"

THEN % OF ^{11}B IS "100 − x"

$$\text{AVERAGE ATOMIC MASS OF B} = \frac{(\text{ISOTOPIC MASS} \times \% \text{ OF } {}^{10}B) + (\text{ISOTOPIC MASS B} \times \%)}{100}$$

$$10.81 = \frac{(10.0129 \times x) + 11.0093 \times (100 - x)}{100}$$

$$= \frac{(10.0129x) + 1100.93 - 11.0093x}{100}$$

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$$1081 = 10.0129x - 11.0093x + 1100.93$$

$$1081 - 1100.93 = -0.9964x$$

$$+19.93 = +0.9964x$$

$$20.00\% = x$$

PERCENTAGE OF ^{10}B = 20.00% ANS.

" " ^{11}B = 80.00% ANS.

QNO: 10. CALCULATE EACH OF FOLLOWING QUANTITIES.

(a) MASS IN GRAMS OF 2.74 MOLES OF $KMnO_4$

MOLAR MASS OF $KMnO_4$ = 39 + 55 + 64 = 158

$$\text{MOLE} = \frac{\text{MASS IN GRAMS IN GRAMS}}{\text{MOLAR MASS OF } KMnO_4}$$

MOLE × MOLAR MASS = MASS IN GRAMS

432.92g = 2.74 × 158 = MASS IN GRAMS

Moles of O atoms in 9.00g of $Mg(NO_3)_2$.

MOLES OF $Mg(NO_3)_2$ = $\frac{\text{MASS IN GRAMS}}{\text{MOLAR MASS}}$ = ~~2~~

MOLAR MASS OF $Mg(NO_3)_2$ = $24 + (14 + 48)_2$ = 148

MOLES OF $Mg(NO_3)_2$ = $\frac{9}{148}$ = 0.0608 MOLES.

1MOLE OF $Mg(NO_3)_2$ CONTAINS MOLE ATOM OF OXYGEN = 6 MOLE

0.0608 MOLES " " " " = 6X0.0608

= 0.3648 MOLES

THUS MOLE ATOM OF OXYGEN = 0.3648 MOLE.

(C) c) Number of O atoms in 10.037 g of $CuSO_4.5H_2O$

MOLAR MASS OF $CuSO_4 \cdot 5H_2O$ = 63.5 + 32 + 64 + 5(18)

= 159.5 + 90

= 249.5 g/mole

MOLES OF $CuSO_4 \cdot 5H_2O$ = $\frac{10.037}{249.5}$ = 0.040 MOLES

1MOLE $CuSO_4 \cdot 5H_2O$ CONTAINS OXYGEN = 9 MOLE ATOM.

0.040 MOLE $CuSO_4 \cdot 5H_2O$ CONTAINS OXYGEN = 9X 0.040

= 0.36 MOLE ATOM

NUMBER OF OXYGEN ATOMS = $0.36 \times 6.022 \times 10^{23}$

ANS = 2.18×10^{23} ATOMS

(d) Mass in kilograms of 2.6×10^{20} molecules of SO_2.

MOLES OF SO_2 = $\frac{\text{NUMBER OF MOLECULES}}{\text{AVOGADROS' NUMBER}}$

= $\frac{2.6 \times 10^{20}}{6.022 \times 10^{23}}$ = 0.43×10^{-3} MOLE

MASS OF SO_2 = MOLES X MOLAR MASS

= 0.43 X 64

= 27.54 g

SO_2: 32+32 = 64

MASS IN kg. = 0.2764 kg.

e) Moles of Cl atoms in 0.822 g $C_2H_4Cl_2$.

(E) MOLE

MOLES OF $C_2H_4Cl_2$ = $\frac{\text{MASS IN GRAMS}}{\text{MOLAR MASS}}$ = $\frac{0.822}{99}$

MOLE OF $C_2H_4Cl_2$ = 8.3×10^{-3} MOLE

P22

1 MOLE $C_2H_4Cl_2$ CONTAINS Cl = 2 MOLE ATOM

8.3×10^{-3} MOLE " " Cl = $2\times8.3\times10^{-3}$ MOLE ATOM

MOLE ATOMS Cl = 1.66×10^{-2} MOLE ATOM.

= 0.0166 MOLE ATOM.

f. f) Mass in grams of 5.136 moles of silver carbonate.

ANS: MASS IN GRAMS = MOLES × FORMULA WT. OF Ag_2CO_3

= $5.136\times((108)_2+12+48)$

= 5.136×276

MASS OF GRAMS. = 1417.536 g.

g Mass in grams of 2.78×10^{21} molecules of CrO_2Cl_2.

MOLES OF CrO_2Cl_2 = $\frac{\text{NO. OF MOLECULES}}{\text{AVOGADRO'S NO.}}$

= $\frac{2.78\times10^{21}}{6.02\times10^{23}}$ = 0.46×10^{-2} MOLE.

MOLAR MASS OF CrO_2Cl_2 = 52+32+71 = 155

MASS IN GRAMS = MOLE × MOLAR MASS

= $0.46\times10^{-2}\times155$ = 0.713 g

h. Number of moles and formula units in 100 g of $KClO_3$.

Formula wt of $KClO_3$ = 39+35.5+48 = 122.5

MOLES OF $KClO_3$ = $\frac{\text{MASS IN GRAM}}{\text{FORMULA WT}}$ = $\frac{100}{122.5}$ = 0.816 mole

NUMBER OF FORMULA UNITS = MOLE × AVOGADRO'S NO.

= $0.816\times6.022\times10^{23}$

= 4.92×10^{23} FORMULA UNITS

i: Number of K^+ ions ClO_3^- ions, Cl atoms, and O atoms in (h).

1 FORMULA UNIT OF $KClO_3$ CONTAINS K^+ = 1 ION

4.92×10^{23} FORMULA UNITS CONTAIN K^+ = 4.92×10^{23}

SIMILARLY 4.92×10^{23} $KClO_3$ CONTAIN ClO_3^{1-} = 4.92×10^{23} ION

Cl ATOMS = 4.92×10^{23} ATOM

OXYGEN ATOMS = $3\times4.92\times10^{23}$

= 14.7×10^{23} ATOM.

OK. = 1.47×10^{24} ATOM

Q.11 Aspartame, the artificial sweetner, has a molecular formula of $C_{14}H_{18}N_2O_5$.

a) What is the mass of one mole of aspartame.
b) How many moles are present in 52 g of aspartame.
c) What is the mass in grams of 10.122 moles of aspartame.
d) How many hydrogen atoms are present in 2.43 g of aspartame.

(a) MOLECULER WT OF ASPARTANE = $C_{14}H_{18}N_2O_5$

$= (12)_{14} + (1)_{18} + (14)_2 + (16)_5$

THUS MASS OF 1 MOLE ASPARTANE = 294 g/MOLE.

(b) MOLE OF ASPARTAME IN 52 g = $\frac{52}{294}$ = 0.177 MOLE

(c) MASS IN GRAMS = MOLE × MOLAR MASS

= 10.122 × 294 = 2975.9 g.

(d) MOLES OF ASPARTAME = $\frac{\text{MASS IN g}}{\text{MOLAR MASS}} = \frac{2.43}{294} = 8.26\times10^{-3}$ MOLE

1 MOLE CONTAIN HYDROGEN = 18 MOLE ATOM

8.26×10^{-3} MOLE CONTAIN H = $18\times8.26\times10^{-3}$

= 0.149 MOLE ATOM

NO. OF HYDROGEN ATOMS = MOLE × AVOGADRO'S NO.

$0.149 \times 6.022\times10^{23}$

$= 8.96\times10^{22}$ ATOMS

Q.12 A sample of 0.600 mole of a metal M reacts completely with excess of fluorine to form 46.8 g MF_x.

a) How many moles of F are present in the sample of MF_x that forms.
b) Which element is represented by the symbol M?

SOLUTION:- EQUATION $M + F_2 \longrightarrow MF_2$

(a) SINCE 1 mole of M produces MF_2 = 1 MOLE

0.6 MOLE M will produce MF_2 = 0.60 mole

1 mole MF_2 contains F = 2 mole atom

0.600 mole MF_2 contain F = 2 × 0.600 = 1.20 MOLE ATOM.

(b) Since 0.60 MOLE MF_2 IS PRODUCED AND ITS MASS IS GIVEN 46.8 g. THUS MOLE = $\frac{\text{MASS IN GRAM}}{\text{FORMULA WT}}$

FORMULA WT = $\frac{\text{MASS IN g}}{\text{MOLE}}$

FORMULA WT = $\frac{46.8}{0.60}$ = 78 g/mole

FORMULA WT. OF MF_2 = $M + 2(F) = 78$

$M + 2(19) = 78$

$M = 78 - 38$

$M = 40$

This is atomic wt of "Ca". Thus M = "Ca"

Q No. 14 CALCULATE % OF NITROGEN IN FOLLOWING FERTILIZERS.

(a)

$\% N = \frac{\text{WT OF NITROGEN}}{\text{M WT}} \times 100$

FERTILIZER	MOLECULER WT	% N
NH_3	$14+3 = 17$	$= \frac{14}{17} \times 100 = 82.35$
NH_2CONH_2	$14+2+12+16+14+2 = 60$	$= \frac{28}{60} \times 100 = 46.67$
$(NH_4)_2SO_4$	$28+8+32+64 = 132$	$= \frac{28}{132} \times 100 = 21.21\%$
NH_4NO_3	$14+4+14+48 = 80$	$= \frac{28}{80} \times 100 = 35\%$

(b)

	MOLECULER WT.	% P	% N
$NH_4H_2PO_4$	$14+4+2+31+64 = 125$	$\frac{31}{125} \times 100 = 26.96$	12.17
$(NH_4)_2HPO_4$	$28+8+1+31+64 = 132$	$\frac{31}{132} \times 100 = 23.48$	21.21
$(NH_4)_3PO_4$	$42+12+31+64 = 149$	$\frac{31}{149} \times 100 = 20.8$	28.1

Q No. 15

Q.15 Glucose $C_6H_{12}O_6$ is the most important nutrient in the cell for generating chemical potential energy. Calculate the mass % of each element in glucose and determine the number of C, H and O atoms in 10.5 g of the sample.

MOLAR MASS OF GLUCOSE = $C_6H_{12}O_6$

$= 72 + 12 + 96 = 180$

% CARBON = $\frac{\text{WT OF CARBON}}{\text{M. WT OF GLUCOSE}} \times 100 = \frac{72}{180} \times 100 = 40\%$

% OXYGEN = $\frac{96}{180} \times 100$ = 53.4 %

% HYDROGEN = $\frac{12}{180} \times 100$ = 6.6 %

MOLES OF GLUCOSE = $\frac{\text{MASS IN GRAMS}}{\text{MOLAR MASS}} = \frac{10.5\,g}{180} = 0.058$ MOLE

MOLES OF CARBON = 6 X 0.058 = 0.35 MOLE ATOM

MOLES OF OXYGEN = 6 X 0.058 = 0.35 MOLE ATOM

MOLES OF HYDROGEN = 12 X 0.058 = 0.70 MOLE ATOM

ATOMS OF CARBON = MOLE X AVOGADRO'S NUMBER

$= 0.35 \times 6.022 \times 10^{23}$

$= 2.11 \times 10^{23}$

ATOMS OF OXYGEN = $0.35 \times 6.022 \times 10^{23} = 2.11 \times 10^{23}$ ATOM.

$= 0.70 \times 6.022 \times 10^{23} = 4.22 \times 10^{23}$ ATOM

Q. **Q.16** Ethylene glycol is used as automobile antifreeze. It has 38.7% carbon, 9.7 % hydrogen and 51.6% oxygen. Its molar mass is 62.1 grams mol⁻¹. Determine its empirical formula.

ELEMENT	%	No. OF GRAM ATOMS	ATOMIC RATIO	EMPIRICAL FORMULA
CARBON	38.7	$\frac{38.7}{12} = 3.22$	$\frac{3.22}{3.22} = 1$	CH_3O
HYDROGEN	9.7	$\frac{9.7}{1} = 9.7$	$\frac{9.7}{3.22} = 3$	
OXYGEN	51.6	$\frac{51.6}{16} = 3.22$	$\frac{3.22}{3.22} = 1$	

EMPIRICAL FORMULA WT = CH_3O

$= 12 + 3 + 16 = 31$

MOLECULER WT. = 62.1

$n = \frac{\text{M Wt}}{\text{Empirical formula Wt}} = \frac{62.1}{31} = 2$

MOLECULER FORMULA = (EMPIRICAL FORMULA) X n

$(CH_3O) \times 2 = C_2H_6O_2$.

Q NO. **Q.17** Serotenin (Molecular mass = 176 g mol⁻¹) is a compound that conducts nerve unpulses in brain and muscles. It contains 68.2 % C, 6.86 % H, 15.09 % N, and 9.08 % O. What is i·· molecular formula.

SOLUTION

ELEMENT	%	MOLE = %/At.Wt	ATOMIC RATIO			EMPIRICAL FORMULA
C	68.2	$\frac{68.2}{12} = 5.68$	$\frac{5.68}{0.56}$	= 10.14	= 10	$C_{10}H_{12}N_2O_1$
H	6.86	$\frac{6.86}{1} = 6.86$	$\frac{6.86}{0.56}$	= 12.25	= 12	
O	9.08	$\frac{9.08}{16} = 0.56$	$\frac{0.56}{0.56}$	= 1	= 1	
N	15.09	$\frac{15.09}{14} = 1.07$	$\frac{1.07}{0.56}$	= 1.9	= 2	

EMPIRICAL FORMULA WT = $C_{10}H_{12}N_2O_1$

$(12)_{10} + (1)_{12} + (14)_2 + (16)_1 = 176$

GIVEN MOLECULAR WT = 176

$n = \frac{m.Wt}{Emp.f.Wt} = \frac{176}{176} = 1$

Molecular formula = (Empirical formula) × n

$= (C_{10}H_{12}N_2O_1) \times 1 = C_{10}H_{12}N_2O_1$

Q.18 An unknown metal M reacts with S to form a compound with a formula M_2S_3. If 3.12 g of M reacts with exactly 2.88 g of sulphur, what are the names of metal M and the compound M_2S_3.

SOLUTION: EQUATION OF REACTION.

$2M + 3S \longrightarrow M_2S_3$

MOLES OF S = $\frac{\text{MASS IN GRAM}}{\text{ATOMIC WT}} = \frac{2.88}{32} = 0.09$ MOLE

3 MOLE S requires M = 2 MOLE

1 " " " " $= \frac{2}{3}$

0.09 " " m = $\frac{2}{3} \times 0.09 = 0.06$ MOLE

GIVEN MASS OF m = 3.12 g

THIS IS MASS OF 0.06 MOLE OF "m" which reacts with 0.09 MOLE OF "S".

MOLE = $\frac{\text{MASS in GRAMS}}{\text{ATOMIC WT}}$

ATOMIC WT = $\frac{\text{MASS IN GRAM}}{\text{MOLE}} = \frac{3.12}{0.06} = 52$

THIS IS ATOMIC WT OF "Cr" = Cr_2S_3

THUS M_2S_3 =

Q.19 The octane present in gasoline burns according to the following equation.

$2C_8H_{18}(l) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(l)$

a) How many moles of O_2 are needed to react fully with 4 moles of octane?

b) How many moles of CO_2 can be produced from one mole of octane?

c) How many moles of water are produced by the combustion of 6 moles of octane?

d) If this reaction is to be used to synthesize 8 moles of CO_2 how many grams of oxygen are needed? How many grams of octane will be used?

SOLUTION:

(a) 2 MOLE OCTANE REQUIRE OXYGEN = 25 MOLE

1 " " " " $= \frac{25}{2} \times 4 = 50$ MOLE

4 " " " "

GIVEN Al = 0.741
REQUIRED Al = 0.625
UNREACTED = 0.116 MOLE   MOLE X Al WT
MASS OF UNREACTED Al = 0.116 X 27 = 3.123 g.

✓ Q.21 A mixture of two liquids, hydrazine N_2H_4 and N_2O_4 are used as a fuel in rockets. They produce N_2 and water vapours. How many grams of N_2 gas will be formed by reacting 100 g of N_2H_4 and 200 g of N_2O_4.

$$2N_2H_4 + N_2O_4 \longrightarrow 3N_2 + 4H_2O$$

SOLUTION:

MOLES OF N_2H_4 = $\frac{100 g}{32}$ = 3.125 mole

MOLES OF N_2O_4 = $\frac{200 g}{92}$ = 2.174 mole

2 MOLE HYDRAZINE (N_2H_4) PRODUCES N_2 = 3 MOLE
1 " " " = $\frac{3}{2}$
3.125 MOLE " " " = $\frac{3}{2}$ x 3.125
N_2 = 4.688 MOLE

SIMILARLY
1 MOLE N_2O_4 PRODUCES N_2 = 3 MOLE
2.174 " " N_2 = 3 X 2.174 = 6.522 MOLE

SINCE N_2H_4 PRODUCES LEAST AMOUNT OF N_2.
THEREFORE N_2H_4 IS LIMITING REACTANT
AMOUNT OF N_2 = 4.688 MOLE
MASS OF N_2 = 4.688 X 28 = **131.26 g**

Q NO 22 $SiO_2 + 3C \longrightarrow SiC + 2CO$
When 100 kg sand is reacted with excess of carbon, 51.4 kg of SiC is produced. What is the percentage yield of SiC.

MASS OF SiO_2 = 100 kg = 100000 g, ACTUAL YIELD = 51.4 kg
MOLES OF SiO_2 = $\frac{100000}{60}$ = 1666.67 MOLE
1 MOLE SiO_2 PRODUCES SiC = 1 MOLE
1666.67 MOLE " SiC = 1666.67
MASS OF SiC = 1666.67 X 40
= 66666.67 g
THUS THEORATICAL YIELD = 66.66667 kg

% YIELD = $\frac{\text{ACTUAL YIELD}}{\text{THEORATICAL YIELD}}$ x 100
= $\frac{51.4}{66.67}$ x 100 = 77%

2 MOLE OCTANE PRODUCES CO_2 = 16 MOLE
1 MOLE OCTANE PRODUCES $CO_2 = \frac{16}{2}$ = 8.0 MOLE

(c) 2 MOLE OCTANE PRODUCES H_2O = 18 MOLE
1 " " " $H_2O = \frac{18}{2}$
6 " " " $= \frac{18}{2} \times 6$ = 54 MOLE.

(d) 16 MOLE CO_2 REQUIRES OXYGEN = 25 MOLE
1 " " " " $= \frac{25}{16}$
8 " " " " $= \frac{25}{16} \times 8$ = 12.5 MOLE

MASS OF OXYGEN = 12.5×32 = 400g.

Q No. 20 Calculate the number of grams of Al_2S_3 which can be prepared by the reaction of 20 g of Al and 30 g of sulphur. How much the non-limiting reactant is in excess?

SOLUTION. $2Al + 3S \longrightarrow Al_2S_3$

GIVEN Al = 20g
S = 30g

MOLE OF Al = $\frac{20}{27}$ = 0.741 MOLE
MOLE OF S = $\frac{30}{32}$ = 0.938 MOLE

2 MOLE Al ~~requ~~ produces Al_2S_3 = 1 MOLE
1 MOLE Al produces $Al_2S_3 = \frac{1}{2}$
0.741 MOLE Al produces $Al_2S_3 = \frac{1}{2} \times 0.741$ = **0.370 MOLE**

3 MOLE S produces Al_2S_3 = 1 MOLE
1 " " " " $= \frac{1}{3}$
0.938 MOLE will produce $Al_2S_3 = \frac{1}{3} \times 0.938$
Al_2S_3 = 0.313 MOLE

Since S produces least amount of products.
Therefore S is LIMITING REACTANT.
HENCE Al_2S_3 = 0.313 MOLE
MASS OF Al_2S_3 = 0.313 ×150
= 46.95 g.

NON LIMITING REACTANT

3 MOLE S REQUIRE Al = 2 MOLE
1 " " " $= \frac{2}{3}$
0.938 " " " $= \frac{2}{3} \times 0.938$ = 0.625 MOLE

2 MOLE OCTANE PRODUCES CO_2 = 16 MOLE
1 MOLE OCTANE PRODUCES CO_2 = $\frac{16}{2}$ = 8.0 MOLE

(c) 2 MOLE OCTANE PRODUCES H_2O = 18 MOLE
1 " " " H_2O = $\frac{18}{2}$
6 " " " = $\frac{18}{2} \times 6$ = 54 MOLE.

(d) 16 MOLE CO_2 REQUIRES OXYGEN = 25 MOLE
1 " " " " = $\frac{25}{16}$
8 " " " " = $\frac{25}{16} \times 8$ = 12.5 MOLE
MASS OF OXYGEN = 12.5 × 32 = 400 g.

Q No. 20 Calculate the number of grams of Al_2S_3 which can be prepared by the reaction of 20 g of Al and 30 g of sulphur. How much the non-limiting reactant is in excess?

SOLUTION. $2Al + 3S \longrightarrow Al_2S_3$

GIVEN Al = 20 g
S = 30 g
MOLE OF Al = $\frac{20}{27}$ = 0.741 MOLE
MOLE OF S = $\frac{30}{32}$ = 0.938 MOLE

2 MOLE Al ~~requ~~ produces Al_2S_3 = 1 MOLE
1 MOLE Al produces Al_2S_3 = $\frac{1}{2}$
0.741 MOLE Al produces Al_2S_3 = $\frac{1}{2} \times 0.741$ = 0.370 MOLE

3 MOLE S produces Al_2S_3 = 1 MOLE
1 " " " " = $\frac{1}{3}$
0.938 MOLE will produce Al_2S_3 = $\frac{1}{3} \times 0.938$
Al_2S_3 = 0.313 MOLE

Since S produces least amount of products.
Therefore S is LIMITING REACTANT.
HENCE Al_2S_3 = 0.313 MOLE
MASS OF Al_2S_3 = 0.313 × 150
= 46.95 g

<u>NON LIMITING REACTANT</u>

3 MOLE S REQUIRE Al = 2 MOLE
1 " " " = $\frac{2}{3}$
0.938 " " = $\frac{2}{3} \times 0.938$ = 0.625 MOLE