Chapter 1
Basic Concepts
Isotopes

Definition:

Atoms of the same element can possess different masses but same atomic numbers. Such atoms of an element are called isotopes.

Explanation

This phenomenon of isotopy was first discovered by Soddy. Isotopes are different kind of atoms of the same element having same atomic number, but different atomic masses due to same number of protons and electrons but different number of neutrons. The isotopes of an element possess same chemical properties and same position in the periodic table.

Examples

- Carbon has three isotopes written as $_6C^{12}$, $_6C^{13}$ and $_6C^{14}$ and expressed as C-12, C-13 and C-14. Each of these have 6-protrons and 6 electrons. These isotopes have 6, 7 and 8 neutrons, respectively.
- Hydrogen has three isotopes written as $_1H^1$, $_1H^2$ and $_1H^3$ called protium, deuterium and tritium.
- Oxygen has three, nickel has five, calcium has six, palladium has six, cadmium has nine and tin has eleven isotopes.

Relative Abundance of Isotopes

The percentage of one isotope of an element as compared to other isotopes of the same element occurring naturally is called relative abundance of isotopes.

The properties of a particular element mostly correspond to the most abundant isotope of that element. The relative abundance of the isotopes of elements can be determined by mass spectrometry.

Facts about Isotopes

- At present above 280 different isotopes occur in nature.
- They include 40 radioactive isotopes as well.
- About 300 unstable radioactive isotopes have been produced through artificial disintegration.
- The elements like arsenic, fluorine, iodine and gold etc have only a single isotope. They are called mono-isotopic elements.
- The elements of odd atomic number almost never possess more than two stable isotopes.
- The elements of even atomic number usually have larger number of isotopes and isotopes whose mass numbers are multiples of four are particularly abundant.
- For example, O-16, Mg-24, Si-28, Ca-40 and Fe-56 form nearly 50% of the earth's crust.
- Out of 280 isotopes that occur in nature, 154 have even mass number and even atomic number.

Determination of Relative Atomic Mass of Isotopes by Mass Spectrometer

Definition:

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

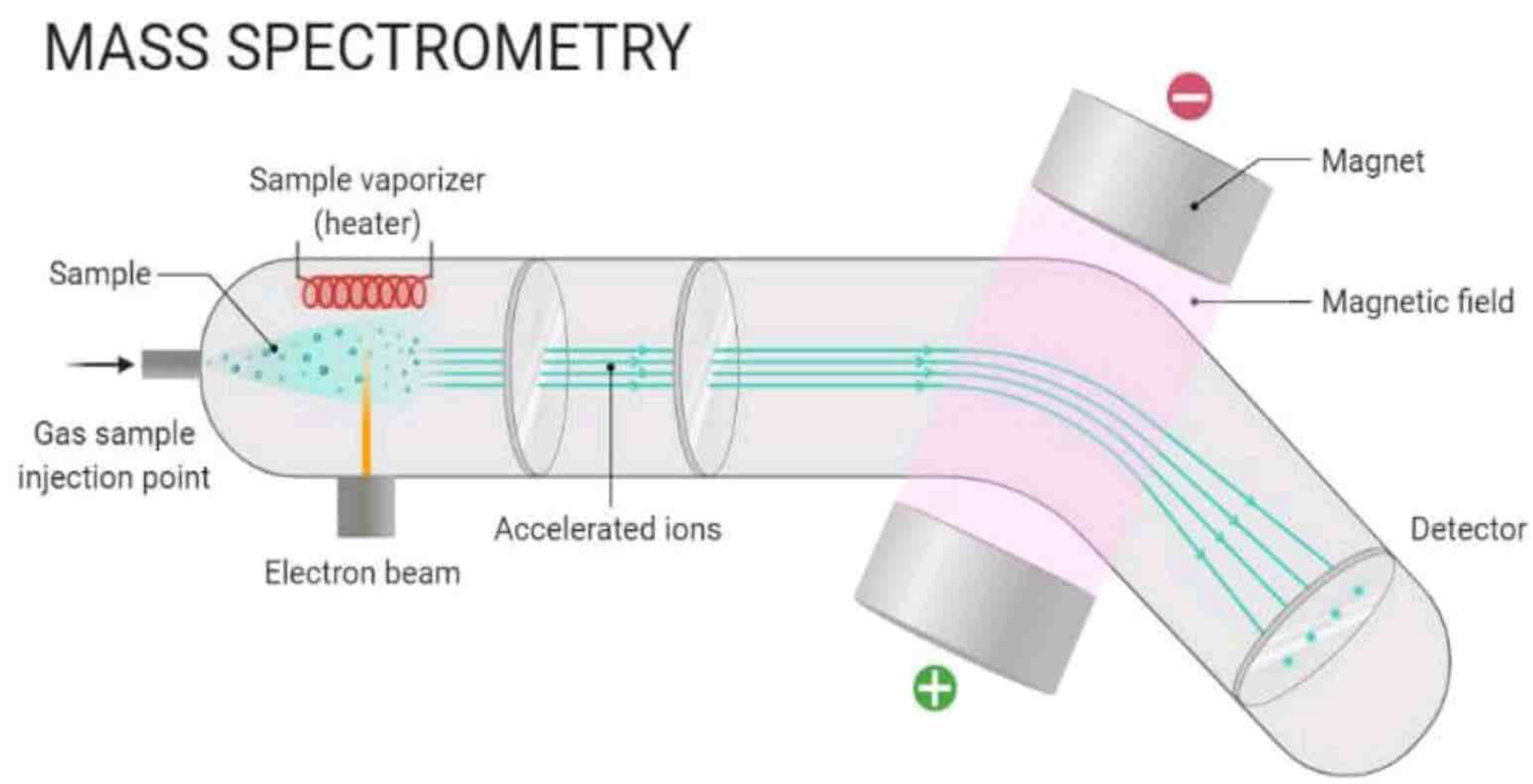
History:

- **Aston's mass spectrograph** was designed to identify the isotopes of an element on the basis of their atomic masses.
- **Dempster's mass spectrometer** was designed for the identification of elements which were available in solid state.

Construction and working:

The instrument has following components:

1. **Vaporization Chamber:**
    - The substance, the analysis of which is required, is converted into vapour state in the vaporization chamber.
    - The pressure of these vapours is kept very low, that is, 10^{-6} to 10^{-7} torr.



2. **Ionization Chamber:**
   - Vapors are allowed to enter the ionization chamber.
   - Fast moving electrons are thrown upon them.
   - The atoms in vapour state are ionized to cations.
   - The positively charged ions of isotopes of an element have different masses.
3. **Electric Field:**
   - The positive ions enter the electric field.
   - Electric Field is applied between perforated plates.
   - A potential difference (E) of 500-2000 volts is applied.
   - The ions are accelerated.

4. **Magnetic Field:**
   - The ions are then allowed to pass through a magnetic field of strength (H).
   - The magnetic field makes the ions to move in a circular path and then fall on the electrometer.
   - In this way ions are separated on the basis of their (m/e) values.

5. **Electrometer:**
   - It is also called ion collector and develops the electric current.
   - The strength of current thus gives the relative abundance of ions.
6. **Mathematical expression:**

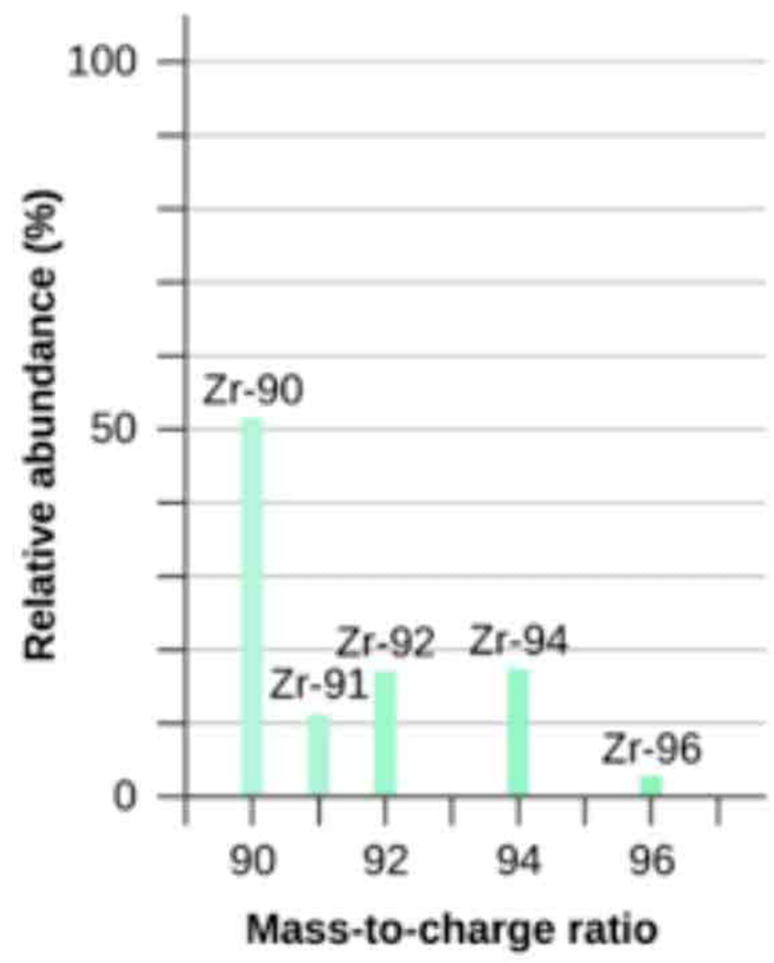
   The mathematical relationship for m/e ratio of isotopic ions is:

$$\frac{m}{e} = \frac{H^2r^2}{2E}$$

Where 'H' is the strength of magnetic field
'E' is the strength of electrical field
'r' is the radius of circular path adopted by isotopes in the magnetic field

7. **Comparison with C-12:**
   - The same experiment is performed with C-12 and the current strength is compared.
   - This comparison allows us to measure the exact mass number of the isotope of the element under study.
8. **Modern spectrograph:**
   - In modern spectrograph ions strike a detector.
   - Ionic current is amplified and fed into recorder.
   - Recorder plots a graph.
9. **Mass spectrum:**

The graph between mass to charge ratio on x-axis and relative abundance on Y-axis is known as mass spectrum.



Empirical Formula

Definition

It is the simplest formula that gives the small whole number ratio between the atoms of different elements present in a compound. In an empirical formula of a compound, AxBy, there are x atoms of an element A and y atoms of an element B.

Examples

The empirical formula of glucose ($C_6H_{12}O_6$) is CH_2O and that of benzene (C_6H_6) is CH.

Steps to Calculate Empirical Formula

Empirical formula of a compound can be calculated following the steps mentioned below:

1. Percentage composition

Determination of the percentage composition.

2. Number of Gram atoms

Finding the number of gram atoms of each element. Divide the mass of each element (% of an element) by its atomic mass.

3. Determination of Atomic Ratio

Determination of the atomic ratio of each element. Divide the number of moles of each element (gram atoms) by the smallest number of moles.

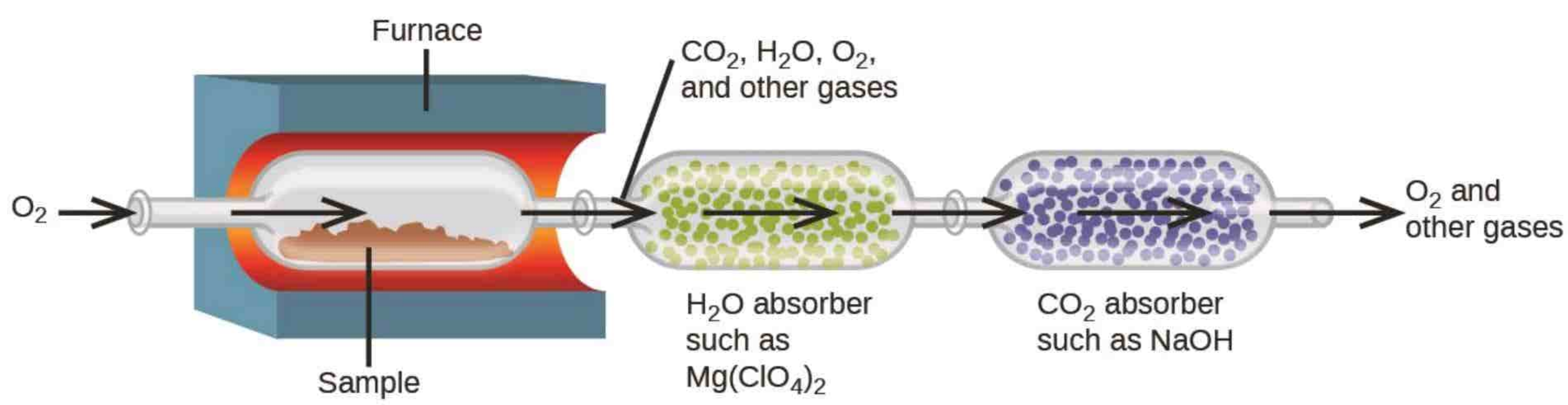
4. Multiply with a Suitable Digit

If the atomic ratio is simple whole number, it gives the empirical formula, otherwise multiply with a suitable digit to get the whole number atomic ratio.

Combustion Analysis

Definition: It is a technique used for finding empirical formula of the organic compounds which simply consist of carbon, hydrogen and oxygen.

Diagram:



Working:

1. **Furnace**

- A weighed sample of the organic compound is placed in the combustion tube.
- This combustion tube is fitted in a furnace.
- Oxygen is supplied to burn the compound.
- Carbon and hydrogen of organic compound are converted to CO_2 and H_2O, respectively.

2. **H_2O absorber**

- H_2O vapours are absorbed in magnesium per chlorate Mg $(ClO_4)_2$ solution.

3. **CO_2 absorber**

- CO_2 gas is absorbed in 50% KOH solution.
- The difference in the masses of these absorbers before absorbing gases and after absorbing gases gives the amount of CO_2 and H_2O produced.

Formulas to calculate the percentages of the elements:

Following formulas are used to calculate the percentages of carbon, hydrogen and oxygen respectively in the unknown organic compound.

% of carbon

$$\text{\% of carbon} = \frac{mass\ of\ CO2}{mass\ of\ organic\ compound} \times \frac{12.00}{44.00} \times 100$$

% of hydrogen

$$\text{\% of hydrogen} = \frac{mass\ of\ H2O}{mass\ of\ orgnaic\ compound} \times \frac{2.016}{18} \times 100$$

% of oxygen The percentage of oxygen is obtained by the method of difference.

$$\text{\% of oxygen} = 100 - (\text{\% of carbon} + \text{\% of hydrogen})$$

Molecular Formula

Definition

That formula of a substance which is based on the actual molecule is called molecular formula.

Examples

Molecular formula of benzene is C_6H_6 while that of glucose is $C_6H_{12}O_6$.

The empirical formulas of benzene and glucose are CH and CH_2O respectively, so for these compounds the molecular formulas are the simple multiple of empirical formulas.

$$\text{Molecular formula} = n\ (\text{Empirical formula})$$

Where 'n' is a simple integer.

$$n = \frac{\text{Molecular formula}}{\text{Empirical formula}}$$

Compounds with Same Molecular Formula and Empirical Formula

Those compounds whose empirical and molecular formulae are the same are numerous.

Examples

H_2O, CO_2, NH_3 and $C_{12}H_{22}O_{11}$ have same empirical and molecular formulas. Their simple multiple 'n' is unity.

Mole

Definition of mole

The atomic mass, molecular mass, formula mass or ionic mass of the substance expressed in gram is called molar mass of the substance.

Formula

$$\text{Number of gram atoms or moles of an element} = \frac{\text{Mass of an element in grams}}{\text{Molar mass of an element}}$$

Example
1 gram atom of hydrogen= 1.008 g
1 gram atom of carbon= 12.000 g
1 gram atom of uranium= 238.0 g

Formula
$$\text{Number of gram molecules or moles of molecule} = \frac{\text{Mass of molecular substance in grams}}{\text{Molar mass of the substance}}$$

Example
1 gram molecule of water= 18.0 g
1 gram molecule of H_2SO_4= 98.0 g
1 gram molecule of sucrose= 342.0 g

Formula
$$\text{Number of gram formulas or moles of a substance} = \frac{\text{Mass of the ionic substance in grams}}{\text{Formula mass of the ionic substance}}$$

Example
1 gram formula of NaCl= 58.50 g
1 gram formula of Na_2CO_3= 106 g
1 gram formula of $AgNO_3$= 170 g

Formula
$$\text{Number of gram ions or moles of a species} = \frac{\text{Mass of the ionic species in grams}}{\text{Formula mass of the ionic species}}$$

Example
1 gram ion of OH^-= 17 g
1 gram ion of SO_4^{2-}= 96 g
1 gram ion of CO_3^{2-}= 60 g

Avogadro's Number

Definition

Avogadro's number is the number of atoms, molecules and ions in one gram atom of an element, one gram molecule of a compound and one gram ion of a substance, respectively.

Examples
1.008 g of hydrogen = 1 mole of hydrogen = 6.02×10^{23} atoms of H
23 g of sodium = 1 mole of Na = 6.02×10^{23} atoms of Na
238 g of uranium = 1 mole of U = 6.02×10^{23} atoms of U
6.02×10^{23} is the number of atoms in one mole of the element.

18 g of H_2O =1 mole of water =6.02×10^{23} molecules of water
180 g of glucose = 1 mole of glucose = 6.02×10^{23} molecules of glucose
342 g of sucrose = 1 mole of sucrose = 6.02×10^{23} molecules of sucrose
One mole of different compounds has different masses but has the same number of molecules.

When we take into consideration the ions, then
96 g of SO_4^{2-} = 1 mole of SO_4^{2-} = 6.02×10^{23} ions of SO_4^{2-}
62 g of NO_3^- = 1 mole of NO_3^- = 6.02×10^{23} ions of NO_3^-

Formulas
$$\text{Number of atoms of an element} = \frac{\text{Mass of the element} \times N_A}{\text{Atomic mass}}$$

Number of molecules of a compound= $\frac{\text{Mass of the compound}}{\text{Molecular mass}} \times N_A$

Number of ions of an ionic species= $\frac{\text{Mass of the ion}}{\text{Ionic mass}} \times N_A$

Explanation through examples

1. In 18 g of water there are present 6.02 x 10^{23} molecules of H_2O, 2 x 6.02 x 10^{23} atoms of hydrogen and 6.02 x 10^{23} atoms of oxygen.
2. In 98g of H_2SO_4, it has twice the Avogadro's number of hydrogen atoms, four times the Avogadro's number of oxygen atoms and the Avogadro's number of sulphur atoms.
3. Dissolve 9.8 g of H_2SO_4 in sufficient quantity of H_2O to get it completely ionized. It has 0.1 moles of H_2SO_4. It will yield 0.2 mole or 0.2 x 6.02 x 10^{23} H^+ and 0.1 moles or 0.1 x 6.02 x 10^{23} SO_4^{2-} etc. Total positive charges will be 0.2 x 6.02 x 10^{23} and the total negative charges will be 0.2 x 6.02 x 10^{23}. The total mass of H^+ is (0.2 x 1.008) g and that of SO_4^{2-} is (0.1 x 96) g.

Stoichiometry

Definition

Stoichiometry is a branch of chemistry which tells us the quantitative relationship between reactants and products in a balanced chemical equation.

Assumptions

Following are the assumptions of stoichiometry:

1. All the reactants are completely converted into the products.
2. No side reaction occurs.
3. The law of conservation of mass and the law of definite proportions are obeyed.

Studied Relationships

The following type of relationships can be studied with the help of balanced chemical equation:

a. Mass-mass Relationship

If we are given the mass of one substance, we can calculate the mass of the other substances involved in the chemical reaction.

b. Mass-mole Relationship or Mole-mass Relationship

If we are given the mass of one substance, we can calculate the moles of other substance and vice versa.

c. Mass-volume Relationship

If we are given the mass of one substance, we can calculate the volume of the other substances and vice-versa.

d. Mole-mole Relationship

If we are given the moles of one substance, we can calculate the moles of the other substances.

Limiting Reactant

Definition

The limiting reactant is a reactant that controls the amount of the product formed in a chemical reaction due to its smaller amount.

Examples

From Daily Life

1. If we have 30 "kababs" and five breads "having 58 slices", then we can only prepare 29 "sandwiches". One "kabab" will be extra (excess reactant) and "slices" will be the limiting reactant.

2. Burning of wood in excess oxygen. In this case, wood is the limiting reactant and oxygen is the excess reactant.

From Chemistry

Consider the reaction between hydrogen and oxygen to form water.

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

We will get 2 moles (36g) of water because 2 moles (4g) of hydrogen react with 1 mole (32 g) of oxygen according to the balanced equation. Since less hydrogen is present as compared to oxygen, so hydrogen is a limiting reactant.

Identification of Limiting Reactant

To identify a limiting reactant, the following three steps are performed:

1. Calculate the number of moles from the given amount of reactant.

2. Find out the number of moles of product with the help of a balanced chemical equation.

3. Identify the reactant which produces the least amount of product as limiting reactant.

<u>Yield</u>

Definition

Actual Yield

The amount of the product obtained in a chemical reaction is called the actual yield.

Theoretical Yield

The amount of the product obtained through balanced chemical equation is called the theoretical yield.

Actual Yield less than Theoretical Yield

Actual yield is always less than theoretical yield due to the following reasons:

1. A practically inexperienced worker has many shortcomings and cannot get the expected yield.

2. The processes like filtration, separation by distillation, separation by a separating funnel, washing, drying and crystallization if not properly carried out, decrease the actual yield.

3. Some of the reactants might take part in a competing side reaction and reduce the amount of the desired product.

Efficiency of Reaction

The efficiency of a reaction is expressed by comparing the actual and theoretical yields in the form of percentage (%) yield.

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

Numericals

6. Silver has atomic number 47 and 16 known isotopes but only two occur naturally i-e ^{107}Ag and ^{109}Ag. Given the following mass spectrometric data. Calculate the average atomic mass of silver.

Data:

Isotopes	Mass (amu)	% Abundance
^{107}Ag	106.90509	51.84%
^{109}Ag	108.90476	48.16%

To Find:

Average atomic mass = ?

Solution:

Average atomic mass = (106.90509 x 51.84) + (108.90476 x 48.16) / 100

Average atomic mass = 107.87 amu

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

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

To Find:

Let % abundance of ^{10}B = x =?

Then % abundance of ^{11}B = 100 – x =?

Solution:

Average atomic mass = [(10.0129 .x) + (11.0093 . (100-x)] / 100 = 10.81 amu

10.0129 x + 1100.93 – 11.0093x = 1081

-0.9964x = 1081 – 1100.93

-0.9964x = -19.93

x = 19.93/ 0.9964

x = 20.002 %

Hence % abundance of ^{10}B = x = 20.002%

% abundance of ^{11}B = 100 – x = 100 – 20.002 = 79.998%

9. Justify the following statement:

(a) 23 grams of Sodium and 238 grams of Uranium have equal number of atoms in them.

1mole of Sodium = 23 grams

1 mole of Uranium = 238 grams

Since 1 mole of each element contains Avogadro's number of atoms i-e $N_A = 6.02 \times 10^{23}$ atoms. Hence 1 mole of each of Sodium and Uranium contain equal number of atoms i.e. 6.02×10^{23}

(b) Mg atom is twice heavier than carbon

One carbon atom contains 6 protons and 6 neutrons in its nucleus and its atomic mass is 12 amu

While one atom of Magnesium contains 12 protons and 12 neutrons in its nucleus and its atomic mass is 24 amu

Thus, Mass of Mg / Mass of Carbon = 24/ 12 = 2

Hence, One atom of Magnesium is twice heavier than that of one carbon atom

(c) 180g of glucose and 342g of sucrose have same number of molecules but different number of atoms

180 g of glucose = 1 mole of glucose

342 g of sucrose = 1 mole of sucrose

1 mole of each compound contains Avogadro's number of molecules. Hence both 1 mole of glucose (180grams) and 1mole of sucrose (342 grams) contain equal Avogadro's number of molecules $N_A = 6.02 \times 10^{23}$.

Since one molecule of glucose ($C_6H_{12}O_6$) contains 24 atoms. Whereas, one molecule of sucrose ($C_{12}H_{22}O_{11}$) contains 45 atoms.

(d) 4.9g of H_2SO_4 when completely ionized in water have equal number of positive and negative charges but the number of positive charge ions are twice the number of negatively charge ions.

$H_2SO_4 \rightleftharpoons 2H^+ + SO_4^{2-}$

This balance equation shows that 1 molecule of H_2SO_4 produces
Number of positively charged ions = H^+ = 2
Number of negatively charged ions = SO_4^{2-} = 1
Number of positive charges = 2 (Due to two H^+ ions)
Number of negative charges = 2 (Due to two negative charges on SO_4^{2-})
Hence, whatever be the amount of H_2SO_4. It will always produce equal number of positive and negative charges but number of positively charged ions will be twice the number of negatively charged ions.

(e) One mg of K_2CrO_4 has thrice the number of ions than the number of formula units when ionized in water

$K_2CrO_4 \rightleftharpoons 2K^+ + CrO_4^{2-}$
This equation shows that 1 formula unit of K_2CrO_4 produces two K^+ ions and one CrO_4^{2-} ion in solution. Thus a total of three ions are produced by ionization of 1 formula unit of K_2CrO_4.
Hence whatever be the amount of the K_2CrO_4 number of ions in its solution will always be thrice than the number of its formula units.

(f) Two grams of H_2, 16g of CH_4 and 44 g of CO_2 occupy separately the volume of 22.414dm^3, although the size and the masses of molecules of these gases are very different from each other.
2 grams of H_2 = 1 mole = N_A = 6.02 x 10^{23} molecules
16 grams of CH_4 = 1 mole = N_A = 6.02 x 10^{23} molecules
44 grams of CO_2 = 1 mole = N_A = 6.02 x 10^{23} molecules
In gases distance between two molecules is approximately 300 times than its molecular size. Thus, volume occupied by the gas molecules does not depend upon the size or the mass of molecules while it only depends upon the number of molecules. Hence, equal number of molecules of H_2, CH_4 and CO_2 at STP will occupy same volume i.e 22.414 dm^3. This is called as Avogadro's Law

10. Calculate each of the following:
(a) Mass in grams of 2.74 moles of $KMnO_4$
Data: Moles of $KMnO_4$=2.74 mol

To Find: Mass in grams of $KMnO_4$=?

KMnO4 (M=39 + 55 + 64 = 158 g mol-)

Formula: Number of moles $=\frac{\text{Mass in grams}}{\text{Molar mass}}$

Solution: Mass=Mole × Molar mass

Mass of $KMnO_4$ =No. of moles × Molar mass

= 2.74 × 158

= **432.9 g**

(b) Moles of O atoms in 9.00g of $Mg(NO_3)_2$
Data: Mass of $Mg(NO_3)_2$ =9g

To Find: Moles of moles of oxygen atoms =?

$Mg(NO_3)_2$(M=39 + 55 + 64 = 158 g mol^{-1})

Formula: Number of moles $=\frac{\text{mass in grams}}{\text{molar mass}}$

Solution:

Moles of $Mg(NO_3)_2 = \frac{9}{148}$ = 0.061 mol

1 moles of $Mg(NO_3)_2$ contains oxygen =6 mol

0.061 moles of $Mg(NO_3)_2$ contains oxygen =6 $\times$ 0.0061

= 0.366 moles of Oxygen

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

Data: Mass of $CuSO_4.5H_2O$ = 10.037g

To Find: No. of Oxygen atoms = ?

$CuSO_4$ $5H_2O$ (M=63.5 + 32 + 64 + 90=249.5 g mol^{-1})

Formula: Number of moles of $CuSO_4.5H_2O$ = $\frac{\text{mass}}{\text{molar mass}}$

Solution: Number of moles of $CuSO_4.5H_2O$ = $\frac{10.037}{249.5}$

= 0.04 moles of $CuSO_4.5H_2O$

1 Moles of $CuSO_4.5H_2O$contains = 9 moles of oxygen

0.04 moles of $CuSO_4$ $5H_2O$ contains = 9x 0.04

= 0.36 moles of oxygen

1 mole of oxygen atoms = N_A = 6.02 x 10^{23}atoms of oxygen

0.36 moles of oxygen contain = 6.02 x 10^{23} x 0.36 = 2.167 x 10^{23} atoms of oxygen

(d) Mass in kilogram of 2.6 $\times 10^{20}$ molecules of SO_2

Data: Molecules of SO_2 = 2.6×10^{20}

To Find: Mass in kilogram of SO_2 = ?

SO_2 (M=32 + 32 = 64 g mol^{-1})

Formula: Number of particles = $\frac{\text{mass}}{\text{molar mass}}$ x N_A

Solution: Mass = $\frac{\text{Number of particles x molar mass}}{N_A}$

Mass = $\frac{2.6\times10^{20} \text{ x } 64}{6.02\text{x}10^{23}}$

= $27.64\text{x}10^{-3}$ g

= $27.64\text{x}10^{-3}\text{x}10^{-3}$

= 27.64 x 10^{-6}kg

= $2.764\text{x}10^{-5}$ kg

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

Data: Mass of $C_2H_4Cl_2$ = 0.822g

To Find: Moles of Cl atoms = ?

$C_2H_4Cl_2$ (M=24 + 4+71 = 99 g mol^{-1})

Solution:

99g of $C_2H_4Cl_2$ contains moles of Cl = 2

1g of $C_2H_4Cl_2$ contains moles of Cl = $\frac{2}{99}$

0.822 g $C_2H_4Cl_2$ contains moles of Cl = $\frac{2}{99} \times 0.822$

= **0.017 moles of Cl**

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

Data: Moles of Silver Carbonate: = 5.136 mol

To Find: Mass in grams of Silver Carbonate = ?

Ag_2CO_3 (M=2(107.87) + 12 + 48 = 275.74 g mol^{-1}

Formula: Moles = $\frac{\text{mass}}{\text{molar mass}}$

Solution: Mass = moles x molar mass

Mass of Ag_2CO_3 = No. of moles $\times$moles of mass

= 5.136×275.74

= **1416.2 g**

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

Data: Molecules of $CrCO_2Cl_2$ = 2.78×10^{21}

To Find: Mass in grams of CrO_2Cl_2= ?

CrO_2Cl_2 (M=52+32+71= 155 g mol^{-1})

Formula: Number of particles = $\frac{\text{mass in grams}}{\text{molar mass}} \times N_A$

Solution: Mass = $\frac{\text{Number of particles x molar mass}}{N_A}$

Mass of CrO_2Cl_2 = $\frac{2.78 \times 10^{21} \times 155}{6.02 \times 10^{23}}$

= **0.715 g**

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

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

Data: Mass of $KClO_3$ = 100g

To Find:

Number of moles of $KClO_3$ = ?

Number of Formula units of $KClO_3$ = ?

$KClO_3$ (M=39+35.5+48= 122.5 g mol^{-1}

$$\text{Number of moles} = \frac{\text{mas in grams}}{\text{molar mass}}$$

$$\text{Number of particles} = \text{number of moles x } N_A$$

i. Calculation of Number of moles:

$$\text{Moles of } KClO_3 = \frac{100}{122.5} = 0.816 \text{ moles of } KClO_3$$

ii. Calculation of Number of formula units:

Number of formula units of $KClO_3$ = $0.816x\ 6.02 \times 10^{23}$

$= 4.9 \times 10^{2}$

Calculation of Number of K^+ ions, ClO_3^- ion, Cl atoms and O atoms

$KClO_3$: K^+ ions

1 : 1

$4.9 \times 10^{23} : 4.9 \times 10^{23}$

Number of K^+ ions = 4.9×10^{23}

$KClO_3$: ClO_3^- ions

1 : 1

$4.9 \times 10^{23} : 4.9 \times 10^{23}$

Number of ClO_3^- ions = 4.9×10^{23}

$KClO_3$: Chlorine atoms

1 : 1

$4.9 \times 10^{23} : 4.9 \times 10^{23}$

Number of Chlorineatoms = 4.9×10^{23}

$KClO_3$: Oxygen atoms

1 : 3

$4.9 \times 10^{23} : 3 \times 4.9 \times 10^{23}$

$: 1.473 \times 10^{24}$

Number of Oxygen atoms = 1.473×10^{24}

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

(a) What is the mass of one mole of aspartame?

Data: Molecular formula of aspartame = $C_{14}H_{18}N_2O_5$

To Find: Mass of one mole of aspartame (Molar mass) = ?

Solution: Molar mass of $C_{14}H_{18}N_2O_5$ = 14(12) +18(1) +2(14) +5(16)

= 168+18+28+80

= 294 $gmol^{-1}$

1 mole of aspartame = **294 g mol^{-1}**

(b) How many moles are present in 52g of Aspartame?

Data: Mass of aspartame = 52

To Find: Moles in 52 grams of aspartame = ?

Formula: Number of moles = $\frac{\text{mass in gram}}{\text{molar mass}}$

Solution: Mass = mole x molar mass

Moles of Aspartame = ?

Moles of aspartame = $\frac{52}{294}$ = **0.17 mol**

(c) What is the mass in grams of 10.122 moles of Aspartame.

Data: Moles of aspartame = 10.122 mol

To Find: Mass in grams of Aspartame = ?

Formula: Moles = $\frac{\text{mass in gram}}{\text{molar mass}}$

Solution: Mass = moles x molar mass

Mass of aspartame = 10.122×294

= **2975.86 g**

(d) How many hydrogen atoms are present in 2.43g of aspartame?

Data: Mass of aspartame = 2.43 g

To Find: Hydrogen atoms in Aspartame = ?

Formula: Number of particles = $\frac{\text{mass}}{\text{molar mass}} \times N_A$

Solution: Number of particles of aspartame = $\frac{2.43}{294} \times 6.02 \times 10^{23} = 4.97 \times 10^{21}$

$C_{14}H_{18}N_2O_5$	:	H atoms
1	:	18
4.97×10^{21}	:	$4.97 \times 10^{21} \times 18$
	:	8.95×10^{22}
Number of H atoms	=	**8.95×10^{22} atoms of hydrogen**

12. A sample of 0.600 mole of a metal M reacts completely with excess of fluorine to form 46.8g MF_2

(a) How many moles of F present in the sample of MF_2 that forms.

Data:

M	**+**	**F_2**	**$\rightarrow$**	**MF_2**
Mole=0.6		excess		Mass= 46.8 g

calculation of Moles of F in MF_2

M	:	MF_2
M	:	MF_2
1 mole	:	1mole
0.600 mol	:	0.600 mo

Moles of MF_2 = 0.6 mole

Moles of F in MF_2 = 0.6x2=**1.2 mole**

(b) Which element is represented by the symbol M?

$$\text{Moles of } MF_2 = \frac{\text{mass}}{\text{molar mass}}$$

$$\text{Molar mass} = \frac{\text{mass}}{\text{moles}}$$

$$= \frac{46.8}{0.600}$$

$$= 78 \text{g mole}^{-1}$$

Molar mass of MF_2 = 78

M + 2F = 78

M + 2(19) = 78

M = 78 – 38 = 40

Molar mass of M = 40 g mol^{-1}

Molar mass of '40' g mol^{-1} identifies Calcium (Ca)

14. (a) Calculate the percentage of nitrogen in the four important fertilizer i.e.

(i) NH_3 (ii) NH_2CONH_2 (iii) $(NH_4)_2SO_4$ (iv) NH_4NO_3

To Find: % age of Nitrogen = ?

Formula: %age of Nitrogen $= \frac{\text{Mass of Nitrogen in fertilizer}}{\text{Molar mass}} \times 100$

(i) NH_3 (Ammonia)

Molar mass of NH_3 = $14 + 3 = 17\ g\ mol^{-1}$

%age of Nitrogen = $\frac{\text{Mass of Nitrogen in fertilizer}}{\text{Molar mass}} \times 100$

= $\frac{14}{17} \times 100 =$ **82.35 %**

(ii) Urea (NH_2CONH_2)

Molar mass of NH_2CONH_2 = $60\ g\ mol^{-1}$

%age of N = $\frac{28}{60} \times 100 =$ **46.67 %**

(iii) $(NH_4)_2SO_4$ (Ammonium Sulphate)

Molar mass of $(NH_4)_2SO_4$ = $132\ g\ mol^{-1}$

%age of N = $\frac{28}{132} \times 100 =$ **21.2 %**

(iv) NH_4NO_3 (Ammonium Nitrate)

Molar mass of NH_4NO_3 = $80\ g\ mol^{-1}$

% age of N = $\frac{28}{80} \times 100 =$ **35 %**

(b) Calculate the percentage of nitrogen and phosphorus in each of the following:

(i) $NH_4H_2PO_4$ **(ii)** $(NH_4)_2HPO_4$ **(iii)** $(NH_4)_3PO_4$

(i) $NH_4H_2PO_4$ Ammonium hydrogen phosphate

Molar mass of $NH_4H_2PO_4$ = $115\ g\ mol^{-1}$

% age of N = $\frac{14}{115} \times 100 =$ **12.17%**

% age of P = $\frac{31}{115} \times 100 =$ **26.96%**

(ii) $(NH_4)_2HPO_4$ (Diammonium Hydrogen Phosphate)

Molar mass of $(NH_4)_2HPO_4$ = $132\ g\ mol^{-1}$

% age of N = $\frac{28}{132} \times 100 =$ **21.21%**

% age of P = $\frac{31}{132} \times 100 =$ **23.48%**

(iii) $(NH_4)_3PO_4$ (Ammonium Phosphate)

Molar mass of $(NH_4)_3PO_4$ = $3(14 + 4) + 31 + 4 (16)$

= $3(18) + 31 + 64$

= $54 + 31 + 64 = 149 \text{ g mol}^{-1}$

% age of N = $\frac{42}{149} \times 100$

= **28.19%**

% age of P = $\frac{\text{Mass of P}}{\text{Molar mass}} \times 100$

= $\frac{31}{149} \times 100$

= **20.8 %**

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.

Data: Mass of Glucose = 10.5 g

To Find:

(i) Mass % of each element in Glucose = ?

(ii) No. of atoms of C, H & O in 10.5g Glucose = ?

(i) Molar mass of $C_6H_{12}O_6$ = $72 + 12 + 96$

= 180 g mol^{-1}

% of C = $\frac{\text{Mass of Carbon}}{\text{Molar mass}} \times 100$

= $\frac{72}{180} \times 100 = \mathbf{40\ \%}$

% of H = $\frac{12}{180} \times 100 = \mathbf{6.66\ \%}$

% of O = $\frac{96}{180} \times 100 = \mathbf{53.34\ \%}$

(ii) Mass of glucose = 10.5g

Moles of glucose = $\frac{10.5}{180} = 0.058 \text{ mol}$

1 mole glucose contain C-atoms = $6 \times 6.02 \times 10^{23}$

0.058 mole glucose contains C-atoms = $6 \times 6.02 \times 10^{23} \times 0.058$

= $\mathbf{2.1 \times 10^{23}}$ **atoms**

1 mole glucose contains H-atoms = $12 \times 6.02 \times 10^{23} \times 0.058$

0.058 moles glucose contains H-atoms	=	$12\times6.02\times10^{23}\times0.058$
	=	**4.2×10^{23} atoms**
1 mole $C_6H_{12}O_6$ contains O-atoms	=	$6 \times 6.02 \times 10^{23}$
0.058 mole $C_6H_{12}O_6$ contains O-atoms	=	$6 \times 6.02 \times 10^{23} \times 0.058$
	=	**2.1×10^{23} atoms**

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. Determine its empirical formula.

Data:

% of Carbon	=	38.7%
% of Hydrogen	=	9.7%
% of Oxygen	=	51.6%
Molar mass of Ethylene glycol	=	62.1 g mol^{-1}

To Find:

Empirical formula of Ethylene glycol	=	?
Molecular formula of Ethylene glycol	=	?

	C	:	H	:	O
Moles	$\frac{38.7}{12}$	:	$\frac{9.7}{1}$	:	$\frac{51.6}{16}$
	3.225	:	9.7	:	3.225
Mole ratios	$\frac{3.225}{3.225}$	:	$\frac{9.7}{3.225}$	:	$\frac{3.225}{3.225}$
Atomic ratio	1	:	3	:	1

Empirical formula of Ethylene glycol is CH_3O

Empirical formula mass = CH_3O = 12 + (1x 3) + 16 = 31 $gmol^{-1}$

Molecular Mass = n (Empirical Mass)

n= Molecular mass/ Empirical mass

n = 62.1 $gmol^{-1}$ / 32 $gmol^{-1}$

n= 2

Molecular Formula = n (Empirical Formula)

= 2 (CH_3O)

Molecular Formula = $C_2H_6O_2$

17. Serotenin (Molar mass = 176g mol-1) is a compound that conducts nerve impulses in brain and muscles. H contains 68.2% C, 6.86% H, 15.09% N and 9.089%O. What is its molecular formula?

Data: %ages of elements in Serotenin:

% age of Carbon = 68.2%

% age of Hydrogen = 6.86%

% age of Oxygen = 9.08%

Molar mass of Serotenin = 176 g mol^{-1}

To Find: Molecular formula of Serotenin = ?

i. Finding empirical formula:

	C	:	H	:	N	:	O
Moles	$\frac{68.2}{12}$	:	$\frac{6.86}{1}$	:	$\frac{15.09}{14}$	:	$\frac{9.08}{16}$
	5.68	:	6.86	:	1.08	:	0.57
Molar ratios	$\frac{5.68}{0.57}$	:	$\frac{6.86}{0.57}$	:	$\frac{1.08}{0.57}$	:	$\frac{0.57}{0.57}$
Atomic ratio	10	:	12	:	2	:	1

Empirical formula is $C_{10}H_{12}N_2O$

ii. Finding Molecular formula:

Empirical formula mass = 120+12+28+16

= 176g

$$n = \frac{\text{Molar mass}}{\text{Empirical mass}} = \frac{176}{176} = 1$$

Molecular formula = n × empirical formula

= 1 ($C_{10}H_{12}N_2O$)

Molecular formula of Serotenin is $C_{16}H_{12}N_2O$

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

Data:

Mass of M = 3.12 g

Mass of S = 2.88

To Find:

Unknown metal M = ?

Compound M_2S_3 = ?

Solution: $2M + 3S \rightarrow M_2S_3$

3.12g $\frac{2.88}{32}$ = 0.09 mole

Identification = ?

S : M

According to equation: 3mol : 2 mol

3/3 : 2/3

1 : 2/3

1 x 0.09 : 0.09 x 2/3

0.09 : **0.06 mol**

Moles of M = 0.06

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar mass}}$$

Molar mass = Mass

Moles

= 3.12/0.06= 52 g mol^{-1} = molar mass of Chromium

The metal 'M' is Chromium and compound 'M_2S_3' is Cr_2S_3

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

$$2C_8H_{18(l)} + 25O_{2(g)} \rightarrow 16CO_{2(g)} + 18H_2O_{(l)}$$

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

$2C_8H_{18\,(l)} + 25O_{2\,(g)} \rightarrow 16CO_{2\,(g)} + 18H_2O_{\,(l)}$

(Octane)

To Find: Moles of O_2 = ?

2 moles of C_8H_{18} need O_2 = 25 mol

1 moles of C_8H_{18} need O_2 = $\frac{25}{2}$

4 moles of C_8H_{18} need O_2 = $\frac{25}{2} \times 4$

= 50 moles of O_2 are required

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

To Find: Moles of CO_2 = ?

2 moles C_8H_{18} produce CO_2 = 16 moles

1 mole C_8H_{18} produce CO_2 = $\frac{16}{2}$

= 8 moles of CO_2 can be produced

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

To Find: Moles of Water = ?

2 moles C_8H_{18} produce H_2O = 18 moles

1 mole C_8H_{18} produce H_2O = $\frac{18}{2}$

6 mole C_8H_{18} produce H_2O = $\frac{18}{2} \times 6$ = 54 moles of water are produced

(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?

To Find:

Mass of Oxygen needed = ?

Mass of Octane used = ?

16 moles CO_2 need O_2 = 25 moles

8 moles CO_2 need O_2 = $\frac{25}{16} \times 8$ = 12.5 moles

Mass of O_2 = 12.5×32 = **400 g**

16 moles CO_2 need C_8H_{18} = 2 moles

8 moles CO_2 need C_8H_{18} = $\frac{2}{16} \times 8$ = 1 mol

= **114 g** mass of Octane used

20. Calculate the number of grams of Al_2S_3 which can be prepared by the reaction of 20g of Al and 30g of Sulphur. How much the non-limiting reactant is in excess?

Data:

2Al	**+**	**3S**	**→**	**Al_2S_3**
mass=20 g		mass=30 g		mass=?
molar mass=27 $gmol^{-1}$		molar mass=32 $gmol^{-1}$		molar mass=150 $gmol^{-1}$
mole=20/27		mole=30/32		
= 0.74mol		= 0.9375mol		

i. **Identification of limiting reactant and calculation of mass of Al_2S_3**

Solution: To identify limiting reactant moles of both reactants shall be compared with product

	Al	**:**	**Al_2S_3**
According to equation:	2 mole	:	1 mole
	2/2	:	1/2
	1	:	1/2
	1x0.74	:	0.74 x 1/2
	3.125	:	0.37 moles of Al_2S_3
	S	**:**	**Al_2S_3**
According to equation:	3 mole	:	1 mole
	3/3	:	1/3
	1	:	1/3
	0.9375 x 1	:	0.9375 x 1/3
		:	0.3125moles of Al_2S_3

As number of moles of product obtained from Sulphur are less so Sulphur is limiting reactant.

Actually 0.3125 mole Al_2S_3shall be obtained.

Mass of Al_2S_3	=	No. of moles × Molar mass
	=	0.3125× 150
	=	**46.87 g**

ii. **Calculation of excess amount of non-limiting reactant:**

Solution: For this purpose, moles of both reactants shall be compared

	S	**:**	**Al**
According to equation:	3 mole	:	2 mole
	3/3	:	2/3
	1	:	2/3
	0.9375 x 1	:	0.9375 x 2/3
		:	0.625 moles of Al

For 0.9375 moles of S, we need only 0.623 moles of Al, however we have been provided 0.74 moles of Al.

Excess moles of Al	=	given mole- required mole
	=	0.74 - 0.625
	=	0.115
Excess mass of Al	=	mole x molar mass
	=	0.115 x 27
	=	**3.105 g**

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 vapors. How many grams of N_2 gas will be formed by reacting 100g of N_2H_4 and 200g of N_2O_4.

Data: $2N_2H_4 + N_2H_4 \rightarrow 3N_2 + 4H_2O$

mass of N_2H_4=100 g   mass of N_2O_4=200 g   mass=?

molar mass=32 $gmol^{-1}$molar mass= 92 $gmol^{-1}$ Molar mass=28 $gmol^{-1}$

mole=100/32=3.125   mole=200/92=2.17

Identification of limiting reactant and calculation of mass of N_2:

Solution: To identify limiting reactant moles of both reactants shall be compared with product

	N_2H_4	**:**	**N_2**
According to equation:	2mol	:	3 mol
	2/2	:	3/2
	1	:	3/2
	1x3.125	:	3.125 x 3/2
	3.125mol	:	4.6875mol
	N_2O_4	**:**	**N_2**
According to equation:	1mol	:	3 mol
	1x 2.17	:	3x 2.17
	2.17	:	6.51 mol

As number of moles of product obtained from N_2H_4 are less so N_2H_4 is limiting reactant. Actually 4.6875 moles of N_2 shall be obtained.

Mass of N_2 = No. of moles × Mol. Mass

= 4.6875× 28

= **131.25 g**

22. Silicon Chloride (SiC) is an important ceramic material. It is produced by allowing sand (SiO_2) to react with carbon at high temperature.

$$SiO_2 + 3C \rightarrow SiC + 2CO$$

When 100kg sand is reacted with an excess of carbon, 51.4 kg of SiCis produced.

What is the % yield?

Data: SiO_2 + 3C → SiC + 2CO

Mass=100Kg   mass = 51.4Kg

=100,000 g   = 514,00 g(Actual yield)

Molar mass= 60 $gmol^{-1}$ Molar mass = 40 $gmol^{-1}$

To Find: % age yield = ?

Formula: % age Yield : $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100$

i. Calculation of Theoretical yield of SiC

According to Equation

60g SiO_2 produces = 40g SiC

100,000 g SiO_2 produces = $\frac{40}{60} \times 100{,}000$

= **66666.6 g SiC**

ii. Calculation of % age yield of SiC

$$\%\text{ age yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

$$= \frac{51400}{66666.6} \times 100$$

= **77.1 %**

Important long questions according to past papers.

1. Define mass spectrometer. Explain the construction and working of mass spectrometer.
2. Describe combustion analysis. OR How can the %age of Carbon, Hydrogen and Oxygen in the given organic compound be estimated by combustion analysis?
3. What is the difference between actual yield and theoretical yield?
4. Define limiting reactant. How is it helpful to control chemical reaction?
5. What is stoichiometry? Give its assumptions and relationships studied. Mention two important laws which help to perform the stoichiometric calculations.
6. Example# 5, 10, 11, 12, 13
7. Exercise Numerical 16, 17, 20, 21