5 Mole Concept and Stoichimetry

SYLLABUS

Mole Concept and Stoichiometry

(i) Gay-Lussac's Law of Combining Volumes; Avogadro's Law.

Idea of mole – a number just as dozen, a gross; Avogadro's Law – statement and explanation; Gay-Lussac's Law of Combining Volumes – statement and explanation; "the mass of 22·4 litres of any gas at S.T.P. is equal to its molar mass" (Questions will not be set on formal proof but may be taught for clear understanding) – simple calculations based on the molar volume.

(ii) Refer to the atomicity of hydrogen, oxygen, nitrogen and chlorine (proof not required).

The explanation can be given using equations for the formation of HCl, NH_3 and NO.

(iii) Relative atomic masses (atomic weight) and relative molecular masses (molecular weights) : either H = 1 or $^{12}C = 12$ will be accepted; molecular mass = 2 × vapour density (formal proof not required). Deduction of simple (empirical) and molecular formula from the percentage composition of a compound; the molar volume of a gas at S.T.P.; simple calculations based on chemical equations; both reacting weight and volumes.

Idea of relative atomic mass and relative molecular mass — standard H atom, or 1/12th of carbon 12 atom.

Relating mole and atomic mass; arriving at gram atomic mass and then gram atom; atomic mass is a number dealing with one atom; gram atomic mass is the mass of one mole of atoms.

Relating mole and molecular mass arriving at gram molecular mass and gram molecule – molecular mass is a number dealing with a molecule, gram molecular mass is the mass of one mole of molecules.

Molecular mass = 2 × vapour density (questions will not be set on formal proof but may be taught for clear understanding); simple calculations based on the formula.

Deduction of simple (empirical) and molecular formula from the percentage composition of a compound.

5A. GAY-LUSSAC'S LAW AND AVOGADRO'S LAW

5.1 INTRODUCTION

You have learnt in Class IX that all gases behave similarly under similar conditions of temperature and pressure as expressed by gas laws.

Gas laws are certain rules applicable to a gas in respect of change in temperature (*T*) or pressure (*P*) or volume (*V*). The change in any one of these affects the other two.

Pressure-Volume Relationship or Boyle's Law

It states that *the volume of a given mass of dry gas is inversely proportional to its pressure at a constant temperature.*

$$P_1V_1 = P_2V_2 = k \text{ at constant temperature}$$

Temperature-Volume Relationship or Charles's Law :

It states that *volume of a given mass of a dry gas is directly proportional to its absolute (kelvin) temperature, if the pressure is kept constant.*

OR

The pressure remaining constant, the volume of a given mass of a dry gas increases or decreases by 1/273 of its volume for each 1°C increase or decrease in temperature respectively.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} = k \text{ at constant pressure}$$

Gas Equation

On combining both the laws we conclude that *the volume of a given mass of a dry gas varies inversely as the pressure and directly as the absolute temperature.*

$$V \propto \frac{1}{P} \times T$$

or $$V = \frac{T}{P} \times \text{constant}$$

or $$\frac{PV}{T} = k \text{ (constant)}$$

If the volume of a given mass of a gas changes from V_1 to V_2, pressure from P_1 to P_2 and temperature from T_1 to T_2 then,

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} = k \text{ (constant)}$$

Standard Temperature Pressure (STP)

Since the volume of a gas changes remarkably with change of temperature and pressure, it becomes necessary to choose standard values of temperature and pressure to which gas volumes can be referred.

The standard values chosen are : ***0°C or 273 K*** for temperature and *1 atmospheric pressure or **760 mm or 76 cm of Hg*** for pressure.

Temperature Scales

Absolute Scale or Kelvin Scale : *A temperature scale with absolute zero (zero kelvin) as the starting point is called the absolute scale or the kelvin scale.*

Absolute zero = 0 K = – 273°C

5.2 GAY-LUSSAC'S LAW OF COMBINING VOLUMES

Gay-Lussac (1805) observed that a simple relation exists between the volumes of hydrogen and oxygen, which react together to form water. Thus, one litre of oxygen requires two litres of hydrogen to form two litres of water vapour. So, he found that oxygen and hydrogen react in the ratio 1 : 2 (by volume). When he extended his study to the volumes of other reacting gases, he noted similar simple relationships. Consequently, he generalised these observations as the ***law of combining volumes of gases.***

Gay-Lussac's Law of Combining Volumes

When gases react, they do so in volumes which bear a simple ratio to one another, and to the volume of the gaseous product, provided that all the volumes are measured at the same temperature and pressure.

Note : Gay-Lussac's Law is valid only for gases. **The volumes of solids and liquids are considered to be zero.**

***The law may be illustrated by the following* examples** involving gases or vapours :

1. **Hydrogen chloride**

   One volume of hydrogen when mixed with one volume of chlorine, gives two volumes of hydrogen chloride gas at the same temperature and pressure.

   $$H_2 + Cl_2 \rightarrow 2HCl$$

   Thus, the ratio 1 : 1 : 2 is simple.

2. **Ammonia**

   One volume of nitrogen combined with three volumes of hydrogen gives two volumes of ammonia at the same temperature and pressure.

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

   Thus, the ratio 1 : 3 : 2 is simple.

3. **Carbon dioxide**

   Two volumes of carbon monoxide on combustion with one volume of oxygen gives two volumes of carbon dioxide at the same temperature and pressure.

   $$2CO + O_2 \rightarrow 2CO_2$$

   Thus, the ratio 2 : 1 : 2 is simple.

Numericals based on Gay-Lussac's Law

Example 1 : *What volume of oxygen would be required to burn completely 200 mL of acetylene (C_2H_2) and what would be the volume of carbon dioxide formed ?*

Solution :

$$C_2H_2 + O_2 \rightarrow CO_2 + H_2O$$

On balancing,

$$\underset{2\text{ vols}}{2C_2H_2} + \underset{5\text{ vols}}{5O_2} \rightarrow \underset{4\text{ vols}}{4CO_2} + \underset{\text{Nil}}{2H_2O}$$

According to the above equation — 2 volumes of acetylene requires 5 volumes of oxygen for complete combustion to produce 4 volumes of carbon dioxide.

∴ 200 mL of acetylene will require

$= \frac{5}{2} \times 200$ mL = 500 mL of oxygen, ***i.e., 500 mL of oxygen is required to produce 200 × 2 = 400 mL of carbon dioxide.***

Answer : 500 mL of oxygen is required to burn completely 200 mL of acetylene thereby producing 400 mL of carbon dioxide.

Example 2 : *If 6 litres of hydrogen and 5·6 litres of chlorine are mixed and exploded, what will be the composition by volume of the resulting gaseous mixture ?*

Solution :

$$\underset{1\text{ vol}}{H_2} + \underset{1\text{ vol}}{Cl_2} = \underset{2\text{ vols}}{2HCl}$$

Since 1 volume of chlorine reacts with 1 volume of hydrogen,

∴ 5·6 litres of chlorine will react with only 5·6 litres of hydrogen.

∴ (6 – 5·6) *i.e.,* 0·4 litres of hydrogen will remain unreacted.

Since the volume of HCl gas formed is twice that of chlorine used,

∴ volume of HCl formed will be

5·6 × 2 = 11·2 litres.

Answer : 11·2 litres of HCl and 0·4 litres of residual hydrogen.

> ***Note :*** The reactant which is completely used up in a reaction is known as **Limiting reagent** or **Limiting reactant**. In example 2 chlorine is the limiting reagent.

Example 3 : *What volume of propane is burnt for every 100 cm³ of oxygen in the reaction.*

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

(Gas volumes measured under the same conditions).

Solution :

From the above reaction, it is clear that for every 5 volumes of oxygen, 1 volume of propane is burnt.

∴ Volume of propane burnt for every 100 cm³ of oxygen $= \frac{1}{5} \times 100 = 20$ cm³

Answer : 20 cm³ of propane is burnt.

Example 4 : *80 cm³ of methane is mixed with 200 cm³ of pure oxygen at room temperature and pressure. The mixture is then ignited when it burns as illustrated by the equation :*

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

Calculate the composition of the resulting mixture if it is cooled to initial room temperature and pressure.

Solution :

$$\underset{1\text{ vol}}{CH_4\,(g)} + \underset{2\text{ vols}}{2O_2\,(g)} \rightarrow \underset{1\text{ vol}}{CO_2\,(g)} + \underset{\text{Nil}}{2H_2O\,(l)}$$

By Gay-Lussac's law,

(i) 1 vol of methane requires 2 vols of oxygen.

∴ 80 cm³ of methane requires

= 2 × 80 = 160 cm³ oxygen

(ii) 1 vol of methane produces 1 vol of carbon dioxide.

∴ 80 cm³ of methane produces 80 cm³ of carbon dioxide.

Hence, the composition of gaseous mixture after reaction is :

Methane	=	(80 – 80) = 0
Carbon dioxide	=	80 cm³
Oxygen	=	(200 – 160) = 40 cm³
Water	=	Negligible.

Example 5 : *40 cm³ of methane reacts with chlorine according to the following equation.*

Equation :

$$CH_4\,[g] + 2Cl_2\,[g] \rightarrow CH_2Cl_2\,[g] + 2HCl\,[g]$$

Calculate the volume of HCl gas formed and chlorine gas required.

Solution :

By Gay-Lussac's law :

1 Vol of CH_4 produces 2 Vols of HCl

∴ 40 cm³ of CH_4 will produce

2 × 40 = **80 cm³ of HCl.**

Since 1 Vol of CH_4 require 2 Vols of chlorine.

$\therefore$ 40 cm^3 of methane will require

$2 \times 40 =$ **80 cm^3 of chlorine.**

Answer : 80 cm^3 of HCl is formed.

80 cm^3 of chlorine is required.

Example 6 : *200 cm^3 of ethylene [C_2H_4] is burnt in just sufficient air (containing 20% oxygen) to form carbon dioxide gas and steam. If all measurements are made at constant pressure and 100°C, find the composition of the resulting mixture.*

Solution :

The reaction involved is given by :

$$\underset{1\ \text{vol.}}{C_2H_4(g)} + \underset{3\ \text{vols.}}{3O_2(g)} \xrightarrow{100^\circ C} \underset{2\ \text{vols.}}{2CO_2(g)} + \underset{2\ \text{vols.}}{2H_2O(g)}$$

By Gay-Lussac's Law;

(*i*) 1 vol. of ethylene requires oxygen = 3 vols

$\therefore$ 200 cm^3 of ethylene will require oxygen $= 3 \times 200 = 600\ cm^3$.

(*ii*) 1 vol. of ethylene produces carbon dioxide = 2 vols

$\therefore$ 200 cm^3 of ethylene will produce carbon dioxide $= 2 \times 200 = 400\ cm^3$.

(*iii*) 1 vol. of ethylene produces steam = 2 vols

$\therefore$ 200 cm^3 of ethylene will produce steam $= 2 \times 200 = 400\ cm^3$.

When oxygen is 20%, unreacted air is = 80%

When oxygen is 600 cm^3, then unreacted air is $= \frac{80 \times 600}{20} = 2400\ cm^3$.

Hence, composition of mixture after reaction :

(*i*) Carbon dioxide = 400 cm^3

(*ii*) Steam = 400 cm^3

(*iii*) Unreacted air = 2400 cm^3.

Example 7 : *20 mL of hydrogen, 10 mL of carbon monoxide and 20 mL of oxygen are exploded in an eudiometer. What will be the volume and composition of the mixture of gases, after cooling to room temperature ?*

Write each reaction separately.

Solution :

$2H_2$	+	O_2	=	$2H_2O$
2 vols		1 vol.		Nil
20 mL		10 mL		

$2CO$	+	O_2	=	$2CO_2$
2 vols		1 vol.		2 vols
10 mL		5 mL		10 mL

Total volume of oxygen used = (10 + 5) *i.e.* 15 mL.

Volume of oxygen left over = 20 – 15 = 5 mL

Volume of carbon dioxide formed = 10 mL

$\therefore$ Total volume of gases = 15 mL

Answer : Total volume of gaseous mixture after cooling to room temperature would be 15 mL, containing 5 mL oxygen and 10 mL carbon dioxide.

Example 8 : *A sample of coal gas contained 45% H_2, 30% CH_4, 20% CO, and 5% C_2H_2 by volume. 100 mL of this gaseous mixture was mixed with 160 mL of oxygen and exploded. Calculate the volume and the composition of the resulting mixture, when cooled to room temperature and pressure.*

Solution :

100 mL of gaseous mixture will contain 45 mL H_2, 30 mL CH_4, 20 mL CO and 5 mL C_2H_2.

Equations representing the involved combustion are :

(a) Hydrogen

$2H_2$	+	O_2	=	$2H_2O$
2 vols		1 vol		Nil
45 mL		22·5 mL		

(b) Methane

CH_4	+	$2O_2$	=	CO_2	+	$2H_2O$
1 vol		2 vol		1 vol.		Nil
30 mL		60 mL		30 mL		

(c) Carbon monoxide

$2CO$	+	O_2	=	$2CO_2$
2 vols		1 vol		2 vols
20 mL		10 mL		20 mL

(d) Acetylene

$2C_2H_2$	+	$5O_2$	=	$4CO_2$	+	$2H_2O$
2 vols		5 vols		4 vols		Nil
5 mL		12·5 mL		10 mL		

Total volume of oxygen used

= (22·5 + 60 + 10 + 12·5) = 105 mL

$\therefore$ Volume of oxygen left $= (160 - 105)$

$= 55$ mL

Total volume of carbon dioxide formed

$= 30 + 20 + 10 = 60$ mL

$\therefore$ Total volume of gases left

$= (55 + 60) = 115$ mL

Answer : 55 mL residual oxygen and 60 mL carbon dioxide and 115 mL total volume.

5.3 AVOGADRO'S LAW

Amedo Avogadro in 1811 put forward a hypothesis based on the relationship between the number of molecules in equal volumes of different gases under similar conditions.

Avogadro's law states that "equal volumes of all gases under similar conditions of temperature and pressure contain the same number of molecules."

This means that one ***litre*** of hydrogen contain the same number of molecules as are present in one ***litre*** of oxygen or in one litre of chlorine or of any other gas, provided the volumes of all gases are measured at the same temperature and pressure.

Avogadro further suggested that the smallest particle of a gaseous element is the molecule and not the atom though it may contain one, two or more atoms.

Example : A molecule of N_2 is made of two atoms of nitrogen, a molecule of NH_3 is made of one atom of nitrogen and three atoms of hydrogen.

He made the following distinction between atoms and molecules of a gaseous element.

*An **atom** is the smallest particle of an element that can take part in a chemical reaction; however, it may or may not exist independently.*

*A **molecule** is the smallest particle of an element or a compound that can exist by itself; it never breaks up except for taking part in a chemical reaction.*

Atomicity

*The number of atoms in a molecule of an element is called its **atomicity.***

(a) Monoatomic

Monoatomic molecule is composed of only one atom.

Examples : Inert gases like Helium, Neon, Argon, etc.

(b) Diatomic

Diatomic molecule is composed of two similar atoms.

Examples : H_2, O_2, Cl_2, N_2, etc.

(c) Triatomic

Triatomic molecule is composed of three similar atoms.

Example : Ozone gas (O_3).

(d) Tetratomic

Tetratomic molecule is composed of four similar atoms.

Example : Phosphorus (P_4).

(e) Octatomic

Octatomic molecule is composed of eight similar atoms.

Example : Sulphur (S_8).

Molecules made up of same type of atoms are homoatomic molecules, e.g. phosphorus (P_4), ozone (O_3) etc. while molecules made up of different types of atoms are hetero-atomic molecules, e.g. HCl, NH_3 etc.

Examples based on Avogadro's Law

Example 9 : *Under same conditions of temperature and pressure, you collect 2 litres of carbon dioxide, 3 litres of chlorine, 5 litres of hydrogen, 4 litres of nitrogen and 1 litre of sulphur dioxide. In which gas sample will there be*

(i) the greatest number of molecules ?

(ii) the least number of molecules ?

Justify your answer.

Solution :

Equal volumes of all gases under similar conditions of temperature and pressure contain the same number of molecules. So, under the same conditions of temperature and pressure, if volume of the gas is decreased, the number of molecules will also decrease.

Hence,

(i) 5 litres of hydrogen contain the *greatest number* of molecules.

(ii) 1 litre of sulphur dioxide contains the *least number* of molecules.

Example 10 : If 50 cc of a gas A contains y molecules, how many molecules of gas B will be present in 25 cc of B under same conditions ?

Solution :

50 cc contains y molecules.

$\therefore$ 25 cc will contain $\frac{y}{2}$ molecule.

According to Avogadro's law, 50 cc of A will contain equal molecules as 50 cc of B.

$\therefore$ 25 cc will contain half of what 50 cc contains.

EXERCISE-5A

1. State :

(a) Gay-Lussac's Law of combining volumes, **(2011)**

(b) Avogadro's law.

2. (a) Define atomicity of a gas. State the atomicity of Hydrogen, Phosphorus and Sulphur.

(b) Differentiate between N_2 and 2N.

3. Explain Why ?

(a) "The number of atoms in a certain volume of hydrogen is twice the number of atoms in the same volume of helium at the same temperature and pressure."

(b) "When stating the volume of a gas, the pressure and temperature should also be given."

(c) Inflating a baloon seems to violate Boyle's law.

NUMERICAL PROBLEMS

4. (a) Calculate the volume of oxygen at STP required for the complete combustion of 100 litres of carbon monoxide at the same temperature and pressure.

$$2CO + O_2 \rightarrow 2CO_2$$

(b) 200 cm^3 of hydrogen and 150 cm^3 of oxygen are mixed and ignited, as per following reaction,

$$2H_2 + O_2 \rightarrow 2H_2O$$

What volume of oxygen remains unreacted ?

5. 24 cc Marsh gas (CH_4) was mixed with 106 cc oxygen and then exploded. On cooling, the volume of the mixture became 82 cc, of which, 58 cc was unchanged oxygen. Which law does this experiment support? Explain with calculations.

6. What volume of oxygen would be required to burn completely 400 mL of acetylene [C_2H_2] ? Also, calculate the volume of carbon dioxide formed.

$$2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O\ (l)$$

7. 112 cm^3 of $H_2S(g)$ is mixed with 120 cm^3 of $Cl_2(g)$ at STP to produce HCl(g) and sulphur(s). Write a balanced equation for this reaction and calculate (i) the volume of gaseous product formed (ii) composition of the resulting mixture.

8. 2500 cc of oxygen was burnt with 600 cc of ethane [C_2H_6]. Calculate the volume of unused oxygen and the volume of carbon dioxide formed, after writing a balanced equation :

Ethane + Oxygen $\rightarrow$ Carbon dioxide + Water vapour

9. What volume of oxygen at STP is required to affect the combustion of 11 litres of ethylene [C_2H_4] at 273°C and 380 mm of Hg pressure?

$$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$$

10. Calculate the volume of HCl gas formed and chlorine gas required when 40 mL of methane reacts completely with chlorine at STP.

$$CH_4 + 2Cl_2 \rightarrow CH_2Cl_2 + 2HCl$$

11. What volume of propane is burnt for every 500 cm^3 of air used in the reaction under the same conditions ? (Assuming oxygen is 1/5th of air)

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

12. 450 cm^3 of nitrogen monoxide and 200 cm^3 of oxygen are mixed together and ignited. Calculate the composition of resulting mixture.

$$2NO + O_2 \rightarrow 2NO_2$$

13. If 6 litres of hydrogen and 4 litres of chlorine are mixed and exploded and if water is added to the gases formed, find the volume of the residual gas. **(2015)**

14. Ammonia may be oxidised to nitrogen monoxide in the presence of a catalyst according to the following equation.

$$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$$

If 27 litres of reactants are consumed, what volume of nitrogen monoxide is produced at the same temperature and pressure ?

15. A mixture of hydrogen and chlorine occupying 36 cm^3 was exploded. On shaking it with water, 4 cm^3 of hydrogen was left behind. Find the composition of the mixture.

16. What volume of air (containing 20% O_2 by volume) will be required to burn completely 10 cm^3 each of methane and acetylene.

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$
$$2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$$

17. LPG has 60% propane and 40% butane : 10 litres of this mixture is burnt calculate the volume of carbondioxide added to atmosphere.

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$
$$2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$$

18. 200 cm^3 of CO_2 is collected at STP when a mixture of acetylene and oxygen is ignited. Calculate the volume of acetylene and oxygen at STP in original mixture

$$2C_2H_2\,(g) + 5O_2\,(g) \rightarrow 4\,CO_2\,(g) + 2H_2O\,(g)$$

19. You have collected (a) 2 litres of CO_2 (b) 3 litres of chlorine (c) 5 litres of hydrogen (d) 4 litres of nitrogen and (e) 1 litre of SO_2, under similar conditions of temperature and pressure. Which gas sample will have :

(a) the greatest number of molecules, and

(b) the least number of molecules ?

Justify your answers.

20. The gases chlorine, nitrogen, ammonia and sulphur dioxide are collected under the same conditions of temperature and pressure. The following table gives the volumes of gases collected and the number of molecules (x) in 20 litres of nitrogen. You are to complete the table giving the number of molecules in the other gases in terms of x.

Gas	Volume (in litres)	Number of molecules
Chlorine	10	
Nitrogen	20	x
Ammonia	20	
Sulphur dioxide	5	

21. If 100 cm^3 of oxygen contains Y molecules, how many molecules of nitrogen will be present in 50 cm^3 of nitrogen under the same conditions of temperature and pressure?

ANSWERS

4. (a) 50 litres (b) 50 cm^3 **6.** 1000 mL of O_2, 800 mL of CO_2 **7.** (i) 224 cm^3 HCl (ii) 224 cm^3 HCl + 8 cm^3 Cl_2 **8.** 400cc, 1200cc. **9.** 8·25 litres **10.** HCl = 80 mL, Cl_2 = 80 mL **11.** 20 cc **12.** Unused NO = 50 cm^3, NO_2 formed = 400 cm^3, Total mixture = 450 cm^3. **13.** 2 litres of hydrogen as HCl form dissolves in water **14.** 12 litres **15.** 20 cm^3 of hydrogen, 16 cm^3 of chlorine **16.** 225 cm^3 **17.** 34 litres **18.** Acetylene 100 cm^3, oxygen 250 cm^3 **19.** (a) H_2 (b) SO_2 **20.** $x/2$, x, $x/4$ **21.** $\frac{y}{2}$ molecules

5B. RELATIVE ATOMIC MASS, RELATIVE MOLECULAR MASS AND MOLE CONCEPT

5.4 RELATIVE ATOMIC MASS (ATOMIC WEIGHT)

Atoms being extremely small, cannot be seen or weighed directly. But indirect methods of physics have enabled us to know the absolute mass of nearly all kinds of atoms. The mass of a hydrogen atom is found to be $1{\cdot}6735 \times 10^{-24}$ g while that of an oxygen atom is $26{\cdot}565 \times 10^{-24}$ g. As these masses are too small, it is not convenient to use kilograms or grams as unit. It has, therefore, been considered appropriate to use the mass of some standard atom as a unit and then relate masses of other atoms to it. The resulting masses of atoms are thus ***Relative Atomic Masses** (RAM)* or ***Atomic Weight***.

In the beginning, the mass of the hydrogen atom (hydrogen element being the lightest) was chosen as a unit and masses of other atoms were compared with it. In 1961, *carbon-12 was finally selected, because its adoption least affected the values of the atomic mass of the various elements on the old standard.*

The **relative atomic mass** or ***atomic weight*** *of an element is the number of times one atom of the element is heavier than $\frac{1}{12}$ times of the mass of an atom of carbon-12. Thus:*

$$\text{Relative atomic mass} = \frac{\textit{Mass of 1 atom of the element}}{\frac{1}{12}\textit{th the mass of one C-12 atom}}$$

Atomic mass is expressed in atomic mass units [a.m.u.]. **Atomic mass unit is defined as 1/12 the mass of carbon atom C-12.**

Fractional atomic masses

It is observed that most atomic masses are not whole numbers. The reason is that most natural elements are a mixture of constant composition containing two or more isotopes.* The relative atomic mass of any element is the weighted average of the relative atomic masses of its natural isotopes. For example, chlorine consists of a mixture of two isotopes of masses 35 and 37 in the ratio of 3:1.

The average relative atomic mass of chlorine

$$= \left(\frac{35 \times 3 + 37 \times 1}{4} = 35{\cdot}5\right)$$

Note : The mass of an ion will be same as that of parent atom. The mass of electron(s) gained or lost in forming the ion can be ignored in comparison to the total mass of the atom.

5.5 RELATIVE MOLECULAR MASS (MOLECULAR WEIGHT)

The relative molecular mass (or molecular weight) of an element or a compound is the number that represents how many times one molecule of the substance is heavier than 1/12 of the mass of an atom of carbon-12.

The Relative Molecular Mass (RMM) is obtained by adding together the relative atomic masses (atomic weights) of all the various atoms present in a molecule.

For example, relative molecular mass of sulphuric acid (H_2SO_4)

$$= 2 \times 1 + 32 + 4 \times 16 = 98$$

Thus, molecular mass of H_2SO_4 is 98 a.m.u. *i.e. one molecule of H_2SO_4 is 98 times as heavy as 1/12 the mass of carbon atom C-12.*

Example 1 : *Calculate the molecular masses (or molecular weights) of the following compounds:*

(a) Copper sulphate crystals, $CuSO_4 \cdot 5H_2O$

(b) Ammonium sulphate, $(NH_4)_2SO_4$

Given that the relative atomic masses of Cu = 63·5, S = 32, O = 16, N = 14 and H = 1.

Solution :

(a) The molecular mass of $CuSO_4 \cdot 5H_2O$

$= 63{\cdot}5 + 32 + (16 \times 4) + 5(1 \times 2 + 16)$

$= 159{\cdot}5 + 90 =$ **249·5 a.m.u.** **Ans.**

(b) The molecular mass of $(NH_4)_2SO_4$

$= 2(14 + 1 \times 4) + 32 + (16 \times 4)$

$= 2 \times 18 + 32 + 64$

$= 36 + 32 + 64 =$ **132 a.m.u.** **Ans.**

5.6 GRAM ATOMIC MASS

The atomic mass of an element expressed in grams is called **gram atomic mass.**

For example, the atomic mass of oxygen is 16 a.m.u., therefore its gram atomic mass is 16 g.

The quantity of the element which weighs equal to its gram atomic mass is called ***one gram atom*** *of that element.*

For example 16 g of oxygen is equal to *one gram atom* of oxygen. Similarly,

1 g atom of sodium = 23 gm of sodium

⇒ 1 g atom of He = 4 gm of helium

5.7 GRAM MOLECULAR MASS

The molecular mass of a substance expressed in grams is called gram molecular mass or molar mass. A sample of a substance with its mass equal to its gram molecular mass is called *one gram molecule of the substance.*

The molecular mass of water (H_2O) is 18·00 a.m.u., and its gram molecular mass is 18·00 g. Similarly, the relative molecular mass of ammonia (NH_3) is 17·00, therefore, its given molecular mass is 17·00 g.

5.8 MOLE CONCEPT

It is not practically possible to find the mass of a minute particle like an atom, a molecule, or an ion etc. Therefore, a collection of 6·022 ´ 10^{23} elementary particles called **Mole** is taken for all practical purposes.

Mole is simply like a dozen or a gross. A dozen is a collection of 12 objects, a gross is a collection

* *Atoms of the same element having same atomic number but different mass number are known as ISOTOPES.*

of 144 objects, similarly **a mole is a collection of 6·022 × 10^{23} particles** (atoms or molecules or ions).

A **mole** is the amount of pure substance containing the same number of chemical units as there are atoms in exactly 12 grams of carbon –12.

Note : When the particle is not specifically mentioned, then one mole is taken as 6·02 × 10^{23} molecule.

Avogadro's number *is defined as the number of atoms present in 12 g (gram atomic mass) of C-12 isotope, i.e., 6·022 × 10^{23} atoms.*

OR

Avogadro's number is the number of elementary units, i.e., atoms, ions or molecules present in one mole of a substance.

It is denoted by N_A.

5.8.1. Mole of atoms

"One mole of atoms contain 6·022 × 10^{23} atoms having mass equal to its gram atomic mass."

Thus, one mole of oxygen atoms contain 6·022 × 10^{23} atoms of oxygen and weighs 16 g.

5.8.2 Mole of Molecules

One mole contain 6·022 × 10^{23} molecules and is equivalent to the gram molecular mass of any given substance. Thus 1 mole of O_2 contain 6·022 × 10^{23} molecules and weighs 32 g.

1 mole of water contain 6·022 × 10^{23} molecules and weighs 18 g.

Mole and molar volume : It has been noticed that one mole of any gaseous molecule occupy 22·4 dm^3 (litre) or 22400 cm^3 (ml) at S.T.P. This volume is known as **molar volume.**

The molar volume of a gas can be defined as the volume occupied by one mole of a gas at S.T.P.

One G.M.M. of the substance = **1 mole** of that substance = **6·02 × 10^{23} molecules** of that substance = **22·4 *l* at S.T.P.** if substance is a gas

Unit : *1 litre = 1 dm^3 = 1000 cm^3 = 1000 ml*

Substance	Moles	Molar Mass (mass in grams)	Volume at S.T.P. in litres or dm^3.	No. of molecules	No. of atoms
Sulphur dioxide (SO_2)	1	64	22·4	6·022 × 10^{23}	S atoms – 6·022 × 10^{23} O atoms – 2 × 6·022 × 10^{23}
Ammonia (NH_3)	1	17	22·4	6·022 × 10^{23}	N atoms – 6·022 × 10^{23} H atoms – 3 × 6·022 × 10^{23}
Calcium Chloride ($CaCl_2$)	1	111	—	6·022 × 10^{23}	Ca^{2+} ions – 6·022 × 10^{23} Cl^- ions – 2 × 6·022 × 10^{23}

Example 2 : *The number of atoms in one mole (one g-atom of an element) is 6 × 10^{23}. Calculate,*

(a) the number of molecules in 14 g of nitrogen gas.

(b) the total number of atoms in 18 g of water.

(c) the number of chloride ions in 111 g of anhydrous calcium chloride.

Solution :

Molecular weight of any substance contain 6 × 10^{23} molecules.

(a) mass of 1 mole of nitrogen is 28 g.

28 g of nitrogen gas contains 6 × 10^{23} molecules.

$\therefore$ 14 g of nitrogen contains $\dfrac{6\times10^{23}\times14}{28}$

= 3 × 10^{23} molecules.

(b) Mass of 1 mole of water is 18 g.

$\therefore$ 18 g of water contains 6 × 10^{23} molecules

But H_2O contains 3 atoms.

(2 atoms of hydrogen + 1 atom of oxygen)

$\therefore$ 18 g of water contains

$3 \times 6 \times 10^{23}$ atoms $= 18 \times 10^{23}$ atoms.

(c) Mass of 1 mole of anhydrous calcium chloride ($CaCl_2$) = 40 + 71 = 111 g

But 1 molecule of $CaCl_2$ contains 2 chloride ions.

$\therefore$ 111 g of anhydrous $CaCl_2$ contains

$2 \times 6 \times 10^{23} = 12 \times 10^{23}$ chloride ions.

Answer : (a) 3×10^{23} molecules.

(b) 18×10^{23} atoms.

(c) 12×10^{23} chloride ions.

Example 3 : *How many moles of NaOH are present in 160 g of it ?*

Solution :

Mass of 1 mole of NaOH = 23 + 16 + 1 = 40 g.

$\therefore$ Moles in 160 g NaOH $= \frac{160}{40} = 4$ moles

Answer : There are 4 moles of NaOH in 160 g of it.

Example 4 : *What is the mass in grams of 2·5 moles of slaked lime $Ca(OH)_2$?*

Solution :

Mass of 1 mole of slaked lime $Ca(OH)_2$

$= 40 + 2 \times (16 + 1)$ g = 74 g.

$\therefore$ 2·5 moles of Ca $(OH)_2$ = 2·5 × 74 = 185 g.

Answer : 2·5 moles of Ca $(OH)_2$ weigh 185 g.

Example 5 : *Calculate the mass of*

(i) *one atom of oxygen and*

(ii) *one molecule of oxygen.*

Solution :

(i) Number of oxygen atoms in 16·0 g of atomic oxygen $= 6{\cdot}022 \times 10^{23}$ atoms.

$\therefore$ Mass of one atom of oxygen

$$= \frac{16\cdot0}{6\cdot022 \times 10^{23}} = 2{\cdot}657 \times 10^{-23} \text{ g}$$

Answer : Mass of one atom of oxygen is $2{\cdot}657 \times 10^{-23}$ g.

(ii) Gram molecular mass of oxygen = 32·0 g

Number of O_2 molecules in 32·0 g of O_2

$= 6{\cdot}022 \times 10^{23}$ molecules

$\therefore$ Mass of one molecule of oxygen

$$= \frac{32\cdot0}{6\cdot022 \times 10^{23}} = 5{\cdot}314 \times 10^{-23} \text{ g.}$$

Answer : Mass of one molecule of oxygen is $5{\cdot}314 \times 10^{-23}$ g.

Example 6 : *How many moles of sulphur atoms and oxygen atoms are present in one mole each of H_2SO_4, H_2SO_3 and SO_2 ?*

Solution :

One mole of H_2SO_4 contains 1 mole of sulphur atoms and 4 moles of oxygen atoms.

One mole of H_2SO_3 has 1 mole of sulphur atoms and 3 moles of oxygen atoms.

One mole of SO_2 has 1 mole of sulphur atoms and 2 moles of oxygen atoms.

Example 7 : *A silicon chip used in an integrated circuit of a microcomputer has a mass of 5.68 mg. How many silicon (Si) atoms are present in this chip?*

Solution :

Gram atomic mass of silicon is 28 g.

1g atom of silicon contain $6{\cdot}022 \times 10^{23}$ atoms

$\therefore$ 28g of silicon contain $6{\cdot}022 \times 10^{23}$ atoms

or 28000 mg of silicon contain $6{\cdot}022 \times 10^{23}$ atom

5.68 mg of silicon contain $\frac{6\cdot022 \times 10^{23}}{28000} \times 5\cdot68$

$= 1{\cdot}22 \times 10^{20}$ atoms

Example 8 : *Find the number of sodium and sulphate ions in 14·2 gram of sodium sulphate.*

Solution :

Molecular mass of Na_2SO_4

$= 2 \times 23 + 1 \times 32 + 4 \times 16 = 142$ a.m.u.

$\therefore$ 142 g Na_2SO_4 = 1 mole

14·2 g Na_2SO_4 = 0·1 mole

Total molecules $= 0{\cdot}1 \times 6{\cdot}02 \times 10^{23}$

$= 6{\cdot}02 \times 10^{22}$

1 molecule of Na_2SO_4 contain two Na^+ ions and one SO_4^{2-} ions.

$\therefore$ Na^+ ions present $= 2 \times 6{\cdot}02 \times 10^{22}$

$= \mathbf{1{\cdot}204 \times 10^{23}}$

SO_4^{2-} ions present $= 1 \times 6{\cdot}02 \times 10^{22}$

$= \mathbf{6{\cdot}02 \times 10^{22}}$

Example 9 : Calculate the total number of electrons present in 1·6 gram of methane.

Solution :

Molecular mass of CH_4

$= 12 + 1 \times 4 = 16$ a.m.u.

16 g = 1 mole

1·6 g = 0·1 mole

1 mole of CH_4 has $6{\cdot}02 \times 10^{23}$ molecule.

0·1 has $6{\cdot}02 \times 10^{22}$ molecule.

1 molecule of CH_4 contain $(6 + 1 \times 4) = 10$ electrons

$6{\cdot}02 \times 10^{22}$ molecules will contain

$= 10 \times 6{\cdot}02 \times 10^{22}$

$= \mathbf{6{\cdot}02 \times 10^{23}}$ **electrons.**

5.9 APPLICATIONS OF AVOGADRO'S LAW

1. It explains Gay-Lussac's law.
2. It determines atomicity of the gases.
3. It determines the molecular formula of a gas.
4. It determines the relation between molecular mass and vapour density.
5. It gives the relationship between gram molecular mass and gram molar volume.

(1) *Avogadro's Law explains Gay-Lussac's Law.*

According to Avogadro's Law, under the same conditions of temperature and pressure, equal volumes of different gases have the same number of molecules.

Since substances react in simple ratio by the number of molecules, volumes of the gaseous reactants and products will also bear a simple ratio to one another. This is what Gay-Lussac's Law says.

H_2	+	Cl_2	→	2HCl
1 vol		1 vol		2 vol (by Gay-Lussac's Law)
n molecules		n molecules		2n molecules (by Avogadro's Law)

(2) *Avogadro's Law predicts atomicity of gases.*

The application of Avogadro's Law to determine atomicity in elementary gases is illustrated by the following examples :

(i) *Atomicity of molecules of hydrogen and chlorine gases.*

In the reaction between hydrogen and chlorine, the volumes bear the following ratio:

H_2	+	Cl_2	→	2 HCl
1 vol		1 vol		2 vol (Gay-Lussac's Law)

Applying Avogadro's Law, x molecules of hydrogen combine with x molecules of chlorine to give $2x$ molecules of hydrogen chloride.

or 1 molecule of H_2 + 1 molecule of Cl_2 → 2 molecules of HCl

or $\frac{1}{2}$ molecule of H_2 + $\frac{1}{2}$ molecule of Cl_2 → 1 molecule of HCl

Since atoms are indivisible, half a molecule each of hydrogen and chlorine must contain at least one atom each and consequently one molecule each of hydrogen and chlorine must contain two atoms each.

(ii) *Atomicity of nitrogen and oxygen.*

Nitrogen	+	oxygen	→	Nitric oxide
N_2 1 vol	+	O_2 1 vol	→	2NO 2 vol
1 molecule of Nitrogen	+	1 molecule of Oxygen	→	2 molecules of Nitric Oxide
$\frac{1}{2}$ molecule	+	$\frac{1}{2}$ molecule	→	1 molecule

Half molecule of nitrogen is equal to one atom of nitrogen as an atom is indivisible.

So, one molecule of nitrogen and one molecule of oxygen contain two atoms each.

(3) *Determination of the molecular formula of a gaseous compound*

Molecular formula is a chemical formula which gives the actual number of atoms of the elements present in one molecule of a compound.

It is experimentally found that 2 volumes of ammonia decompose to give 1 volume of nitrogen and 3 volumes of hydrogen.

	Ammonia	→	Nitrogen	+	Hydrogen
	2 volumes		1 volume		3 volumes (by experiment)
	2 molecules		1 molecule		3 molecules (by Avogadro's Law)
or,	2 molecules		2 atoms		6 atoms
or,	1 molecule		1 atom		3 atoms

(since the molecules of nitrogen and hydrogen are diatomic).

Therefore, one molecule of ammonia contains 1 atom of nitrogen and 3 atoms of hydrogen. Hence, its molecular formula is NH_3.

(4) *It establishes a relationship between the relative vapour density of a gas and its relative molecular mass.*

The **Relative Vapour Density** of a gas (or a vapour) is :

The ratio between the masses of equal volumes of gas (or vapour) and hydrogen under the same conditions of temperature and pressure.

Relative V.D.

$$= \frac{\text{Mass of volume 'v' of the gas under similar conditions}}{\text{Mass of volume 'v' of hydrogen gas under similar conditions}}$$

According to Avogadro's Law, volumes at the same temperature and pressure may be substituted by molecules. Hence,

$$\text{Relative V.D.} = \frac{\text{Mass of 1 molecule of gas or vapour}}{\text{Mass of 2 atoms of hydrogen}}$$

(Molecule of hydrogen contains 2 atoms) By multiplying both sides by 2

$$2 \times \text{Rel. V.D.} = \frac{\text{Mass of 1 molecule of gas or vapour}}{\text{Mass of 1 atom of hydrogen}}$$

$2 \times$ Rel. V.D. = Rel. molecular mass of a gas or vapour.

> ***The relative molecular mass of a gas or vapour is twice its vapour density.***

For example, the relative vapour density of gaseous chlorine is 35·5, hence its relative molecular mass would be $2 \times 35{\cdot}5$ or 71·0.

(5) *It gives the value of molar volume of gases at S.T.P., and introduces the concept of Avogadro's number.*

> The **molar volume** of a gas is *the volume occupied by one gram-molecular mass or simply, by one mole of the gas at S.T.P. It is equal to 22·4 dm³.*

The value of molar volume is deduced from Avogadro's Law as follows :

It is found by direct weighing that 1 dm^3 (or 1 L) of hydrogen at S.T.P. weighs 0·089 g. Therefore, 1 dm^3 of any gas, whose vapour density at S.T.P. is D, would weigh $D \times 0{\cdot}089$ g at S.T.P.

OR

$D \times 0{\cdot}089$ g of any gas occupies 1 dm^3 (1 L).

$\therefore$ D g of any gas will occupy $\frac{1}{0{\cdot}089}$ or 11·2 dm^3 or 11·2 L

or 2D of any gas will occupy 22·4 dm^3 or 22·4 L.

But, $2 \times$ V.D. = Molecular mass.

Thus, gram molecular mass of any gas occupies 22·4 litres at S.T.P., which contains $6{\cdot}02 \times 10^{23}$ molecules of that gas and is equal to 1 mole of that gas.

Example 10 : *Calculate the atomicity of chlorine, if 35·5 g of it occupies 11,200 cm³ at S.T.P.*

Solution :

[1 mole of any substance = 1 g molecular mass of it and occupies 22400 cm^3 at S.T.P.].

1 mole of chlorine = 1 *g* mol. mass of chlorine and occupies 22,400 cm^3 at S.T.P.

Given :

1 gm atom of chlorine occupies 11,200 cm^3.

How many grams of chlorine will occupy 22,400 cm^3 at S.T.P.

$$= \frac{22400 \times 35 \cdot 5}{11200} = 71$$

$\therefore$ gram molecular mass = 71 g

$$\therefore \text{Atomicity} = \frac{\text{Mol. mass}}{\text{At. mass}} = \frac{71}{35{\cdot}5} = 2$$

$\therefore$ Atomicity of Chlorine = 2

Example 11 : *Calculate the volume occupied by 2·8 g of N_2 at S.T.P. ?*

Solution :

Gram molecular mass of $N_2 = 2 \times 14 = 28$ g

28 g of N_2 at S.T.P. occupies 22·4 litres

$\therefore$ 2·8 g of N_2 at S.T.P. will occupy

$$\frac{22{\cdot}4 \times 2{\cdot}8}{28} = 2{\cdot}24 \text{ litres}$$

Example 12 : *Calculate*

(a) *the gram molecular mass of N_2, if 360 cm³ at S.T.P. weighs 0·45 g;*

(b) *the number of gram molecule of water in 4·5 g of water.*

Solution :

The mass of 22·4 L of a gas at S.T.P. is equal to its gram molecular mass.

(a) 360 cm^3 of N_2 at S.T.P. weighs 0·45 g

$\therefore$ 22,400 cm^3 of N_2 will weigh

$\frac{0.45}{360} \times 22{,}400 = 28$ g.

Answer : Gram Molecular Mass of N_2 = 28 g

(b) No. of gram molecules $= \frac{\text{Mass in grams of water}}{\text{Gram molecular mass}}$

$= \frac{4.5\ g}{18\ g} = 0.25$ g molecules.

Answer : 0·25 g molecules are present in 4·5 g of water.

Example 13 : *What volume of oxygen is required to completely burn a mixture of 224 cm^3 of methane and 112 cm^3 of hydrogen at S.T.P. to produce steam and carbon dioxide ? If the steam formed is condensed to water, find the mass of water so formed.*

Solution :

For combustion of methane :

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

According to Gay-Lussac's Law, 1 volume of methane requires 2 volumes of oxygen to burn it completely, giving 1 volume of carbon dioxide and 2 volumes of steam.

Hence, 224 cm^3 of methane will require 448 cm^3 of oxygen for complete combustion and give 224 cm^3 of carbon dioxide and 448 cm^3 of steam.

For combustion of hydrogen :

$$2H_2 + O_2 \rightarrow 2H_2O$$

According to Gay-Lussac's Law, 2 volumes of hydrogen requires 1 volume of oxygen for complete combustion to give 2 volumes of steam.

$\therefore$ 112 cm^3 of hydrogen will require 56 cm^3 of oxygen for complete combustion and thereby form 112 cm^3 of steam.

$\therefore$ Total volume of oxygen needed for burning CH_4 and H_2 = 448 + 56 = 504 cm^3 and total volume of steam formed = 448 + 112 = 560 cm^3.

Molecular mass of H_2O = 2 + 16 = 18 amu
22400 cm^3 of steam at S.T.P. weighs 18 g.

$\therefore$ 560 cm^3 of steam at S.T.P. weighs

$= \frac{18 \times 560}{22400} = 0.45$ g

Answer : Volume of oxygen required is 504 cm^3 and the mass of water formed would be 0·45 g.

Example 14 : *A gas cylinder filled with hydrogen holds 50 g of the gas. The same cylinder holds 200 g of a gas X and 500 g of gas Y. Considering the same conditions of temperature and pressure in the cylinder, calculate the relative molecular masses of gases X and Y.*

Solution :

Vapour Density of a gas

$= \frac{\text{Mass of a certain volume of a gas at S.T.P.}}{\text{Mass of an equal volume of } H_2 \text{ at S.T.P.}}$

For Gas X

Vapour Density $= \frac{200}{50} = 4$

$\therefore$ Relative molecular mass of gas X = 2 × V.D.
= 2 × 4 = 8

For Gas Y

Vapour Density $= \frac{500}{50} = 10$

$\therefore$ Relative molecular mass of gas Y = 2×V.D.
= 2 × 10 = 20

Answer : Relative molecular mass of gas X = 8 and Gas Y = 20

Example 15 : 0·583 g of acetone when vaporized formed 225 cm^3 of vapours at S.T.P. Calculate the gram molecular mass of the gas.
[22·4 L. of any gas at stp = 1 g. mol. mass of gas]

Solution :

225 cm^3 of acetone at S.T.P. weighs = 0·583g

$\therefore$ 22400 cm^3 of acetone at stp will weigh

$= \frac{0.583 \times 22400}{225} = 58.04$ g. ***Ans.***

Example 16 : Calculate the V.D. and molecular mass of CO_2 if 200 mL of the gas at S.T.P. weighs 0·40 g.

[1 L of H_2 at S.T.P. weighs 0·09 g.]

Solution :

200 mL of CO_2 at S.T.P. weighs 0·40 g.

$\therefore$ 1 L of CO_2 at S.T.P. will weigh

$$= \frac{0{\cdot}40 \times 1000}{200} = 2 \text{ g.}$$

$$\text{Vapour density} = \frac{\text{Wt. of 1000 mL of } CO_2 \text{ at S.T.P.}}{\text{Wt. of 1000 mL of } H_2 \text{ at S.T.P.}}$$

$$= \frac{2}{0{\cdot}09} = \mathbf{22{\cdot}22} \qquad \textit{Ans.}$$

Molecular mass = 2 × V.D.

$$= 2 \times 22{\cdot}22 = \mathbf{44{\cdot}44 \text{ g.}} \qquad \textit{Ans.}$$

Example 17 : A gas at a pressure of 700 mm of Hg and a temperature of 57°C occupies 700 mL. If at S.T.P. the mass of the gas is 1.5 g, find the vapour density and the molecular mass of the gas (1 L of H_2 at S.T.P. weighs 0·09 g).

Solution :

P_1 = 700 mm of Hg $\qquad$ P_2 = 760 mm of Hg

V_1 = 700 mL $\qquad$ $V_2 = \chi$ mL

T_1 = 57 + 273 = 330 K $\qquad$ T_2 = 273 K

By Gas equation :

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} = \frac{700 \times 700}{330} = \frac{760 \times V_2}{273}$$

$$\therefore \quad V_2 = \frac{700 \times 700 \times 273}{760 \times 330}$$

$$= \mathbf{533{\cdot}37 \text{ mL}}$$

533·37 mL of the gas at S.T.P. weighs 1·5 g

$\therefore$ 1000 mL of the gas weighs $\dfrac{1{\cdot}5 \times 1000}{533{\cdot}37}$

$$= 2{\cdot}8123 \text{ g}$$

$$\text{Vapour density} = \frac{\text{Wt. of 1000 mL of the gas at S.T.P.}}{\text{Wt. of 1000 mL of } H_2 \text{ at S.T.P.}}$$

$$= \frac{2{\cdot}8123}{0{\cdot}09} = \mathbf{31{\cdot}25}$$

$\therefore$ Molecular mass = 2 × V.D. = 2 × 31·24

$$= \mathbf{62{\cdot}50}$$

Answer : Vapour density of the gas is 31·25

Molecular weight is 62·50.

Example 18 : Calculate the mass of a substance A which in gaseous form occupies 10 dm^3 at 27°C and 700 mm pressure. The molecular mass of A is 60g.

Solution :

P_1 = 700 mm of Hg $\qquad$ P_2 = 760 mm of Hg

V_1 = 10 dm^3 $\qquad$ $V_2 = \chi$ dm^3

T_1 = 27 + 273 = 300 K $\qquad$ T_2 = 273 K

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{700 \times 10}{300} = \frac{760 \times \chi}{273}$$

$$\therefore V_2 = \frac{700 \times 10 \times 273}{300 \times 760} = \mathbf{8{\cdot}38 \ dm^3}$$

Since 22·4 dm^3 of gas A at S.T.P. weighs 60 g

$\therefore$ 8·38 dm^3 of gas A at S.T.P. will weigh

$$\frac{60}{22{\cdot}4} \times 8{\cdot}38 = 22{\cdot}45 \text{ g}$$

$\therefore$ The mass of the substance A is 22·45 g.

Example 19 : Calculate the volumetric composition of water.

Solution :

$$2\,H_2O \rightarrow 2H_2 + O_2$$

2 vol. $\qquad$ 2 vol. $\qquad$ 1 vol.

Total = 3 vol.

Hence, the volumetric composition of water is

$$H_2 = \frac{2}{3} \times 100 = 66{\cdot}66\%$$

$$O_2 = \frac{1}{3} \times 100 = 33{\cdot}33\%$$

or $\quad H_2 : O_2$ in H_2O = 66·66 : 33·33 = 2 : 1 by volume.

Answer : 2 : 1 by volume.

Important Formulae for Revision

Mole and Gram Atomic Mass :	One mole of atoms	=	$6{\cdot}022 \times 10^{23}$ atoms
		=	Gram atomic mass of an element
		=	1 g atom of the element.
Mole and Gram Molecular Mass :	One mole of molecules	=	$6{\cdot}022 \times 10^{23}$ molecules
		=	Gram molecular mass
		=	1 g molecule of the compound.
Mole in terms of Volume :	One mole of a gas	=	22·4 litres at S.T.P.

Moles of an element $= \dfrac{\text{Mass of the element}}{\text{Atomic mass or G.A.W.}}$	**Moles of a compound** $= \dfrac{\text{Mass of the compound}}{\text{Molecular mass or G.M.W}}$
Mass of one atom $= \dfrac{\text{Atomic Mass or G.A.W.}}{6{\cdot}022 \times 10^{23}}$	**Mass of one molecule** $= \dfrac{\text{Molecular Mass or G.A.W.}}{6{\cdot}022 \times 10^{23}}$
Number of molecules = Moles $\times 6{\cdot}022 \times 10^{23}$	**Number of atoms** = Moles $\times 6{\cdot}022 \times 10^{23}$.

[a] **A mole of atom** = Avogadro's number of atoms = Gram atomic mass of that element.

Examples

- 1 mole of hydrogen atom [H] $= 6{\cdot}022 \times 10^{23}$ atoms of hydrogen = One gram of hydrogen
- 1 mole of sodium atom [Na] $= 6{\cdot}022 \times 10^{23}$ atoms of sodium = 23g of sodium

[b] **A mole of molecule** = Avogadro's number of molecules = Gram molecular mass of that substance.

Examples

- 1 mole of hydrogen molecule $= 6{\cdot}022 \times 10^{23}$ molecules of hydrogen = 2 grams of hydrogen [$H_2 = 1 \times 2$]
- 1 mole of water molecule $= 6{\cdot}022 \times 10^{23}$ molecules of water = 18 grams of water [$H_2O = 1 \times 2 + 16$]

[c] **A mole of an ionic compound** $= 6{\cdot}022 \times 10^{23}$ formula units of that compound = Mass equal to the formula mass of that substance.

Example

- 1 mole of $CaCl_2 = 6{\cdot}022 \times 10^{23}$ $CaCl_2$ units = 111 g of $CaCl_2$ [$CaCl_2 = 40 + 2 \times 35{\cdot}5$]

  i.e. $6{\cdot}022 \times 10^{23}$ Ca^{2+} ions

  i.e. 1 mole of Ca^{2+} ions

  and $2 \times 6{\cdot}022 \times 10^{23}$ Cl^- ions

  i.e. 2 moles of Cl^- ions

EXERCISE-5B

1. (a) The relative atomic mass of Cl atom is 35·5 a.m.u. Explain this statement.
   (b) What is the value of Avogadro's number ?
   (c) What is the value of molar volume of a gas at S.T.P.?
2. Define or explain the terms :
   (a) vapour density, (b) molar volume,
   (c) relative atomic mass, (d) relative molecular mass,
   (e) Avogadro's number, (f) Gram atom,
   (g) Mole.
3. (a) What are the main applications of Avogadro's Law ?
   (b) How does Avogadro's Law explain Gay-Lussac's Law of combining volumes?
4. Calculate the relative molecular masses of :
   (a) Ammonium chloroplatinate $(NH_4)_2 PtCl_6$,
   (b) Potassium chlorate, (c) $CuSO_4 \cdot 5H_2O$,
   (d) $(NH_4)_2SO_4$ (e) CH_3 COONa,
   (f) $CHCl_3$, (g) $(NH_4)_2$ Cr_2O_7.

5. Find the :
   (a) number of molecules in 73 g of HCl,
   (b) weight of 0·5 mole of O_2,
   (c) number of molecules in 1·8 g of H_2O,
   (d) number of moles in 10 g of $CaCO_3$,
   (e) weight of 0·2 mole of H_2 gas,
   (f) number of molecules in 3·2 g of SO_2.

6. Which of the following would weigh most :
   (a) 1 mole of H_2O, (b) 1 mole of CO_2,
   (c) 1 mole of NH_3, (d) 1 mole of CO ?

7. Which of the following contains maximum number of molecules :
   (a) 4 g of O_2, (b) 4 g of NH_3,
   (c) 4 g of CO_2, (d) 4 g of SO_2 ?

8. Calculate the number of :
   (a) particles in 0·1 mole of any substance,
   (b) hydrogen atoms in 0·1 mole of H_2SO_4,
   (c) molecules in one kg of calcium chloride.

9. How many grams of :
   (a) Al are present in 0·2 mole of it,
   (b) HCl are present in 0·1 mole of it ?
   (c) H_2O are present in 0·2 mole of it,
   (d) CO_2 is present in 0·1 mole of it ?

10. (a) The mass of 5·6 litres of a certain gas at STP is 12 g. What is the relative molecular mass or molar mass of the gas ?
    (b) Calculate the volume occupied at S.T.P. by 2 moles of SO_2.

11. Calculate the number of moles of :
    (a) CO_2 which contain 8·00 g of O_2,
    (b) methane in 0·80 g of methane.

12. Calculate the actual mass of :
    (a) an atom of oxygen, (b) an atom of hydrogen,
    (c) a molecule of NH_3, (d) the atom of silver,
    (e) the molecule of oxygen,
    (f) 0·25 gram atom of calcium.

13. Calculate the mass of 0·1 mole of each of the following:
    (a) $CaCO_3$, (b) $Na_2SO_4 \cdot 10H_2O$,
    (c) $CaCl_2$, (d) Mg.
    (Ca = 40, Na = 23, Mg = 24, S = 32, C = 12, Cl = 35·5, O = 16, H = 1)

14. Calculate the number of oxygen atoms in 0·10 mole of $Na_2CO_3 \cdot 10H_2O$.

15. What mass of Ca will contain the same number of atoms as are present in 3·2 g of S ? **(2015)**

16. Calculate the number of atoms in each of the following :
    (a) 52 moles of He, (b) 52 amu of He,
    (c) 52 g of He.

17. Calculate the number of atoms of each kind in 5·3 grams of sodium carbonate.

18. (a) Calculate the mass of nitrogen supplied to soil by 5 kg of urea $[CO(NH_2)_2]$.
    [O = 16; N = 14; C =12; H = 1]
    (b) Calculate the volume occupied by 320 g of sulphur dioxide at STP [S = 32; O = 16]

19. (a) What do you understand by the statement that 'vapour density of carbon dioxide is 22' ?
    (b) Atomic mass of chlorine is 35·5. What is its vapour density ?

20. What is the mass of 56 cm^3 of carbon monoxide at STP? [C = 12; O = 16]

21. Determine the no. of molecules in a drop of water which weighs ·09 g.

22. The molecular formula for elemental sulphur is S_8. In a sample of 5·12 g of sulphur :
    (a) How many moles of sulphur are present,
    (b) How many molecules and atoms are present ?

23. If phosphorus is considered to contain P_4 molecules, then calculate the number of moles in 100 g of phosphorus?

24. Calculate :
    (a) the gram molecular mass of chlorine if 308 cm^3 of it at STP weighs 0·979 g,
    (b) the volume of 4 g of H_2 at 4 atmosphere,
    (c) the mass of oxygen in 2·2 litres of CO_2 at STP.

25. A student puts his signature with graphite pencil. If the mass of carbon in the signature is 10^{-12}g, calculate the number of carbon atoms in the signature.

26. An unknown gas shows a density of 3 g per litre at 273°C and 1140 mm Hg pressure. What is the gram molecular mass of this gas ?

27. Cost of Sugar ($C_{12}H_{22}O_{11}$) is ₹ 40 per kg; calculate its cost per mole.

28. Which of the following weighs the least ?
    (a) 2 g atom of N (b) 3×10^{25} atoms of carbon
    (c) 1 mole of sulphur (d) 7 g silver

29. Four grams of caustic soda contains :
    (a) $6{\cdot}02 \times 10^{23}$ atoms of it,

(b) 4 g atom of sodium,

(c) $6{\cdot}02 \times 10^{22}$ molecules (d) 4 moles of NaOH.

30. The number of molecules in 4·25 g of ammonia is :

(a) $1{\cdot}0 \times 10^{23}$ (b) $1{\cdot}5 \times 10^{23}$,

(c) $2{\cdot}0 \times 10^{23}$ (d) $3{\cdot}5 \times 10^{23}$.

31. Correct the statements, if required.

(a) One mole of chlorine contains $6{\cdot}023 \times 10^{23}$ atoms of chlorine,

(b) Under similar conditions of temperature and pressure, two volumes of hydrogen combined with two volumes of oxygen will give two volumes of water vapour,

(c) Relative atomic mass of an element is the number of times one molecule of an element is heavier than 1/12 the mass of an atom of carbon [C^{12}],

(d) Under the same conditions of the temperature and pressure, equal volumes of all gases contain the same number of atoms.

ANSWERS

2. (a) 444 (b) 122.5 (c) 249.5 (d) 132 (e) 82, (f) 119·5 (g) 252 **5.** (a) $1{\cdot}2 \times 10^{24}$ (b) 16 g (c) $6{\cdot}02 \times 10^{22}$ (d) 0·1 mole (e) 0·4 g (f) 3×10^{22} **6.** (b) **7.** (b) **8.** (a) $6{\cdot}02 \times 10^{22}$ (b) $1{\cdot}2 \times 10^{23}$ (c) $5{\cdot}42 \times 10^{24}$ **9.** (a) 5·4 g (b) 3·65 g (c) 3·6 g (d) 4·4 g **10.** (a) 48 g (b) 44·8 lts **11.** (a) 0·25 moles (b) ·05 moles **12.** (a) $2{\cdot}657 \times 10^{-23}$ g (b) $1{\cdot}666 \times 10^{-24}$ g (c) $2{\cdot}823 \times 10^{-23}$ g, (d) $1{\cdot}7934 \times 10^{-22}$ g (e) $5{\cdot}314 \times 10^{-23}$ g (f) 10 g of calcium, **13.** (a) 10 g, (b) 32·2 g, (c) 11·1g, (d) 2·4 g, **14.** $7{\cdot}8 \times 10^{23}$ **15.** 4 g **16.** (a) $3{\cdot}131 \times 10^{25}$ (b) 13 atoms (c) $7{\cdot}828 \times 10^{24}$ **17.** Na = $6{\cdot}02 \times 10^{22}$, C = $3{\cdot}01 \times 10^{22}$, O = $9{\cdot}03 \times 10^{22}$ **18.** (a) 2·33 kg (b) 112 *l* **19.** (b) 35·5 **20.** 0·07 g **21.** $3{\cdot}01 \times 10^{21}$ molecules **22.** (a) 0·02 moles (b) $1{\cdot}2 \times 10^{22}$ molecules; $9{\cdot}635 \times 10^{22}$ atoms **23.** 0·805 moles **24.** (a) 71·2 g (b) 11·2 dm^3 (c) 3·14 g **25.** $5{\cdot}019 \times 10^{10}$ atoms **26.** 89·6 g **27.** ₹ 13.68 per mole **28.** (d) **29.** (c) **30.** (b)

5C. PERCENTAGE COMPOSITION, EMPIRICAL AND MOLECULAR FORMULAE

5.10 Percentage composition

Percentage composition of a compound, is the percentage by weight of each element present in it.

Percentage of an element in a compound

$$= \frac{\text{Total wt. of the element in one molecule}}{\text{Gram molecular weight of the compound}} \times 100$$

Example 1 : *Calculate the percentage composition of various elements in :*

Sodium carbonate, Na_2CO_3

Given that the relative atomic masses of O = 16, Na = 23 and C = 12.

Solution :

Relative molecular mass of Na_2CO_3

$= 23 \times 2 + 12 + 16 \times 3$

$= 46 + 12 + 48 = 106$

Since 106 g of Na_2CO_3 contains 46 g of sodium,

$\therefore$ 100 g of Na_2CO_3 contains $\frac{46 \times 100}{106}$ of sodium

$= \frac{4600}{106} = 43{\cdot}4$ g of Sodium

Similarly, 106 g Na_2CO_3 contains 12 g of carbon.

$\therefore$ 100 ” ” $\frac{12 \times 100}{106}$

$= \frac{1200}{106} = 11{\cdot}3$ g of C

Again, 106 g of Na_2CO_3 contains 48 g of oxygen.

$\therefore$ 100 ” ” $\frac{48 \times 100}{106}$

$= \frac{4800}{106} = 45{\cdot}3$ g of O

Answer : In Na_2CO_3 : Na = 43·4%,
C = 11·3% and O = 45·3%

Example 2 : *Find the percentage mass of water in washing soda crystals $Na_2CO_3 \cdot 10H_2O$.*

Solution :

Relative molecular mass of

$Na_2CO_3 \cdot 10H_2O$

$= 23 \times 2 + 12 + 16 \times 3 + 10\,(2 + 16)$

$= 106 + 180 = 286$

286 g of washing soda contains 180 g of water of crystallisation.

$\therefore$ 100 ” ” $\frac{180 \times 100}{286}$

$= \frac{18000}{286} = 62{\cdot}9$ g of H_2O

Answer : The % of H_2O in $Na_2CO_3 \cdot 10H_2O = 62{\cdot}9$

Example 3 : *A given sample has 30% ethyl alcohol by mass, the rest being water. Determine the percentage of C, H and O in the given sample?*

Solution :

Molecular mass of ethyl alcohol

$C_2H_5OH = 2 \times 12 + 5 \times 1 + 16 + 1 = 46$

Carbon present in 46 parts by mass of C_2H_5OH $= 2 \times 12 = 24$ parts by mass.

Carbon present in 30 parts by mass of C_2H_5OH

$= \frac{24}{46} \times 30 = \frac{720}{46} = 15{\cdot}65$

Hydrogen present $= \frac{6}{46} \times 30 = 3{\cdot}91$

Oxygen present $= \frac{16}{46} \times 30 = 10{\cdot}43$

In water (70%)

Hydrogen present $= \frac{2}{18} \times 70 = 7{\cdot}78$

Oxygen present $= \frac{16}{18} \times 70 = 62{\cdot}22$

Percentage of C = 15·65%

Percentage of H = 3·91 + 7·78 = 11·69%

Percentage of O = 10·43 + 62·22 = 72·65%

5.11 EMPIRICAL FORMULA OF A COMPOUND

The empirical formula of a compound is the simplest formula, which gives the simplest ratio in whole numbers of atoms of different elements present in one molecule of the compound.

For example, the empirical formula of hydrogen peroxide is HO. It indicates the simplest ratio (1 : 1) between the hydrogen and oxygen atoms in its molecule whereas its actual formula is H_2O_2.

Similarly, the empirical formula of glucose ($C_6H_{12}O_6$) is CH_2O. It indicates that the ratio of C, H and O atoms in a molecule of glucose is 1 : 2 : 1.

***The empirical formula mass** is the sum of atomic masses of various elements present in the empirical formula.*

Thus, for hydrogen peroxide (H_2O_2), the empirical formula is HO and its empirical formula mass is 1 + 16 = 17.

5.12 DETERMINATION OF EMPIRICAL FORMULA

The empirical formula can be calculated from the percentage composition of the compound as follows :

(1) The percentage weight of oxygen, carbon or hydrogen (necessary elements in every organic compound) is sometimes not given. Hence, it is first determined by subtracting the sum of other elemental percentages from 100.

(2) The percentage weight of each element is divided by its atomic weight. This gives the ratio of the number of atoms in a molecule of the compound.

(3) If the molecule contains water of crystallisation, its percentage weight is divided by the molecular weight of water, *i.e.*, 18. This gives the ratio of the number of water molecules in one molecule of the compound.

(4) *To get a whole number ratio of atoms of different elements and water molecules*, the numbers obtained in steps 2 and 3 are divided by the smallest number.

(5) The empirical formula is now derived by writing the symbols of various elements side by side with the number of atoms of each one as the subscript to the lower right of its symbol.

***Example 4** : The molecular formula of an organic acid is CH_3COOH. What is its empirical formula?*

Solution :

Molecular formula = CH_3COOH, *i.e.*, $C_2H_4O_2$

∴ Ratio of C, H and O is 2 : 4 : 2

Simple ratio is, *i.e.* 1 : 2 : 1

∴ Empirical Formula = CH_2O

***Example 5** : Find the empirical formulae of the compounds with the following percentage compositions :*

(a) Zn = 47·8; Cl = 52·2

(b) Mg = 9·76%, S = 13·01%, O = 26·01%, H_2O = 51·22%

Solution :

(a) *Zn = 47·8; Cl = 52·2*

Element	*Percentage composition*	*Atomic mass*	*Atomic ratio*	*Simplest ratio*
Zn	47·8	65	$\frac{47.8}{65} = 0.73$	$\frac{0.73}{0.73} = 1$
Cl	52·2	35·5	$\frac{52.2}{35.5} = 1.46$	$\frac{1.46}{0.73} = 2$

Thus, the ratio of Zn : Cl atoms = 1 : 2

Answer : The empirical formula of the compound is $ZnCl_2$.

(b) *Mg = 9·76%, S = 13·01%, O = 26·01%, H_2O = 51·22%*

Solution :

Element compound	*Percentage composition*	*Atomic or molecular mass*	*Ratio of atoms/ molecules*	*Simplest ratio*
Mg	9·76	24	$\frac{9.7}{24}$ = 0·406	$\frac{0.406}{0.406} = 1$
S	13·01	32	$\frac{13.01}{32}$ = 0·406	$\frac{0.406}{0.406} = 1$
O	26·01	16	$\frac{26.01}{16}$ = 1·625	$\frac{1.625}{0.406} = 4$
H_2O	51·22	18	$\frac{51.22}{18}$ = 2·84	$\frac{2.84}{0.406} = 7$

Answer : The empirical formula of the compound is $MgSO_4 \cdot 7H_2O$.

***Example 6** : An organic compound has oxygen 26·24% and hydrogen 4·92%. Find its empirical formula.*

Solution :

Every organic compound contains carbon. Find the percentage of carbon :

Percentage of carbon = 100 – (26·24 + 4·92)

= 68·84

Element	Percentage composition	Atomic mass	Atomic ratio	Simplest ratio
C	68·84	12	$\frac{68.84}{12} = 5.74$	$\frac{5.74}{1.64} = 3.5$
H	4·92	1	$\frac{4.92}{1} = 4.92$	$\frac{4.92}{1.64} = 3$
O	26·24	16	$\frac{26.24}{16} = 1.64$	$\frac{1.64}{1.64} = 1$

Multiply 3·5, 3 and 1 by 2 in order to change them in whole number. Thus, the ratio of C : H : O atoms = 7 : 6 : 2.

Answer : The empirical formula is $C_7H_6O_2$.

5.13 DETERMINATION OF MOLECULAR FORMULA

The molecular formula of a compound denotes *the actual number of atoms of different elements present in one molecule of the compound.*

Molecular formula of Blue vitriol is $CuSO_4.5H_2O$.

It gives the information that a molecule of blue vitriol is made of –

1. One atom of copper,
2. One atom of sulphur,
3. Four atoms of oxygen,
4. Five molecules of water of crystallization.

Steps to find the *Molecular formula* of a compound :

(1) Calculate the empirical weight of the compound from its empirical formula,

(2) Divide its molecular weight by empirical weight which gives the number (n).

(3) Multiply the empirical formula by this number to get the molecular formula.

Molecular formula = empirical formula × n

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

For example, the empirical formula of hydrogen peroxide is HO and its empirical formula weight is 1 + 16 = 17, but its molecular weight is 34. Hence its molecular formula is 34/17 = 2 times the empirical formula, *i.e.*, H_2O_2.

Note : The value of n must be rounded off to its nearest whole number.

Example 7 : *Calculate the empirical formula of a compound whose molecular formula is $C_8H_6O_4$ and empirical formula weight is 83.*

Solution :

Molecular weight of $C_8H_6O_4$ is

$= 12 \times 8 + 1 \times 6 + 16 \times 4$

$= 96 + 6 + 64$

$= 166$

n(Empirical formula weight) = Molecular weight

$$\therefore \; n = \frac{\text{Molecular Weight}}{\text{Empirical Formula Weight}}$$

$$= \frac{166}{83} = 2$$

i.e. $C_8H_6O_4$ = empirical formula × n [n = 2]

∴ Empirical Formula = $C_4H_3O_2$

Example 8 : *A compound is found to possess C = 40%, H = 6.7% and O = 53.3%. Its molecular mass is 60. Find the molecular formula of the compound?*

Solution :

Element	Percentage composition	Atomic mass	Atomic ratio	Simplest ratio
C	40	12	$\frac{40}{12}$ = 3·3	$\frac{3.3}{3.3} = 1$
H	6·7	1	$\frac{6.7}{1}$ = 6·7	$\frac{6.7}{3.3} = 2$
O	53·3	16	$\frac{53.3}{16}$ = 3·3	$\frac{3.3}{3.3} = 1$

∴ The empirical formula of the compound is CH_2O.

Let the molecular formula of the compound be n (CH_2O).

∴ Molecular mass = n (12 + 2 + 16) = 30n

30n = 60 or n = 2

Answer : The molecular formula of the compound is $C_2H_4O_2$.

Example 9 : Calculate the empirical and molecular formula of the compound having the following percentage composition :

C = 26·59%; H = 2·22%; O = 71·19%

Its vapour density is 45.

Solution :

Element	Percentage composition	Atomic mass	Atomic ratio	Simplest ratio
C	26·59	12	$\frac{26.59}{12}$ = 2·216	$\frac{2.216}{2.216} = 1$
H	2·22	1	$\frac{2.22}{1}$ = 2·22	$\frac{2.22}{2.216} = 1$
O	71·19	16	$\frac{71.19}{16}$ = 4·45	$\frac{4.45}{2.216} = 2$

Hence, the empirical formula is CHO_2

$$\text{Mol. wt.} = n(\text{Empirical formula weight})$$

$$2 \times \text{V.D.} = n(CHO_2)$$

$$\therefore \quad 2 \times 45 = n(12 + 1 + 32)$$

$$90 = n \times 45$$

$$\frac{90}{45} = n \Rightarrow 2 = n$$

Answer : The molecular formula is $C_2H_2O_4$.

Example 10 : 0·5 gram of an organic compound contains 0·062 g of hydrogen and 0·25 g of oxygen. In the vapour state this compound weighs32 times as heavy as the same volume of hydrogen. Determine the molecular formula.

Solution :

We know, every organic compound contains carbon.

$\therefore$ Mass of carbon = 0·5 – (0·25 + 0·062) = 0·188

Element	Mass in grams	Atomic mass	Number of atoms	Simplest ratio
C	0·188	12	$\frac{0.188}{12} = 0.016$	$\frac{0.016}{0.016} = 1$
H	0·062	1	$\frac{0.062}{1} = 0.062$	$\frac{0.062}{0.016} = 4$
O	0·25	16	$\frac{0.25}{16} = 0.016$	$\frac{0.016}{0.016} = 1$

Thus empirical formula is CH_4O

$$\text{Molecular mass} = 2 \times \text{V.D.}$$

$$= 2 \times 32 = 64$$

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

$$= \frac{64}{32} = 2$$

Hence, the molecular formula is $C_2H_8O_2$.

Example 11 : 3·5 g of nitrogen combines with 2 g of oxygen to form an oxide of nitrogen. What is the empirical formula of this oxide?

Solution :

$$\text{Percentage of } N_2 \text{ is } \frac{3.5}{5.5} \times 100 = \frac{700}{11}$$

$$\text{Percentage of } O_2 = \frac{2}{5.5} \times 100 = \frac{400}{11}$$

Element	Percentage	Atomic mass	Number of atoms	Simplest ratio
N	$\frac{700}{11}$	14	$\frac{700}{11 \times 14} = \frac{50}{11}$	$\frac{50/11}{25/11} = 2$
O	$\frac{400}{11}$	16	$\frac{400}{11 \times 16} = \frac{25}{11}$	$\frac{25/11}{25/11} = 1$

Answer : Hence the empirical formula is N_2O.

Example 12 : On heating copper sulphate crystals, $CuSO_4 . xH_2O$ (10g), anhydrous sulphate (6·4g) was left. Write the formula of the crystalline sulphate.

Solution :

	Mass	Molar mass	Moles	Simple ratio
$CuSO_4$	6·4g	159·5g	$\frac{6.4}{159.5} = 0.04$	$\frac{0.04}{0.04} = 1$
H_2O	3·6g	18g	$\frac{3.6}{18} = 0.2$	$\frac{0.2}{0.04} = 5$

Therefore formula is $CuSO_4 \cdot 5H_2O$.

Example 13 : 0·29 grams of a hydrocarbon with vapour density 29 when burnt completely in oxygen produce 448 ml of carbon dioxide at S.T.P. From the given information, calculate the

(i) mass of carbon dioxide formed.

(ii) mass of element carbon in carbon dioxide.
(iii) mass of hydrogen in hydrocarbon.
(iv) empirical formula of hydrocarbon.
(v) molecular formula of hydrocarbon.

Solution :

(i) 1 mole of $CO_2 = 12 + 2 \times 16 = 44$ g.

22400 ml of CO_2 at S.T.P. weighs = 44g

$\therefore$ 448 ml of CO_2 at S.T.P. weighs

$$\frac{44}{22400} \times 448 = 0{\cdot}88 \text{ g}$$

(ii) 44 g of CO_2 contains 12 g carbon.

0·88 g of CO_2 contains $\frac{12}{44} \times 0{\cdot}88 = 0{\cdot}24$ g

(iii) Weight of hydrogen

= wt of hydrocarbon – wt. of carbon

$= 0{\cdot}29 - 0{\cdot}24 = 0{\cdot}05$ g.

(iv)

Element	Mass	At. mass	Gram atom	Ratio	
Carbon	0·24g	12	$\frac{0{\cdot}24}{12} = 0{\cdot}02$	$\frac{0{\cdot}02}{{\cdot}02} = 1$	2
Hydrogen	0·05g	1	$\frac{0{\cdot}05}{1} = 0{\cdot}05$	$\frac{{\cdot}05}{0{\cdot}02} = 2{\cdot}5$	5

Empirical formula is C_2H_5.

(v) Molecular weight = 2 × Vapour density

$= 2 \times 29 = 58$

Molecular wt. = n(Empirical formula wt.)

$58 = n(12 \times 2 + 1 \times 5)$

$\frac{58}{29} = n$

$2 = n$

$\therefore$ Molecular formula = C_4H_{10}

EXERCISE-5C

1. Give three kinds of information conveyed by the formula H_2O.
2. Explain the terms, empirical formula and molecular formula.
3. Give the empirical formula of :
   (a) C_6H_6, (b) $C_6H_{12}O_6$,
   (c) C_2H_2, (d) CH_3COOH.
4. Find the percentage mass of water in the Epsom salt $MgSO_4{\cdot}7H_2O$.
5. Calculate the percentage of phosphorus in :
   (a) Calcium hydrogen phosphate $Ca(H_2PO_4)_2$,
   (b) Calcium phosphate $Ca_3(PO_4)_2$.
6. Calculate the percentage composition of :
   Potassium chlorate, $KClO_3$.
7. Find the empirical formula of the compounds with the following percentage composition. :
   Pb = 62·5%; N = 8·5%; O = 29·0%.
8. Calculate the mass of iron in 10 kg of iron ore which contains 80% of pure ferric oxide.
9. If the empirical formula of two compounds is CH and their vapour densities are 13 and 39 respectively. Find their molecular formula. **(2015)**
10. Find the empirical formula of a compound containing 17·7% hydrogen and 82·3% nitrogen.
11. On analysis, a substance was found to contain :
    C = 54·54%, H = 9·09%, O = 36·36%.
    The vapour density of the substance is 44, calculate :
    (a) its empirical formula,
    (b) its molecular formula.
12. An organic compound, whose vapour density is 45, has the following percentage composition,
    H = 2·22%; O = 71·19%; and remaining carbon.
    Calculate :
    (a) its empirical formula, (b) its molecular formula.
13. An organic compound contains 4·07% hydrogen, 71·65% chlorine and remaining carbon. Its molar mass is 98·96. Find its,
    (a) Empirical formula (b) Molecular formula
14. A hydrocarbon contains 4·8 g of carbon per gram of hydrogen. Calculate :
    (a) the gram atom of each,
    (b) find the empirical formula,
    (c) find molecular formula, if its vapour density is 29.
15. Combine 0·2 g atom of silicon with 21·3 g of chlorine. Find the empirical formula of the compound formed.
16. A gaseous hydrocarbon contains 82·76% of carbon. Given that its vapour density is 29, find its molecular formula. **(2016)**
17. In a compound of magnesium (Mg = 24) and nitrogen (N = 14), 18 g of magnesium combines with 7g of nitrogen.
    Deduce the simplest formula by answering the following questions :

(a) How many gram-atoms of magnesium are equal to 18g ?

(b) How many gram-atoms of nitrogen are equal to 7g of nitrogen ?

(c) Calculate simple ratio of gram-atoms of magnesium to gram-atoms of nitrogen and hence the simplest formula of the compound formed.

18. Barium chloride crystals contain 14·8% water of crystallisation. Find the number of molecules of water of crystallisation per molecule.

19. Urea is a very important nitrogenous fertilizer. Its formula is CON_2H_4. Calculate the percentage of nitrogen in urea (C = 12, O = 16, N = 14 and H = 1).

20. Determine the formula of the organic compound if its molecule contains 12 atoms of carbon. The percentage compositions of hydrogen and oxygen are 6·48 and 51·42 respectively.

21. (a) A compound with empirical formula AB_2, has the vapour density equal to its empirical formula weight. Find its molecular formula.

(b) A compound with empirical formula AB has vapour density three times its empirical formula weight. Find the molecular formula.

22. A hydride of nitrogen contains 87·5 per cent by mass of nitrogen. Determine the empirical formula of this compound.

23. A compound has O = 61·32%, S = 11·15%, H = 4·88% and Zn = 22·65%. The relative molecular mass of the compound is 287 a.m.u. Find the molecular formula of the compound, assuming that all the hydrogen is present as water of crystallisation.

ANSWERS

4. 51·2% **5.** (a) 26·5%, (b) 20% **6.** K = 31·83%, Cl = 28·98%, O_2 = 39·18% **7.** Pb $(NO_3)_2$, **8.** 5·6 kg **9.** C_2H_2, C_6H_6 **10.** NH_3 **11.** (a) C_2H_4O (b) $C_4H_8O_2$ **12.** (a) CHO_2 (b) $C_2H_2O_4$ **13.** (a) CH_2Cl (b) $C_2H_4Cl_2$ **14.** (i) 0·4, 1 (ii) C_2H_5 (iii) C_4H_{10} **15.** $SiCl_3$ **16.** C_4H_{10} **17.** (a) $\frac{3}{4}$ gram atoms (b) $\frac{1}{2}$ gram atoms (c) Mg_3N_2 **18.** 2 **19.** 46·67% **20.** $C_{12}H_{24}O_{12}$ **21.** (a) A_2B_4 (b) A_6B_6 **22.** NH_2. **23.** $ZnSO_4.7H_2O$

5D. CALCULATIONS BASED ON CHEMICAL EQUATIONS

5.14 CHEMICAL EQUATIONS

A chemical equation is a balanced account of a chemical transaction. It is not merely a qualitative statement, but it also gives quantitative information of a chemical reaction.

Chemical equations give information about :

(i) moles and masses of various reactants and products,

(ii) volumes of gaseous reactants and products measured at S.T.P.

Information conveyed by the chemical equation :

Consider the equation :

$$2\ KClO_3(s) \xrightarrow{MnO_2,\ \Delta} 2\ KCl(s) + 3O_2\ (g)$$

(i) ***The molecular proportion of substances:*** Here, 2 molecules of solid potassium chlorate on heating in the presence of manganese dioxide give 2 molecules of solid potassium chloride and 3 molecules of oxygen gas.

(ii) ***The relative masses of substances :*** 2 × 122·5 g = 245 g of potassium chlorate gives 2 × 74·5 g = 149 g of KCl and 3 × 32 = 96 g of oxygen.

(iii) ***The volumes of gaseous substances:*** 3 × 22·4 L = 67·2 L of oxygen at S.T.P. is evolved when 245 g of potassium chlorate is heated.

In solving the problems, proceed as follows:

(i) Write a **balanced** equation for the reaction.

(ii) Write down the molecular weights of each substance below its formula.

(iii) *The same units of weight and volume should be used throughout.*

(iv) Write down the product of 22·4 L and the number of molecules below each gaseous substance.

5.14.1 Problems based on reacting weights

Example 1 : *A solution of common salt when added to silver nitrate solution yields a precipitate of silver chloride (0·28 g). Find the mass of sodium chloride in the solution, and also the mass of sodium nitrate formed.*

Solution :

$NaCl + AgNO_3$	$\rightarrow AgCl$	$+ NaNO_3$
(23 +35·5)	(108 + 35·5)	(23 + 14 + 48)
58·5	143·5	85

Since 143·5 g of silver chloride is formed from 58·5 g of NaCl,

$$\therefore 0{\cdot}28 \text{ g} \quad ,, \quad ,, \quad = \frac{58{\cdot}5 \times 0{\cdot}28}{143{\cdot}5} = \frac{16{\cdot}38}{143{\cdot}5} = 0{\cdot}114 \text{ g of NaCl}$$

Also 58·5 g of sodium chloride forms 85 g of $NaNO_3$.

$$\therefore 0{\cdot}114 \text{ g} \quad ,, \quad ,, \quad = \frac{85 \times 0{\cdot}114}{58{\cdot}5} = \frac{9{\cdot}69}{58{\cdot}5} = 0{\cdot}165 \text{ g of } NaNO_3$$

Answer : 0·114 g NaCl is present in the solution and 0·165 g $NaNO_3$ is formed.

5.14.2 Problems based on mass-volume relationship

Example 2 : *Find the mass of lead formed by the reduction of 342·5 g of red lead (Pb_3O_4) in a current of hydrogen and also find the volume of hydrogen used up at NTP.*

Solution :

Pb_3O_4 +	$4H_2 \rightarrow$	$3Pb$ +	$4H_2O$
207×3 + 16×4	4×22·4 dm^3	3×207	
= 685 g	= 89·6 dm^3	= 621 g	

685 g of red lead yields 621 g of lead.

$$\therefore 342{\cdot}5\text{g} \quad ,, \quad ,, \quad ,, = \frac{621 \times 342{\cdot}5}{685} = 310{\cdot}5 \text{ g of lead}$$

Also, 685g of red lead uses 89·6 dm^3 of hydrogen at NTP.

$\therefore$ 342·5 ,, ,, dm^3

= 44·8 dm^3 of hydrogen.

Answer : 310·5 g of lead is formed and 44·8 dm^3 of hydrogen at NTP is used up.

Example 3 : *If 16·4 grams of calcium nitrate is heated, calculate the volume of (a) nitrogen dioxide obtained at stp and (b) the mass of calcium oxide obtained.*

Solution :

$$2Ca(NO_3)_2 \rightarrow 2CaO + 4NO_2 + O_2$$

2[40 + 2(14+48)] = 328 g; 112 g; 4 moles

(a) 328 g of $Ca(NO_3)_2$ liberates 4 × 22·4 L NO_2.

$$\therefore 16{\cdot}4 \text{ g will liberate } \frac{4 \times 22{\cdot}4}{328} \times 16{\cdot}4 = 4{\cdot}48 \text{ L } NO_2 \text{ at STP}$$

(b) 328 g of calcium nitrate gives 112 g of CaO ∴

$$16{\cdot}4 \text{ g will give } = \frac{112 \times 16{\cdot}4}{328} = 5{\cdot}6 \text{ g of CaO.}$$

Answer : (a) 4·48 litres of NO_2.

(b) 5·6 g of CaO.

Example 4 : *On heating 12·4 g of copper (II) carbonate in a crucible only 7·0 g of copper (II) oxide was produced. What was the percentage yield of copper (II) oxide ? [at. wt. Cu = 64, C = 12, O = 16]*

Solution :

Equation for the reaction

$$\underbrace{\underset{64 + 12 + 48}{CuCO_3}}_{124\text{ g}} \longrightarrow \underbrace{\underset{64 + 16}{CuO}}_{80\text{ g}} + CO_2$$

124 g of copper carbonate produces 80 g of copper oxide, on heating.

∴ 124 g of copper carbonate produces $\frac{80}{124} \times 12{\cdot}4 = \mathbf{8\ g}$

Thus expected yield is 8·0 g

Actual yield is 7·0 g

$$\therefore \text{ Percentage yield} = \frac{7}{8} \times 100 = 87{\cdot}5\%$$

Some difficult problems, not in syllabus

Example 5 : *Find how many grams of zinc must be dissolved in dilute hydrochloric acid to produce 300 cm³ of hydrogen at 27°C and 700 mm pressure.*

Solution :

First convert the volume of hydrogen to NTP.

$P = 700$ mm; $V = 300$ cm³;

$T = (273 + 27)$ K = 300 K

$P' = 760$ mm; $V' = ?$

$T' = (273 + 0)$ K = 273 K

By using the gas equation,

$$\frac{PV}{T} = \frac{P'V'}{T'}$$

$$\text{or} \quad \frac{700 \times 300}{300} =$$

$$\text{or} \quad 300 \times 760V' = 700 \times 300 \times 273$$

$$\therefore \quad V' = \frac{700 \times 300 \times 273}{300 \times 760}$$

$$= \frac{19110}{76} = 251{\cdot}5 \text{ cm}^3$$

$$Zn + 2HCl = ZnCl_2 + H_2$$

65 g   22·4 dm³.

22400 cm³.

Since 22400 cm³ of hydrogen at NTP are obtained from 65 g of Zn.

$$\therefore 251{\cdot}5 \text{ cm}^3 \quad " \quad " = \frac{65 \times 251{\cdot}5}{22400} \text{ g}$$

$$= \frac{16347{\cdot}5}{22400} = 0{\cdot}729 \text{ g of zinc.}$$

Answer : 0·729 g of zinc must be dissolved to produce the given volume of hydrogen.

Example 6 : *The given sample of hydrochloric acid contains only 20% by weight of HCl, the rest being water. What mass of the acid will be required for reacting completely with 20 g of magnesium oxide ?*

Solution :

$$\underset{(24 + 16)}{MgO} + \underset{2(1 + 35{\cdot}5)}{2HCl} \rightarrow MgCl_2 + H_2O$$

40 g of MgO requires 73 g of HCl.

$$\therefore \quad 20 \text{ g of MgO will require} = \frac{73}{40} \times 20$$

$$= 36{\cdot}5 \text{ g of HCl.}$$

But the given acid is 20%. *i.e.* 20 g of HCl is contained in 100 g sample.

$\therefore$ 36·5 g of HCl is contained in $\frac{100}{20} \times 36{\cdot}5$ g of HCl. = 182·5 g of HCl

Answer : 182·5 g of HCl.

Example 7 : *Hydrogen is generated by the action of steam with heated magnesium. Calculate the mass of magnesium which will produce just sufficient hydrogen to combine with all the oxygen that can be obtained by the complete decomposition of 24·5 g of $KClO_3$.*

Solution :

$$Mg + H_2O \rightarrow MgO + H_2 \quad (1)$$

$$2H_2 + O_2 \rightarrow 2H_2O \quad (2)$$

$$2KClO_3 \rightarrow 2KCl + 3O_2 \quad (3)$$

i.e., $2KClO_3 = 3O_2 = 6H_2 = 6Mg$

2 moles of $KClO_3$ correspond to 6 moles of Mg

or 1 mole of $KClO_3$ (122·5 g) corresponds to

3 × 24 g of Mg.

$\therefore$ 24·5 g of $KClO_3$ will correspond to

$$\frac{72}{122{\cdot}5} \times 24{\cdot}5 = 14{\cdot}4 \text{ g of Mg.}$$

Answer : 14·4 g of Mg.

Example 8. *Find the mass of quick lime (calcium oxide) formed by the decomposition of 200 g of limestone. What mass of carbon dioxide will be evolved ? If the limestone is only 90% pure, what mass of quicklime will be produced ?*

Solution :

$$\underset{(40 + 12 + 48)}{CaCO_3} \rightarrow \underset{(40 + 16)}{CaO} + \underset{(12 + 32)}{CO_2}$$

100 g of limestone on decomposition produces 56 g of CaO.

$$\therefore \quad 200 \text{ g} \quad " \quad " = \frac{56 \times 200}{100}$$

$$= 112 \text{ g}$$

Also, 100 g of limestone evolves 44 g of CO_2.

$$\therefore \quad 200 \text{ g} \quad " \quad " = 88 \text{ g of } CO_2$$

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If the limestone is only 90% pure, weight of pure calcium carbonate in 200 g will be

$$= \frac{90 \times 200}{100} = 180 \text{ g}$$

Since 100 g of pure calcium carbonate produces 56 g of CaO.

$$\therefore \quad 180 \text{ g} \quad " \quad " = \frac{56 \times 180}{100} = 100{\cdot}8 \text{ g}$$

Answer : 200 g pure limestone would produce 112 g quicklime and 88 g carbon dioxide. Also, 200 g of 90% pure limestone would produce 100·8 g quicklime.

EXERCISE-5D

Note : Atomic mass wherever required should be taken from the table given on the preliminary pages, rounded off to the nearest whole number.

1. Complete the following blanks in the equation as indicated.

$CaH_2\ (s) + 2H_2O\ (aq) \rightarrow Ca(OH)_2\ (s) + 2H_2\ (g)$

(a) Moles : 1 mole + ... → .... + ...

(b) Grams : 42 g + ... → .... + ...

(c) Molecules : $6{\cdot}02 \times 10^{23}$ + ... → .... + ...

2. The reaction between 15 g of marble and nitric acid is given by the following equation :

$$CaCO_3 + 2HNO_3 \rightarrow Ca(NO_3)_2 + H_2O + CO_2$$

Calculate :

(a) the mass of anhydrous calcium nitrate formed,

(b) the volume of carbon dioxide evolved at STP.

3. 66 g ammonium sulphate is produced by the action of ammonia on sulphuric acid.

Write a balanced equation and calculate :

(a) mass of ammonia required,

(b) the volume of the gas used at STP,

(c) the mass of acid required.

4. The reaction between red lead and hydrochloric acid is given below :

$$Pb_3O_4 + 8HCl \rightarrow 3PbCl_2 + 4H_2O + Cl_2$$

Calculate : (a) the mass of lead chloride formed by the action of 6.85 g of red lead, (b) the mass of chlorine and (c) the volume of chlorine evolved at STP.

5. Find the mass of KNO_3 required to produce 126 kg of nitric acid. Find whether a larger or smaller mass of $NaNO_3$ is required for the same purpose.

$$KNO_3 + H_2SO_4 \rightarrow KHSO_4 + HNO_3$$
$$NaNO_3 + H_2SO_4 \rightarrow NaHSO_4 + HNO_3$$

6. Pure calcium carbonate and dilute hydrochloric acid are reacted and 2 litres of carbon dioxide was collected at 27°C and normal pressure.

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

Calculate : (a) the mass of salt required,

(b) the mass of the acid required.

7. Calculate the mass and volume of oxygen at STP, which will be evolved on electrolysis of 1 mole (18 g) of water.

8. 1·56 g of sodium peroxide reacts with water according to the following equation :

$$2Na_2O_2 + 2H_2O \rightarrow 4NaOH + O_2$$

Calculate : (a) mass of sodium hydroxide formed,

(b) volume of oxygen liberated at STP,

(c) mass of oxygen liberated.

9. (a) Calculate the mass of ammonia that can be obtained from 21·4 g of NH_4Cl by the reaction :

$$2NH_4Cl + Ca(OH)_2 \rightarrow CaCl_2 + 2H_2O + 2NH_3$$

(b) What will be the volume of ammonia when measured at STP ?

10. The reaction between aluminium carbide and water takes place according to the following equation :

$$Al_4C_3 + 12\ H_2O \rightarrow 3CH_4 + 4Al\ (OH)_3.$$

Calculate the volume of methane measured at STP released from 14.4 g of aluminium carbide by excess of water.

11. $MnO_2 + 4HCl \rightarrow MnCl_2 + 2H_2O + Cl_2$

0·02 moles of pure MnO_2 is heated strongly with conc. HCl. Calculate :

(a) mass of MnO_2 used,

(b) moles of salt formed,

(c) mass of salt formed,

(d) moles of chlorine gas formed,

(e) mass of chlorine gas formed,

(f) volume of chlorine gas formed at STP,

(g) moles of acid required,

(h) mass of acid required.

12. Nitrogen and hydrogen react to form ammonia.

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

If 1000 g of H_2 react with 2000 g of N_2.

(a) Will any of the two reactants remain unreacted ? If yes, which one and what will be its mass ?

(b) Calculate the mass of ammonia (NH_3) that will be formed.

ANSWERS

2. (a) 24·6 g (b) 3·36 L **3.** (a) 17 g (b) 22·4 litres (c) 49 g. **4.** (a) 8·34 g, (b) 0·71 g, (c) 0·224 L **5.** 202 Kg KNO_3, 170 Kg $NaNO_3$ *i.e.*, smaller mass of $NaNO_3$ is required **6.** (a) 8·125 g (b) 5·93 g **7.** 16 g, 11·2 L **8.** (a) 1·6 g, (b) 224 cm^3, (c) 0·32 g **9.** (a) 6·8 g, (b) 8·96 L **10.** 6·72 L **11.** (a) 1·74 grams, (b) 0·02 moles, (c) 2·52 g, (d) 0·02 moles, (e) 1·42 g, (f) 0·448 dm^3 at stp, (g) 0·08 moles, (h) 2·92 g **12.** (a) 571·4 g H_2 remains unreacted, (b) 2428·8 g.

MISCELLANEOUS EXERCISES

1. From the equation for burning of hydrogen and oxygen

$$2H_2 + O_2 \rightarrow 2H_2O \text{ (steam)}$$

Write down the number of mole (or moles) of steam obtained from 0·5 moles of oxygen.

2. From the equation,

$$3Cu + 8HNO_3 \rightarrow 3Cu\,(NO_3)_2 + 4H_2O + 2NO$$

Calculate :

(a) mass of copper needed to react with 63 g of HNO_3,

(b) volume of nitric oxide at STP that can be collected.

3. (a) Calculate the number of moles in 7 g of nitrogen.

(b) What is the volume at STP of 7·1 g of chlorine?

(c) What is the mass of 56 cm^3 of carbon monoxide at STP?

4. Some of the fertilizers are sodium nitrate $NaNO_3$, ammonium sulphate $(NH_4)_2SO_4$ and urea $CO(NH_2)_2$. Which of these contains the highest percentage of nitrogen ?

5. Water decomposes to O_2 and H_2 under suitable conditions as represented by the equation below :

$$2\,H_2O \rightarrow 2H_2 + O_2$$

(a) If 2500 cm^3 of H_2 is produced, what volume of O_2 is liberated at the same time and under the same conditions of temperature and pressure ?

(b) The 2500 cm^3 of H_2 is subjected to $2\frac{1}{2}$ times increase in pressure (temp. remaining constant). What volume of H_2 will now occupy ?

(c) Taking the volume of H_2 calculated in 5(b), what changes must be made in kelvin (absolute) temperature to return the volume to 2500 cm^3 pressure remaining constant.

6. Urea $[CO(NH_2)_2]$ is an important nitrogeneous fertilizer, and is sold in 50 kg sacks. What mass of nitrogen is in one sack of urea ?

7. Find the molecular formula of a hydrocarbon having vapour density 15, which contains 20% of Hydrogen.

8. The following experiment was performed in order to determine the formula of a hydrocarbon. The hydrocarbon X is purified by fractional distillation.

0·145 g of X were heated with dry copper (II) oxide and 224 cm^3 of carbon dioxide was collected at STP.

(a) Which elements does X contain ?

(b) What was the purpose of Copper (II) oxide ?

(c) Calculate the empirical formula of X by the following steps :

(i) Calculate the number of moles of carbon dioxide gas.

(ii) Calculate the mass of carbon contained in this quantity of carbon dioxide and thus the mass of carbon in sample X.

(iii) Calculate the mass of hydrogen in sample X.

(iv) Deduce the ratio of atoms of each element in X (empirical formula).

9. A compound is formed by 24 g of X and 64 g of oxygen. If atomic mass of X = 12 and O = 16, calculate the simplest formula of compound.

10. A gas cylinder filled with hydrogen holds 5 g of the gas. The same cylinder holds 85 g of gas X under same temperature and pressure. Calculate :

(a) vapour density of gas X,

(b) molecular weight of gas X.

11. (a) When carbon dioxide is passed over red hot carbon, carbon monoxide is produced according to the equation :

$$CO_2 + C \rightarrow 2CO$$

What volume of carbon monoxide at STP can be obtained from 3 g of carbon ?

(b) 60 cm^3 of oxygen was added to 24 cm^3 of carbon monoxide and mixture ignited. Calculate :

(i) volume of oxygen used up and

(ii) volume of carbon dioxide formed.

12. How much calcium oxide is obtained by heating 82 g of calcium nitrate ? Also find the volume of NO_2 evolved :

$$2Ca(NO_3)_2 \rightarrow 2CaO + 4NO_2 + O_2 \qquad \textbf{(2016)}$$

13. The equation for the burning of octane is :

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$$

(i) How many moles of carbon dioxide are produced when **one** mole of octane burns?

(ii) What volume, at STP, is occupied by the number of moles determined in (i) ?

(iii) If the relative molecular mass of carbon dioxide is 44, what is the mass of carbon dioxide produced by burning **two** moles of octane ?

(iv) What is the empirical formula of octane ?

14. Ordinary chlorine gas has two isotopes $^{35}_{17}Cl$ and $^{37}_{17}Cl$ in the ratio of 3 : 1. Calculate the relative atomic mass of chlorine.

15. Silicon (Si = 28) forms a compound with chlorine (Cl = 35·5) in which 5·6 g of silicon combines with 21·3 g of chlorine. Calculate the empirical formula of the compound.

16. An acid of phosphorus has the following percentage composition; Phosphorus = 38·27%; hydrogen = 2·47%; oxygen = 59·26%. Find the empirical formula of the acid and its molecular formula, given that its relative molecular mass is 162.

17. Calculate the mass of substance 'A' which in gaseous form occupies 10 litres at 27°C and 700 mm pressure. The molecular mass of 'A' is 60.

18. A gas cylinder can hold 1 kg of hydrogen at room temperature and pressure. (a) What mass of carbon dioxide can it hold under similar conditions of temperature and pressure? (b) If the number of molecules of hydrogen in the cylinder is X, calculate the number of carbon dioxide molecules in the cylinder. Give reasons for your answer.

19. Following questions refer to one mole of chlorine gas.

(a) What is the volume occupied by this gas at STP?

(b) What will happen to volume of gas, if pressure is doubled ?

(c) What volume will it occupy at 273°C ?

(d) If the relative atomic mass of chlorine is 35·5, what will be the mass of 1 mole of chlorine gas?

20. (a) A hydrate of calcium sulphate $CaSO_4 \cdot xH_2O$ contains 21% water of crystallisation. Find the value of x.

(b) What volume of hydrogen and oxygen measured at STP will be required to prepare 1·8 g of water.

(c) How much volume will be occupied by 2 g of dry oxygen at 27°C and 740 mm pressure ?

(d) What would be the mass of CO_2 occupying a volume of 44 litres at 25°C and 750 mm pressure.

(e) 1 g of a mixture of sodium chloride and sodium nitrate is dissolved in water. On adding silver nitrate solution, 1·435 g of AgCl is precipitated. $AgNO_3$ (aq) + NaCl (aq) $\rightarrow$ AgCl (s) + $NaNO_3$

Calculate the precentage of NaCl in the mixture.

21. (a) From the equation :

$$C + 2H_2SO_4 \rightarrow CO_2 + 2H_2O + 2SO_2$$

Calculate :

(i) the mass of carbon oxidized by 49 g of sulphuric acid.

(ii) The volume of sulphur dioxide measured at STP, liberated at the same time.

(b) (i) A compound has the following percentage composition by mass : carbon 14·4%, hydrogen 1·2% and chlorine 84·5%. Determine the empirical formula of this compound. Work correct to 1 decimal place. (H = 1; C = 12; Cl = 35·5)

(ii) The relative molecular mass of this compound is 168, so what is its molecular formula ?

22. Find the percentage of

(a) oxygen in magnesium nitrate crystals [$Mg(NO_3)_2 \cdot 6H_2O$].

(b) boron in $Na_2B_4O_7.10H_2O$. [H = 1, B = 11, O = 16, Na = 23].

(c) phosphorus in the fertilizer superphosphate $Ca(H_2PO_4)_2$.

23. A gas occupied 360 cm^3 at 87°C and 380 mm Hg pressure. If the mass of gas is 0·546 g, find its relative molecular mass.

24. Solid ammonium dichromate decomposes as under :

$$(NH_4)_2Cr_2O_7 \rightarrow N_2 + Cr_2O_3 + 4H_2O$$

If 63 g of ammonium dichromate decomposes. Calculate:

(a) the quantity in moles of $(NH_4)_2Cr_2O_7$

(b) the quantity in moles of nitrogen formed

(c) the volume of N_2 evolved at STP.

(d) what will be the loss of mass ?

(e) calculate the mass of chromium (III) oxide formed at the same time. **(2015)**

25. Hydrogen sulphide gas burns in oxygen to yield 12·8 g of sulphur dioxide gas as under :

$$2H_2S + 3O_2 \rightarrow 2H_2O + 2SO_2$$

Calculate the volume of hydrogen sulphide at STP.

Also, calculate the volume of oxygen required at STP. which will complete the combustion of hydrogen sulphide determined in (litres).

26. Ammonia burns in oxygen and the combustion, in the presence of a catalyst, may be represented by; $2NH_3 + 2\frac{1}{2}\ O_2 \rightarrow 2NO + 3H_2O$ [H = 1; N = 14; O = 16]. What mass of steam is produced when 1·5 g of nitrogen monoxide is formed ?

27. If a crop of wheat removes 20 kg of nitrogen per hectare of soil, what mass of the fertilizer, calcium nitrate $Ca(NO_3)_2$ would be required to replace the nitrogen in a 10 hectare field ?

28. Concentrated nitric acid oxidises phosphorus to phosphoric acid according to the following equation:

$$P + 5HNO_3 \rightarrow H_3PO_4 + 5NO_2 + H_2O$$

If 6·2 g of phosphorus was used in the reaction calculate:

(a) Number of moles of phosphorus taken and mass of phosphoric acid formed.

(b) What mass of nitirc acid will be consumed at the same time ?

(c) What would be the volume of steam produced at the same time if measured at 760 mm Hg pressure and 273°C ?

29. 112 cm^3 of a gaseous fluoride of phosphorus has a mass of 0·63 g. Calculate the relative molecular mass of the fluoride. If the molecule of the fluoride contains only one atom of phosphorus, then determine the formula of the phosphorus fluoride. [F = 19; P = 31].

30. Washing soda has the formula $Na_2CO_3.10H_2O$. What mass of anhydrous sodium carbonate is left when all the water of crystallization is expelled by heating 57·2 g of washing soda ?

31. A metal M forms a volatile chloride containing 65·5% chlorine. If the density of the chloride relative to hydrogen is 162·5, find the molecular formula of the chloride (M=56).

32. A compound X consists of 4·8% carbon and 95·2% bromine by mass :

(i) Determine the empirical formula of this compound working correct to one decimal place (C = 12; Br = 80).

(ii) If the vapour density of the compound is 252, what is the molecular formula of the compound ?

33. The reaction : $4N_2O + CH_4 \rightarrow CO_2 + 2H_2O + 4N_2$ takes place in the gaseous state. If all volumes are measured at the same temperature and pressure, calculate the volume of dinitrogen oxide (N_2O) required to give 150 cm^3 of steam.

34. Samples of the gases O_2, N_2, CO_2 and CO under the same conditions of temperature and pressure contain the same number of molecules x. The molecules of oxygen occupy V litres and have a mass of 8 g under the same conditions of temperature and pressure.

What is the volume occupied by :

(a) x molecules of N_2,

(b) $3x$ molecules of CO,

(c) What is the mass of CO_2 in grams ?

(d) In answering the above questions, which law have you used ?

35. The percentage composition of sodium phosphate as determined by analysis is 42·1% sodium, 18·9% phosphorus and 39% oxygen. Find the empirical formula of the compound.

36. What volume of oxygen is required to burn completely a mixture of 22·4 dm^3 of methane and 11·2 dm^3 of hydrogen into carbon dioxide and steam ?

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$
$$2H_2 + O_2 \rightarrow 2H_2O$$

37. The gases hydrogen, oxygen, carbon dioxide, sulphur dioxide and chlorine are arranged in order of their increasing relative molecular masses. Given 8 g of each gas at STP, which gas will contain the least number of molecules and which gas the most ?

38. 10 g of a mixture of sodium chloride and anhydrous sodium sulphate is dissolved in water. An excess of barium chloride solution is added and 6·99 g of barium sulphate is precipitated according to the equation given below :

$$Na_2SO_4 + BaCl_2 \rightarrow BaSO_4 + 2NaCl$$

Calculate the percentage of sodium sulphate in the original mixture.

39. When heated, potassium permanganate decomposes according to the following equation :

$$2KMnO_4 \rightarrow \underbrace{K_2MnO_4 + MnO_2}_{\text{solid residue}} + O_2$$

(a) Some potassium permanganate was heated in the test tube. After collecting one litre of oxygen at room temperature, it was found that the test tube had undergone a loss in mass of 1·32 g. If one litre of hydrogen under the same conditions of temperature and pressure has a mass of 0·0825 g, calculate the relative molecular mass of oxygen.

(b) Given that the molecular mass of potassium permanganate is 158, what volume of oxygen (measured at room temperature) would be obtained by the complete decomposition of 15·8 g of potassium permanganate ? (Molar volume at room temperature is 24 litres).

40. (a) A flask contains 3·2 g of sulphur dioxide. Calculate the following :

(i) The moles of sulphur dioxide present in the flask.

(ii) The number of molecules of sulphur dioxide present in the flask.

(iii) The volume occupied by 3·2 g of sulphur dioxide at STP.

(S = 32, O = 16)

(b) An experiment showed that in a lead chloride solution, 6·21 g of lead is combined with 4·26 g of chlorine. What is the empirical formula of this chloride ? (Pb = 207; Cl = 35·5).

41. The volumes of gases A, B, C and D are in the ratio, 1 : 2 : 2 : 4 under the same conditions of temperature and pressure.

(i) Which sample of gas contains the maximum number of molecules ?

(ii) If the temperature and the pressure of gas A are kept constant, then what will happen to the volume of A when the number of molecules is doubled ?

(iii) If this ratio of gas volume refers to the reactants and products of a reaction, which gas law is being observed ?

(iv) If the volume of A is actually 5·6 dm^3 at STP, calculate the number of molecules in the actual volume of D at STP (Avogadro's Number is 6 x 10^{23}).

(v) Using your answer from (iv), state the mass of D if the gas is dinitrogen oxide (N_2O).

42. The equation given below relate to the manufacture of sodium carbonate (molecular weight of Na_2CO_3 = 106).

1. $NaCl + NH_3 + CO_2 + H_2O \rightarrow NaHCO_3 + NH_4Cl$

2. $2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$

Equations (1) and (2) are based on the production of 21·2 g of sodium carbonate.

(a) What mass of sodium hydrogen carbonate must be heated to give 21·2 g of sodium carbonate?

(b) To produce the mass of sodium hydrogen carbonate calculated in (a), what volume of carbon dioxide, measured at STP, would be required ?

43. A sample of ammonium nitrate when heated yields 8·96 litres of steam (measure at STP).

$$NH_4NO_3 \rightarrow N_2O + 2H_2O$$

(i) What volume of dinitrogen oxide is produced at the same time as 8·96 litres of steam ?

(ii) What mass of ammonium nitrate should be heated to produce 8·96 litres of steam ? (Relative molecular mass of ammonium nitrate is 80).

(iii) Determine the percentage of oxygen in ammonium nitrate (O = 16).

44. Given that the relative molecular mass of copper oxide is 80, what volume of ammonia (measured at STP) is required to completely reduce 120g of copper oxide?

The equation for the reaction is :

$$3CuO + 2NH_3 \rightarrow 3Cu + 3H_2O + N_2$$

45. (a) Calculate the number of moles and the number of molecules present in 1·4 g of ethylene gas. What is the volume occupied by the same amount of ethylene?

(b) What is the vapour density of ethylene ?

46. (a) Calculate the percentage of sodium in sodium aluminium fluoride (Na_3AlF_6) correct to the nearest whole number. (F = 19; Na = 23; Al = 27)

(b) 560 ml of carbon monoxide is mixed with 500 ml of oxygen and ignited. The chemical equation for the reaction is as follows :

$$2\,CO + O_2 \rightarrow 2\,CO_2$$

Calculate the volume of oxygen used and carbon dioxide formed in the above reaction.

2009

47. (a) A gas cylinder of capacity of 20 dm^3 is filled with gas X the mass of which is 10 g. When the same cylinder is filled with hydrogen gas at the same temperature and pressure the mass of the hydrogen is 2 g, hence the relative molecular mass of the gas is :

(i) 5 (ii) 10

(iii) 15 (iv) 20

(b) (i) Calcium carbide is used for the artificial ripening of fruits. Actually the fruit ripens because of the heat evolved while calcium carbide reacts with moisture. During this reaction calcium hydroxide and acetylene gas is formed. If 200 cm^3 of acetylene is formed from a certain mass of calcium carbide, find the volume of oxygen required and carbon dioxide formed during the complete combustion. The combustion reaction can be represented as below.

$$2C_2H_{2(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(g)}$$

(ii) A gaseous compound of nitrogen and hydrogen contains 12·5% hydrogen by mass. Find the molecular formula of the compound if its relative molecular mass is 37.

[N = 14, H = 1].

(c) (i) A gas cylinder contains 24×10^{24} molecules of nitrogen gas. If Avogadro's number is 6×10^{23} and the relative atomic mass of nitrogen is 14, calculate :

(1) Mass of nitrogen gas in the cylinder

(2) Volume of nitrogen at STP in dm^3.

(ii) Commercial sodium hydroxide weighing 30 g has some sodium chloride in it. The mixture on dissolving in water and subsequent treatment with excess silver nitrate solution formed a precipitate weighing 14·3 g. What is the percentage of sodium chloride in the commercial sample of sodium hydroxide ? The equation for the reaction is

$$NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$$

[Relative molecular mass of NaCl = 58; AgCl = 143]

(iii) A certain gas 'X' occupies a volume of 100 cm^3 at S.T.P. and weighs 0·5 g. Find its relative molecular mass.

2010

48. (a) (i) LPG stands for liquefied petroleum gas. Varieties of LPG are marketed including a mixture of propane (60%) and butane (40%). If 10 litre of this mixture is burnt, find the total volume of carbon dioxide gas added to the atmosphere. Combustion reactions can be represented as :

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$
$$2C_4H_{10(g)} + 13O_{2(g)} \rightarrow 8CO_{2(g)} + 10H_2O_{(g)}$$

(ii) Calculate the percentage of nitrogen and oxygen in ammonium nitrate. [Relative molecular mass of ammonium nitrate is 80, H = 1, N = 14, O = 16].

(b) 4·5 moles of calcium carbonate are reacted with dilute hydrochloric acid.

(i) Write the equation for the reaction.

(ii) What is the mass of 4·5 moles of calcium carbonate ? (Relative molecular mass of calcium carbonate is 100).

(iii) What is the volume of carbon dioxide liberated at STP ?

(iv) What mass of calcium chloride is formed ? (Relative molecular mass of calcium chloride is 111).

(v) How many moles of HCl are used in this reaction ?

2011

49. (a) (i) Calculate the volume of 320 g of SO_2 at STP. (Atomic mass : S = 32 and O = 16).

(ii) State Gay-Lussac's Law of combining volumes.

(iii) Calculate the volume of oxygen required for the complete combustion of 8·8 g of propane (C_3H_8). (Atomic mass : C = 12, O = 16, H = 1, Molar Volume = 22·4 dm^3 at stp.)

(b) (i) An organic compound with vapour density = 94 contains

C = 12·67%, H = 2·13%, and Br = 85·11%. Find the molecular formula.

[Atomic mass : C = 12, H = 1, Br = 80]

(ii) Calculate the mass of :

(1) 10^{22} atoms of sulphur.

(2) 0·1 mole of carbon dioxide.

[Atomic mass : S = 32, C = 12 and O = 16 and Avogadro's Number = 6×10^{23}]

2012

50. (a) Concentrated nitric acid oxidises phosphorus to phosphoric acid according to the following equation:

$$P + 5HNO_3 \text{ (conc.)} \rightarrow H_3PO_4 + H_2O + 5NO_2$$

If 9·3 g of phosphorus was used in the reaction, calculate :

(i) Number of moles of phosphorus taken.

(ii) The mass of phosphoric acid formed.

(iii) The volume of nitrogen dioxide produced at STP.

(b) (i) 67·2 litres of hydrogen combines with 44·8 litres of nitrogen to form ammonia under specific conditions as :

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

Calculate the volume of ammonia produced. What is the other substance, if any, that remains in the resultant mixture ?

(ii) The mass of 5·6 dm^3 of a certain gas at STP is 12·0 g. Calculate the relative molecular mass of the gas.

(iii) Find the total percentage of Magnesium in magnesium nitrate crystals, $Mg(NO_3)_2 \cdot 6H_2O$.

[Mg = 24, N = 14, O = 16 and H = 1]

2013

51. (a) (i) What volume of oxygen is required to burn completely 90 dm^3 of butane under similar conditions of temperature and pressure ?

$$2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$$

(ii) The vapour density of a gas is 8. What would be the volume occupied by 24·0 g of the gas at STP ?

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(iii) A vessel contains **X** number of molecules of hydrogen gas at a certain temperature and pressure. How many molecules of nitrogen gas would be present in the same vessel under the same conditions of temperature and pressure?

(b) O_2 is evolved by heating $KClO_3$ using MnO_2 as a catalyst.

$$2KClO_3 \xrightarrow{MnO_2} 2KCl + 3O_2$$

(i) Calculate the mass of $KClO_3$ required to produce 6·72 litre of O_2 at STP.
[atomic masses of K = 39, Cl = 35·5, O = 16].

(ii) Calculate the number of moles of oxygen present in the above volume and also the number of molecules.

(iii) Calculate the volume occupied by 0·01 mole of CO_2 at STP.

2014

52. (a) (i) Oxygen oxidises ethyne to carbon dioxide and water as shown by the equation:

$$2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$$

What volume of ethyne gas at s.t.p. is required to produce 8·4 dm^3 of carbon dioxide at STP ?
[H = 1, C = 12, O = 16]

(ii) A compound made up of two elements X and Y has an empirical formula X_2Y. If the atomic weight of X is 10 and that of Y is 5 and the compound has a vapour density (V.D.) 25, find its molecular formula.

(b) A cylinder contains 68 g of Ammonia gas at STP

(i) What is the volume occupied by this gas ?

(ii) How many moles of ammonia are present in the cylinder ?

(iii) How many molecules of ammonia are present in the cylinder ?

ANSWERS

1. 1 mole **2.** (a) 24 gm (b) 5·6 L **3.** (a) 0·25 moles (b) 2·24 dm^3 (c) 0·07g **4.** Urea **5.** (a) 1250 cm^3 (b) $\frac{5000}{7}$ cm^3 (c) increased to 3·5 times **6.** 23·3 kg **7.** (a) CH_3 (b) C_2H_6 **8.** (c) (i) 0·01 moles (ii) 0·12 g (iii) 0·025 g (iv) C_2H_5 **9.** XO_2 **10.** (a) 17 (b) 34 a.m.u. **11.** (a) 11·2 L (b) (i) 12 cm^3 (ii) 24 cm^3 **12.** (a) 28g, 22·4 *l* **13.** (a) (i) 3g (ii) 11·2 *l*, (b) (i) $CHCl_2$ (ii) $C_2H_2Cl_4$ **14.** 35·5 **15.** $SiCl_3$ **16.** H_2PO_3 ; $H_4P_2O_6$ **17.** 22·45 g **18.** (a) 22 kg (b) X **19.** (a) 22·4 dm^3 (b) 11·2 dm^3 (c) 44·8 dm^3 (d) 71 g **20.** (a) 2 (b) 2·24 *l*, 1·12 *l* (c) 1·58 *l* (d) 78·14 g (e) 58·5 **21.** (i) 8 moles (ii) 179·2 dm^3 (iii) 704 g (iv) C_4H_9 **22.** (a) 75%, (b) 11·5%, (c) 26·5% **23.** 89·6 a.m.u. **24.** (a) 0·25 moles (b) 0·25 moles (c) 5·6 litres (d) 25 g (e) 38 g **25.** H_2S = 4·48 L O_2 = 6·72 L **26.** 1·35 g **27.** 1171·4 kg **28.** (a) 0·2 mole, 19·6 g (b) 63 g (c) 8·96 L **29.** 126, PF_5 **30.** 21·2 g **31.** M_2Cl_6 **32.** (i) CBr_3 (ii) C_2Br_6 **33.** 300 cc **34.** (a) 1 V litres (b) 3 V litres (c) 11g **35.** Na_3PO_4. **36.** 44·8 dm^3 + 5·6 dm^3 = 50·4 dm^3. **37.** **Hydrogen** : maximum; **chlorine** : minimum. **38.** 42·6% **39.** (a) 32 (b) 1·2 L **40.** (a) (i) 0·05 (ii) 3×10^{22} molecules (iii) 1·12 L (b) $PbCl_4$ **41.** (i) D (ii) 2 V (iii) Gay-Lussac's Law (iv) 6×10^{23} (v) 44 **42.** (a) 33·6g (b) 8·96 dm^3 **43.** (i) 4·48 litres (ii) 16 g (iii) 60% **44.** 22·4 litres **45.** (a) 0·05 moles, 3×10^{22} molecules, 1·12 litres (b) 14 **46.** (a) 32·8% (b) oxygen 280 ml, carbon dioxide 560 ml **47.** (a) 10 (b) (i) 500 cc, 400 cc (ii) N_2H_4 (c) (1) 1120 g (2) 896 dm^3 (ii) 19·33% (iii) 112 g **48.** (a) (i) 18 + 16 = 34 litres (ii) N = 35%, O = 60% (b) (i) $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$ (ii) 450 g (iii) 100·8 litres (iv) 499·5 g (v) 9 moles of HCl **49.** (a) (i) 112 litres (ii) 22·4 litres (b) (i) $C_2H_4Br_2$ (ii) (1) 0·533 g (2) 4·4 g **50.** (a) (i) 585 dm^3 (ii) 33·6 litres (iii) X number of molecules (b) (i) 24·5 g (ii) 0·3, $1·806 \times 10^{23}$ (iii) 0·224 litres **52.** (a) (i) 4·2 dm^3 (ii) X_4Y_2 (b) (i) 89·6 dm^3 (ii) 4 moles (iii) $2·4 \times 10^{24}$.