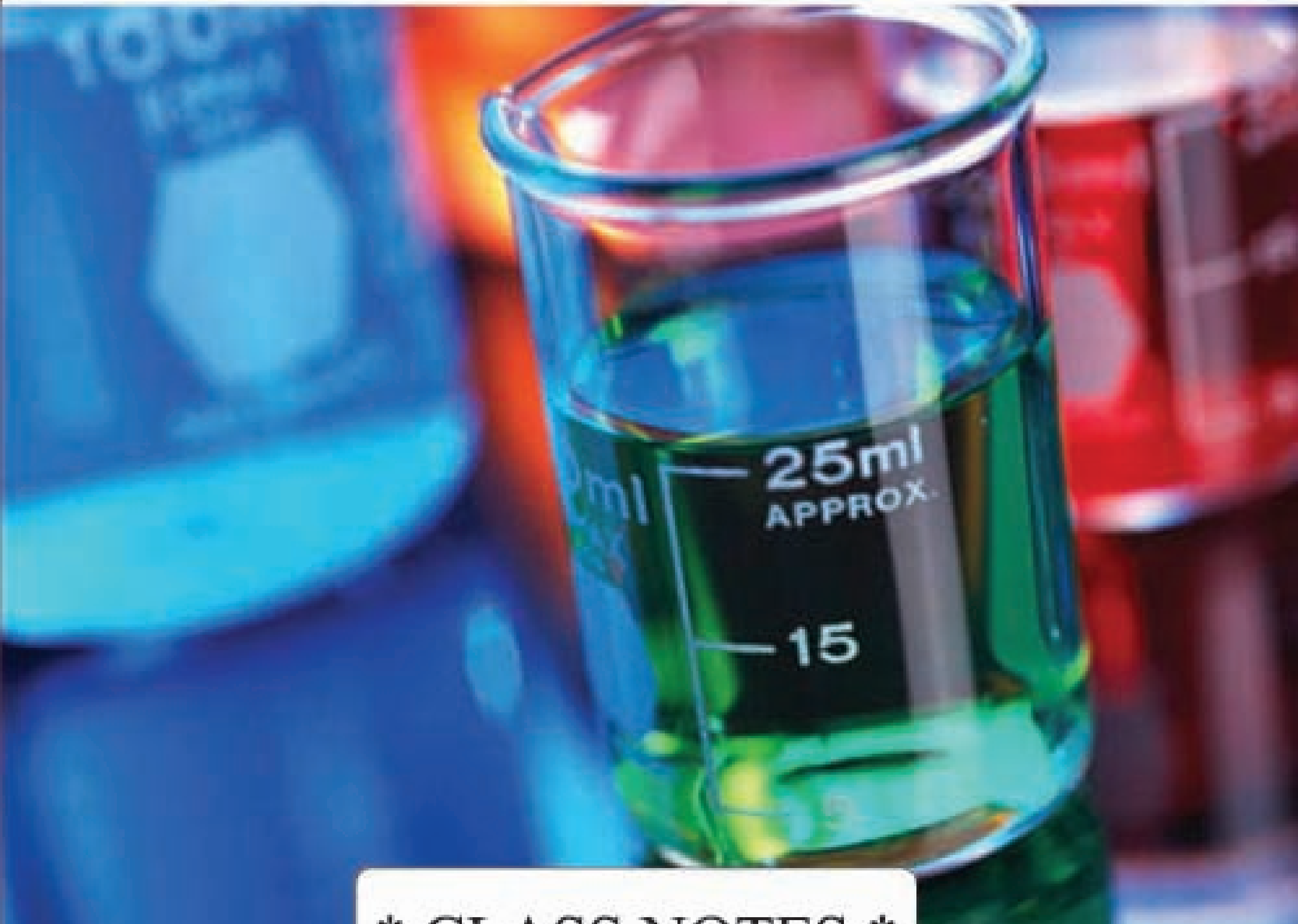
CHEMISTRY



* CLASS NOTES *
CHEMISTRY
Volume-1

CAREER POINT

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Preface

Being involved in preparing students for competitive examination since 1993, we realized that students require books which should be self-sufficient, relevant and in student's friendly language like class notes. The book should develop sound understanding of fundamentals and also enhance questions solving ability of students. The book in your hand has been prepared to achieve these objectives.

We have developed this book using class notes of top faculty members of Career Point who have been successfully preparing students for JEE and Pre-Medical for more than two decade. Structure of book is such that you will feel like you are virtually attending the class of a Teacher. We firmly believe that the book in this form will definitely help a genuine, hardworking student to achieve target.

We have tried our best to keep errors out of this book. Comment and criticism from readers will be highly appreciated and incorporated in the subsequent edition.

We wish to utilize the opportunity to place on record our special thanks to all faculty members of Career Point and Content Development Team for their efforts to make this wonderful book.

Career Point Ltd.

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1 Basic Concepts of Chemistry

1. INTRODUCTION

(a) A symbol mole (mol) is defined as the quantity of a given substance that contains as many particles as the number of atoms in exactly 12gram (gm) of carbon (atomic mass 12).

(b) The number of carbon atoms present in a 12gm of sample of carbon-12, is known as Avogadro's number (symbol. N_A). The value of N_A is 6.023×10^{23}. Thus, one mole of any given substance contain N_A number of molecules.

(c) Whenever there are no existence of actual molecules, e.g., ionic compounds, viz, NaCl, $CaCl_2$ etc., 1 mole is constituted by N_A number of formula units. For N_2, H_2, O_2 etc., individual molecules can exist, therefore, 1 mole of the given substances is equal to the sum of the weight of N_A number of molecules of given gases.

(d) Monoatomic gases like He, Ne etc., where individual molecule is made of one atom only, 1 mole of the given substances is the sum of the weight of N_A number of atoms.

(e) For metallic substances viz, Fe, Co, Ni, Na, K, etc; Individual molecule never exist. Therefore, 1 mole of those metals represented by the weight of N_A number of such metal atoms and thus, 1 mole of the given substances is equal to gram atomic weight of those metals.

(f) Mole concept is based on the application of the law of conservation of atoms, first proposed by Dalton as an extension of conservation of mass. It is utilited to solve the problems of stoichiometry, means the calculation of quantities of reactants and products involved in a chemical reaction.

(g) Before studying the mole concept, students are advised to clear their concept in symbol, valency and chemical formula.

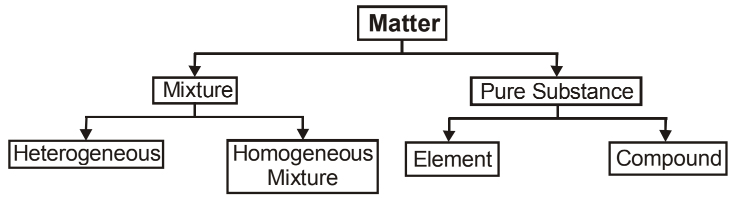
2. MATTER

Chemistry deals with structure, composition and properties of matter.

Matter may be defined as anything which has mass and occupies space. e.g. water, air, milk, salt, sand, oxygen etc.

Matter may be classified into three states namely solid, liquid and gas. At macroscopic level matter can be classified as mixture or pure substance as :

NOTES



2.1 Mixture :

Mixture contains two or more pure substances present in it in any ratio.

Ex. Air, solution of sugar in water, sugar and chalk, Tea etc.

2.2 Homogeneous mixture :

These are the mixtures having uniform composition throughout and in which individual components can not be distinguished from one another.

Ex. Sugar solution in water, air, C_2H_5OH (ethyl alcohol) in water etc.

2.3 Heterogeneous Mixture :

Mixtures whose composition is not uniform throughout and individual components can be distinguished from one another are heterogeneous mixtures.

Ex. $BaSO_4$ or AgCl in water, $CaCO_3$ (Solid) with CO_2 (gas), sand and sugar etc.

2.4 Element :

Purest substance which can not be separated into other elementary materials and which contains only one kind of atoms is known as element.

Ex. Hydrogen, oxygen, sodium, Helium etc.

Elements are made of some smallest particle which may be atom or molecule

Note : (i) Metals have atoms as smallest constituent unit.

(ii) Non-metals like He, Ne, Ar, Kr, Xe, Rn also have atoms as smallest constituent unit.

(iii) Carbon, Boron, silicon etc. are found to form gaint molecules but for calculation we again have atoms as smallest constituent units.

(iv) Remaining elements like hydrogen (H_2), oxygen (O_2), chlorine (Cl_2), phosphorous (P_4), sulphur (S_8) etc. are found to have molecules as smallest constituent unit.

2.5 Compound :

(I) When two or more elements combine in a definite proportion to give a new substance which has different properties from constituent elements then new substance formed is known as a compound.
Ex. Calcium carbonate($CaCO_3$), Sodium chloride (NaCl), Carbondioxide (CO_2) etc.

(II) In case of compound smallest constituent unit is molecule which is formed by the combination of atom of constituent elements.

(III) In ionic compounds molecule is not formed but in these compounds smallest constituent unit is ion.
Ex. NaCl (Na^+ and Cl^- ions), KCl (K^+ and Cl^- ion), $CaCO_3$ (Ca^{+2} ion and CO_3^{-2} ion) etc.

3. VALENCY AND FORMULA OF SOME RADICALS (CATIONS AND ANIONS)

Monovalent ions

CATIONS		ANIONS	
Ammonium	NH_4^+	Hydroxide	OH^-
Sodium	Na^+	Nitrate	NO_3^-
Potassium	K^+	Nitrite	NO_2^-
Rubidium	Rb^+	Permanganate	MnO_4^-
Cesium	Cs^+	Bisulphite	HSO_3^-
Silver	Ag^+	Bicarbonate	HCO_3^-
Copper (Cuprons)	Cu^+	Dihydrogen phosphate	$H_2PO_4^-$
Gold (Aurons)	Au^+	Chlorate	ClO_3^-
		Perchlorate	ClO_4^-
		Hypochlorite	ClO^-
		Iodate	IO_3^-
		Periodate	IO_4^-
		Meta aluminate	AlO_2^-
		Meta borate	BO_2^-
		Cyanide	CN^-
		Isocyanide	NC^-
		Cyanate	CNO^-
		Isocyanate	NCO^-

Divalent ions

CATIONS		ANIONS	
Magnesium	Mg^{2+}	Carbonate	CO_3^{-2}
Calcium	Ca^{2+}	Sulphate	SO_4^{-2}
Stroncium	Sr^{2+}	Sulphite	SO_3^{-2}
Barium	Ba^{2+}	Sulphide	S^{-2}
Zinc	Zn^{2+}	Thiosulphate	$S_2O_3^{-2}$
Cadmium	Cd^{2+}	Tetrathionate	$S_4O_6^{-2}$
Nickel	Ni^{2+}	Oxalate	$C_2O_4^{-2}$
Copper (Cupric)	Cu^{2+}	Silicate	SiO_3^{-2}
Mercury (Mercuric)	Hg^{2+}	Hydrogen phosphate	HPO_4^{-2}
Lead (Plumbus)	Pb^{2+}	Manganate	MnO_4^{-2}
Tin (Stannous)	Sn^{2+}	Chromate	CrO_4^{-2}
Iron (Ferrous)	Fe^{2+}	Dichromate	$Cr_2O_7^{-2}$
		Zincate	ZnO_2^{-2}
		Stannate	SnO_3^{-2}

Trivalent ions

CATIONS		ANIONS	
Iron (Ferric)	Fe^{3+}	Phosphate	PO_4^{-3}
Aluminium	Al^{3+}	Borate (orthoborate)	BO_3^{-3}
Chromium	Cr^{3+}	Aluminate	AlO_3^{-3}
		Arsenate	AsO_4^{-3}
Gold (Auric)	Au^{3+}	Arsenite	AsO_3^{-3}
		Nitride	N^{-3}
		Phosphide	P^{-3}

4. FORMULA OF SIMPLE COMPOUNDS

S.NO.	Name of Compound	Symbols with Valency	Formula
1.	Calcium chloride	$Ca^{2+}\ Cl^{1-}$	$CaCl_2$
2.	Magnesium sulphate	$Mg^{2+}\ SO_4^{2-}$	$Mg_2(SO_4)_2$ ~ $MgSO_4$ (Simple Ratio)
3.	Stannic Sulphide	Sı	SnS_2
4.	Potassium perchlorate	$K^{1+}\ ClO_4^{1-}$	$KClO_4$
5.	Sodium Zincate	$Na^{1+}\ ZnO_2^{2-}$	Na_2ZnO_2
6.	Magnesium bicarbonate	$Mg^{2+}\ HCO_3^{1-}$	$Mg(HCO_3)_2$
7.	Sodium carbonate	$Na^{1+}\ CO_3^{2-}$	Na_2CO_3
8.	Ammonium Oxalate	$NH_4^{1+}\ CO_3^{2-}$	$(NH_4)_2C_2O_4$
9.	Sodium thiosulphate	$Na^{1+}\ S_2O_3^{2-}$	$Na_2S_2O_3$
10.	Potassium permanganate	$K^{1+}\ MnO_4^{1}$	$KMnO_4$
11.	Sodium Iodate	$Na^{1+}\ IO_3^{1-}$	$NaIO_3$
12.	Sodium periodate	$Na^{1+}\ IO_4^{1-}$	$NaIO_4$

5. LAWS OF CHEMICAL COMBINATIONS

There are five laws which were first proposed in the 18th century and early of 19th century to be followed by the chemical reactions. At that time these laws was universally accepted by the scientists, because each of these laws was supported by the Dalton's atomic theory. After the discovery of nuclear reactions, isotopes and complex biochemical and organic molecules, few these laws has been found to be not true. But till date, for simple chemical reaction, these laws are used. The five laws are :

(a) The law of conservation of mass (Levoisier 1774).

(b) The law of constant proportion or definte proportions (proust 1799)

(c) The law of multiple proportions (Dalton 1803)

(d) Law of equivalent or reciprocal or combining proportions or weights. (Richter 1872).

(e) Gay Lussac's law of combining volume.

(a) Law of conservation of mass :

"The law states that matter can neither be created nor destroyed, and it is therefore, often called the law of indestructibility of matter." As per this law in all chemical and physical changes, the total mass of the substances involves remain unchanged.

★ **POAC (Principle of atomic conservation)**

This laws has been further extended by Dalton as atoms cannot be created or destroyed. In any chemical reaction, the total number of atoms of an element present in the reactants side must be equal to the total number of atoms of the same element on the products side. This law will be applied in balancing the chemical equation.

(b) The law of constant proportion or definite proportion :

"This law states that, a given compound, wherever obtained and however prepared contain its component elements in a fixed ratio by weight." As for example, water obtained from any source contains hydrogen and oxygen combined in a ratio of 2 : 16 = 1 : 8 by weight.

(c) The law of multiple proportion :

It states that when two elements combine to form more than one compound, the several weights of first element that combine with the fixed weight of second one are in the ratio of their small whole numbers. Nitrogen and oxygen combine together to form different oxides. The results of this experimentation is as follows :

Oxides of nitrogen	Weight of nitrogen	Weight of oxygen	Fixed weight of nitrogen	Weight of oxygen which combined with fixed weight of nitrogen
N_2O	28	16	28	16
NO	14	16	28	32
N_2O_3	28	48	28	48
N_2O_4	28	64	28	64
N_2O_5	28	80	28	80

The ratio of the weights of oxygen that combine with 28 gm of nitrogen in the given oxides

= 16 : 32 : 48 : 64 : 80 = 1 : 2 : 3 : 4 : 5

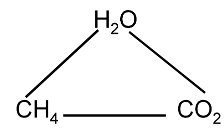
(d) Law of Equivalent, Reciprocal and combining proportions or weights:

This law, which is sometimes known as the law of equivalent weights, deals with the case in which two elements combine with the third element. It states that :

"When two or more elements (A, B, C, etc.) combine separately with another element (D), then the respective weights of these elements (A, B, C) which combine with a fixed weight of another (D) are in the same proportion as those in which they combine to form compounds within themselves (AB, BC, etc) or in simple multiples of these proportions."

This law is explained as follows -

consider two elements hydrogen and oxygen which combine with a third element ; say carbon, to form CH_4 and CO_2 respectively.



In CH_4, weight of hydrogen which combine with 12 gm of carbon divided by weight of oxygen which combine with 12 gm of carbon in CO_2 is equal to $\frac{4}{32} = \frac{1}{8}$

In H_2O, $\frac{\text{weight of H}}{\text{weight of O}} = \frac{2}{16} = \frac{1}{8}$

Thus the two ratios are simply related

(e) Gay Lussac's law of combining volume (Gay Lussac 1808)

This law is applicable for gaseous reactants only. This law states, when gases react, the volume of the reacting gases are in the ratio of small whole numbers, provided the volumes are measured at the same temperature and pressure. If the reactants and products both are gases, then at the same temperature and pressure, the volume of reactant and product gases are also in the ratio of small whole numbers.

As for example, at the same temperature, 1 volume of nitrogen and 3 volume of hydrogen combine to form 2 volume of ammonia. The ratio of the reactants and product gases is 1 : 3 : 2, a ratio of small whole numbers. This concept will be further applied in solving the problem of Eudiometry.

➢ *Examples Based On* **: LAW OF CHEMICAL COMBINATIONS**

Ex.1 In which of the following reaction Gay Lussac law is applicable -

(A) $AgNO_3(s) + NaCl(solution) \rightarrow AgCl(s) + NaNO_3(solution)$

(B) $C(s) + 2\ Cl_2(g) \rightarrow CCl_4(\ell)$

(C) $NH_3(g) + HCl(\ell) \rightarrow NH_4Cl(s)$

(D) $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

Sol. **(D)** In option (D) only two reactants are gaseous, therefore Gay Lussac's law is applicable

Ex.2 1.375 gm of pure cupric oxide was reduced by heating in a current of pure dry hydrogen and the mass of copper that remained 1.0980 gms. In another experiment, 1.179 g of pure copper was dissolved in pure HNO_3 and the resulting copper nitrate converted into cupric oxide by ignition. The mass of copper oxide formed was 1.476 g. Show that the results illustrate to law of constant composition with in the limits of experimental error.

Sol. 1.375 gm of pure cupric oxide gave 1.098 gm of Cu and hence,

percentage of Cu in the oxide = $\frac{1.098}{1.375} \times 100 = 79.85\%$ in another experiment, 1.179 gm of pure copper gave 1.476 gm of the oxide and hence.

percentage of Cu is the oxide = $\frac{1.179}{1.476} \times 100 = 79.87\%$

Since, both the oxides have almost the same percentage of Cu and hence, of oxygen, the result is obeying the law of constant composition.

6. BRIEF DISCUSSION ABOUT SOME OF THE TERMS USED IN MOLE CONCEPT

6.1 Atomic weight and atomic mass unit (amu) :

(a) The atomic weight (or atomic mass) of an element may be defined as the average relative weight (or mass) of an atom of the element with respect to the $\frac{1}{12}$ th mass of an atom of carbon (mass number 12)

(b) Thus,

$$\text{atomic weight} = \frac{\text{Weight of an atom of the element}}{\text{Weight of an atom of C (mass no. 12)}} \times 12$$

(c) The multiplying factor 12 is used to avoid atomic weights less than unity.

(d) If we express atomic weight in grams, it becomes gram atomic weight (symbol gm-atom). For example, atomic weight of oxygen = 16, therefore, 1 gm-atom oxygen = 16 gm of oxygen, similarly, atomic weight of Ag = 108, therefore, 540 gm Ag = $\frac{540}{108}$ = 4 gm-atom of Ag.

(e) Thus the number of gm-atom of an element = $\frac{W}{A}$, where w is the given mass of the element of atomic weight A.

(f) 1 gm-atom of any element contain N_A number of atoms.

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(g) The atomic weight of H = 1.008, therefore, the weight of single H-atom $= \frac{1.008}{N_A}$ gm $= \frac{1.008}{6.023 \times 10^{23}} = 1.673 \times 10^{-24}$gm.

Similarly, $\frac{1}{12}$ th of the mass of a carbon atom (mass number 12)

$= \frac{1}{12} \times \frac{12}{6.023 \times 10^{23}} = \frac{1}{6.023 \times 10^{23}} = 1.667 \times 10^{-24}$ gm.

The atomic mass unit (amu or u) is defined as the $\frac{1}{12}$ th of the mass of single carbon atom of mass number 12.
Thus, 1 amu or u = 1.667×10^{-24} gm = 1.667×10^{-27} kg.

➢ *Examples Based On* **: ATOMIC WEIGHT**

Ex.3 Calculate the number of atoms present in 4.8 gm of oxygen gas.

Sol. The weight of oxygen. i.e, w = 4.8 gm.
The atomic mass of oxygen. i.e, A = 16 gm.

$\therefore$ number of gm-atom of oxygen $= \frac{w}{A} = \frac{4.8}{16}$

$\therefore$ The number of oxygen atoms present in the given gas $= \frac{4.8}{16} \times N_A$

Ex.4 The radius of hydrogen atom is 0.53 Å. Assuming the hydrogen atom as spherical, calculate the atomic weight of Hydrogen. (Given the density of hydrogen atom = 2.675 gm/ml)

Sol. The volume of each hydrogen atom $= \frac{4}{3}\pi r^3$

Now r = 0.53 Å = 0.53×10^{-8} cm.

$\therefore \frac{4}{3}\pi r^3 = \frac{4}{3} \times \pi \times (0.53)^3 \times 10^{-24}$ cm^3

$\therefore$ mass of each hydrogen atom $= \frac{4}{3} \times \pi \times (0.53)^3 \times 10^{-24} \times 2.675$ gm

$\therefore$ atomic weight of hydrogen

$= N_A \times$ mass of one hydrogen atom

$= 6.023 \times 10^{23} \times \frac{4}{3} \times \pi \times (0.53)^3 \times 10^{-24} \times 2.675$

$= 10.04 \times 10^{-1} = 1.004$ **Ans.**

6.2 Molecular Weight and Formula Weight :

(a) Molecular weight is defined as the weight of a molecule of a substance compared to the $\frac{1}{12}$ th of the mass of a carbon atom (mass number = 12). For compounds, molecular weight is the sum of the atomic weights of all atoms present in the molecule.

(b) In ionic compounds, as for example, NaCl, $CaCl_2$, H_2SO_4, etc. there are no existence of molecules. The individual units are stabilized by the electrostatic interactions of the ion pairs present in the ionic compounds. so, word, molecular weight is not applicable.

(c) Instead of "molecular weight" we use a new term known as "formula weight". "Formula weight" is defined as the total weights of atoms present in the formula of the compound.

(d) Thus, formula weight of $Fe_2 (SO_4)_3$

$= 56 \times 2 + 32 \times 3 + 16 \times 12 = 400$

(At. wts of Fe and S are 56 and 32 respectively).

6.3 The average atomic mass and average molecular mass :

(a) After the discovery of isotopes, it has been concluded that same element can be present in the nature with different atomic masses.

(b) Let us consider, an element x, is available in the earth as isotopes of ${}_{n}X^{a_1}$, ${}_{n}X^{a_2}$, ..., ${}_{n}X^{a_n}$ the percentage abundance of the given isotopes in earth are x_1 , x_2, ..., x_n respectively.

$\therefore$ the average atomic mass of X $(A_x) = \dfrac{a_1x_1 + a_2x_2 + ...+ a_nx_n}{100}$

Calculation of average molar mass :

(c) Let us consider, in a container,

n_1 moles of substance x_1 (mol. wt M_1) present

n_2 moles of substance x_2 (mol. wt M_2) present

....

n_n moles of substance x_n (mol. wt M_n) present

hence, the total number of moles of substance present in the container $= n_1 + n_2 + ... + n_n$

Total mass of the substance present in the container $= n_1M_1 + n_2M_2 + ... + n_nM_n$

$\therefore$ the average molar mass of the substance present in the container $= \dfrac{n_1M_1 + n_2M_2 + ...n_nM_n}{n_1 + n_2 + ...n_n}$

➢ *Examples Based On* : **MOLECULAR WEIGHT & FORMULA WEIGHT**

Ex.5 How many grams of nitrogen gas can be maximum obtained from 720 kg of urea ?

Sol. The molecular formula of urea is $CO(NH_2)_2$.

Its molecular wt = 12 + 16 + 28 + 4 = 60

From 60 kg of urea 28 kg of nitrogen can be obtained

From 720 kg of urea $\dfrac{28}{60} \times 720 = 336$ kg

Ex.6 Naturally occurring chlorine is 75.53% Cl^{35} which has an atomic mass of 34.969 amu and 24.47% Cl^{37} which has a mass of 36.966 amu. Calculate the average atomic mass of chlorine -

(A) 35.5 amu (B) 36.5 amu

(C) 71 amu (D) 72 amu

Sol. **(A)** Average atomic mass =

% of I isotope × Its atoms mass + % II isotope × its atomic mass/100 $= \dfrac{75.53 \times 34.969 + 24.47 \times 36.96}{100} = 35.5$ amu.

Ex.7 The molar composition of polluted air is as follows :

Gas	At. wt.	Mole Percentage Composition
Oxygen	16	16%
Nitrogen	14	80%
carbon dioxide	–	03%
sulphurdioxide	–	01%

What is the average molecular weight of the given polluted air ? (Given, atomic weights of C and S are 12 and 32 respectively).

Sol. $M_{avg} = \dfrac{\sum_{j=1}^{j=n} n_j M_j}{\sum_{j=1}^{j=n} n_j}$, Here $\sum_{j=1}^{j=n} n_j = 100$

$$\therefore \quad M_{avg} = \frac{16 \times 32 + 80 \times 28 + 44 \times 3 + 64 \times 1}{100}$$

$$= \frac{512 + 2240 + 132 + 64}{100} = \frac{2948}{100} = 29.48$$ **Ans.**

7. UNITS OF PRESSURE, VOLUME AND TEMPERATURE; CONVERSION OF VOLUME OF GASES INTO MASS

(a) In the chemical calculation, sometimes the conversion of volume of gas into its weight is necessary.

(b) To do so, we are to apply the combined gas law. For example, at absolute temperature T (273 + t ºC) and pressure P, the volume of a given mass of gas (molecular weight M) is V, now as per the combined gas law we can write,
PV = nRT (where n indicates number of moles of gas and R is universal gas constant)

As we know, $n = \dfrac{w}{M}$, where w is the mass of gaseous substance).

$$\therefore PV = \frac{w}{M} RT \quad \text{or} \quad w = \frac{PVM}{RT}$$

(c) Therefore if we know, pressure, volume, temperature and molecular weight of gas, we can calculate its mass.

(d) Be careful while putting the value of R. R = 0.082 lit atm/K/mole when units of P and V are atmosphere and Liter respectively. It means, if the pressure of the given gas in other unit, you have to convert into atmosphere (symbol atm). Similarly, if the volume of the gas is given in any other unit, you have to convert it into litre. Similarly, you have to convert the given temperature into absolute temperature.

(e) Density of gas may be calculated as

$$d = \frac{w}{V} = \frac{PM}{RT}$$

(f) Here we are discussing the relation between the various units of pressure, volume and temperature.

Pressure units

The units of pressure of gas are divided into two types

(i) Absolute units (ii) Relative units

(i) Absolute units :

In MKS system : N/m^2 or Pascal (Pa)
N stands for Newton
In CGS system : gm/cm^2
In FPS system : lb/ft^2 or lb per square inch ($lb/inch^2$)

(ii) Relative units :

Bar, atmosphere (symbol atm), torr and cm of Hg.

Relation :

1 atm = 1.01325×10^5 Pa = 1.01325 bar (1bar = 10^5 Pa)
= 14.7 lb/in^2
= 760 torr = 760 mm of Hg (1 torr = 1 mm. of Hg)
= 76 cm of Hg.

Volume units :

1 dm^3 = 1 litre (symbol lit) = 1000 ml = 1000 cm^3
1 m^3 = 10^3 litre = 10^6 ml
1 mm^3 = 10^{-3} cm^3 = 10^{-9} m^3
1 U.S. gallon = 8.34 lb of water at room temperature
1 lb of water = 453.6 gm
1 U.S. gallon = 8.34 × 453.6 ml = 3783 ml = 3.783 lit

Temperature units :

Absolute temperature T (in k) = 273 + t°C where, t is the temperature in centigrade scale.
STP (Standard temp. and pressure) which considerd 273K and 1atm
NTP (Normal temp. and pressure) which considerd 273K and 1atm

➢ *Examples Based On* : **PRESSURE, VOLUME AND TEMPERATURE RELATIONSHIP**

Ex.8 What is the density of H_2S gas at 27°C and 2 atmosphere ?
(A) 2.77 gm/lit (B) 2.25 gm/lit
(C) 2 gm/lit (D) 3.4 gm/lit

Sol. **(A)** Here P = 2 atm, M = 2 × 1 + 32 = 34, T = 273 + 27 = 300 K

$$d = \frac{PM}{RT} = \frac{2 \times 34}{0.082 \times 300} = 2.764 \text{ gm/litre}$$

Ex.9 Density of dry air containing only N_2 and O_2 is 1.146 gm/lit at 740 mm and 300 K. what is % composition of N_2 by weight in the air.
(A) 78% (B) 82%
(C) 73.74% (D) 72.42%

Sol. **(C)** First calculate the av. mol. wt. of the mixture as

$$d = \frac{PM}{RT} \Rightarrow M = \frac{dRT}{P} = \frac{1.146 \times 0.082 \times 300}{\left(\frac{740}{760}\right)} = 28.95$$

Now, let 100 gm of dry air contains x gm N_2 and hence (100 – x) gm O_2. Applying mole conservation :

mole of air = mole of N_2 + mole of O_2

or, $\frac{100}{28.95} = \frac{x}{28} + \frac{100 - x}{32} \Rightarrow x = 73.74$

Ex.10 A 1.225g mass of a volatile liquid is vapourized, giving 400ml of vapour when measured over water at 30ºC and 770 mm. The vapour pressure of water at 30ºC is 32 mm. What is the molecular weight of the substance ?

Sol. w = 1.225 gm, V = 400 ml = 0.4 litre ;

T = 273 + 30 = 303 K

P = (770 – 32) = 738 mm = 0.971 atm

(from Dalton's law of partial pressure, pressure of dry gas = total pressure – vapour pressure of water)

Now, PV = n R T or $PV = \frac{w}{M}RT$

$\therefore M = \frac{wRT}{PV} = \frac{1.225 \times 0.082 \times 303}{0.4 \times 0.971} = 78.36$

8. CONCEPT OF MOLES AND CONCLUSIONS OF AVOGADRO'S THEORY

(a) When we express the molecular weight or formula weight of a given substance in grams, the amount is known to be gm molecular weight or simply mole (symbol mol). Thus, the molecular weight of CO_2 = 44, therefore, 2.2 gm $CO_2 = \frac{2.2}{44}$ = 0.05 mol of CO_2.

(b) We can also express the molecular weight of a given substance in kg or pound (lb), for those cases, we need to introduce two new terms, kg-mole and lb-mole.

(c) The kg-mole and lb-mole can be easily corelated with the mole.

1 kg-mole = (conversion factor) × mole

1 lb-mole = (conversion factor) × mole

(d) 1 kg-mole of CO_2 = 44 kg = 44 × 10^3 gm = 10^3 mole

1 lb-mole of CO_2 = 44 lb = 44 × 453.59 gm = 453.59 mole

[1 lb = 453.59 gm]

Thus, 1 kg-mole = 10^3 mole and 1 lb mole = 453.59 mole

Conclusions of Avogadro's theory

(i) One mole of all substances contain Avogadro's number of particles i.e., 6.023 × 10^{23}

(ii) The volume of 1 mole of gas at NTP (normal temperature and pressure) or STP (Standard temperature and pressure) is 22.4 litre. NTP or STP refers for pressure = 1 atm and T = 0ºC or 273 K.

(iii) Gram molecular weight (mole) may also be regarded as the weight of 22.4 litre of gas at NTP in gm.

NOTES

(iv) Molecular weight = 2 × Vapour density*

* Vapour density is also known as relative density.

$$\text{Vapour density} = \frac{\begin{array}{c}\text{Weight of definite volume ture and pressure}\\ \text{of a gas at definite temperature}\end{array}}{\begin{array}{c}\text{Weight of same volume of hydrogen gas}\\ \text{at the same temperature and pressure}\end{array}}$$

(v) Avogadro's number, N_A = the number of molecules present in 1 mole of a compound = 6.023×10^{23}.

➢ *Examples Based On* : **CONCEPT OF MOLES & AVOGADROS' THEORY**

Ex.11 Assuming the density of water to be 1 gm/ml, the volume occupied by one molecule of water is -

(A) 2.98×10^{-23} ml (B) 2.98×10^{23} ml

(C) 6.02×10^{23} ml (D) 6.02×10^{-23} ml

Sol. **(A)** mass of 1 mole of H_2O = 18gm

volume of 1 mol of H_2O = 18 ml.

∴ volume of one molecule of water

$$= \frac{18}{N_A} = \frac{18}{6.023 \times 10^{23}} = 2.98 \times 10^{-23} \text{ ml}$$

Ex.12 The shape of Tobacco Mosaic virus (TMV) is cylindrical, having length and diameter are 3000 Å and 170 Å respectively. The density of the virus particle is 0.08 gm/cm³. What is the molecular weight of Tobacco Mosaic virus?

Sol. The volume of single TMV virus = $\frac{1}{4}\pi d^2 \ell$

where, d and ℓ stands for diameter and length respectively.

Here,

d = 170 Å

= 170×10^{-8} cm

= 1.70×10^{-6} cm. and

ℓ = 3000 Å

= $3 \times 10^3 \times 10^{-8}$ cm

= 3×10^{-5} cm.

Hence, volume of single virus particle

$$= \frac{1}{4} \times 3.14 \times (1.70)^2 \times 3 \times 10^{-17} \text{ cm}^3 = 6.806 \times 10^{-17} \text{ cm}^3$$

Mass of single virus particle,

= $6.806 \times 10^{-17} \times 0.08$ gm

= $6.806 \times 8 \times 10^{-19}$ gm.

To calculate the molecular weight of TMV virus, it has been considered to be a single unit, hence,

the molecular weight = N_A × mass of one unit.

∴ The molecular weight of TMV virus

= $6.806 \times 8 \times 10^{-19} \times 6.023 \times 10^{23}$

= 327.94×10^4

= 3.2794×10^6 **Ans.**

Ex.13 P and Q are two elements which forms P_2Q_3 and PQ_2. If 0.15 mole of P_2Q_3 weighs 15.9 gm and 0.15 mole of PQ_2 weighs 9.3 gm, what are the atomic weights of P and Q ?

Sol. Let the atomic weight of P and Q are a and b respectively.

$\therefore$ the molecular weight of P_2Q_3 = $2a + 3b = \frac{15.9}{0.15} = 106$ ...(i)

and molecular weight of PQ_2 = $a + 2b = \frac{9.3}{0.15} = 62$ ...(ii)

By solving equation (i) and (ii), we get,

a = 26 and b = 18

$\therefore$ the atomic weights of P and Q are 26 and 18 respectively. **Ans.**

9. EMPIRICAL FORMULA AND MOLECULAR FORMULA

There are two major kinds of chemical formula, viz,

(i) The simplest or empirical formula

(ii) The molecular formula

(i) Empirical formula (simplest formula)

The empirical formula of a compound reflects the simple ratio of atoms present in the formula units of the compound.

(ii) Molecular formula

The molecular formula is the actual number of atoms of the constituent elements that comprise a molecule of the substance.

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

Here n = 1, 2, 3.......

Determination of empirical and molecular formula

Find out the molecular formula of glucose if it contains, C = 40%, H = 6.66% and O =53.33%, vapour density of glucose is 90.

Empirical formula of glucose = CH_2O

Now let its molecular formula = $(CH_2O)_n$

So molecular weight = 30n

$\Rightarrow$ 30n = 180

$\Rightarrow$ n = 6 so molecular formula of glucose = $(CH_2O)_6 = C_6H_{12}O_6$

➢ *Examples Based On* : **EMPIRICAL FORMULA AND MOLECULAR FORMULA**

Ex.14 Red lead is composed of 90.65% lead and 9.35% oxygen. What is the empirical formula of red lead ? (Given at. wt of Pb = 207).

Sol. To solve the given problem we will also follow the earlier procedure. At first we will prepare the table as.

weight percentage	Pb	O
(as given)	90.65	9.35

Element	C	H	O
(as given) % of element	40	6.667	53.333
Atomic ratios of the elements	$\frac{40}{12} = 3.333$	$\frac{6.667}{1} = 6.667$	$\frac{53.333}{16} = 3.333$
Simple atomic ratio	1	2	1

Atomic ratio of elements $\frac{90.65}{207}$ = 0.4379

$\frac{9.35}{16}$ = 0.5843

since atoms cannot be divided further, therefore the atomic ratios of both Pb and O present in one molecule of red lead cannot be fractional. To proceed, we will first eliminate the fractional part of an element by dividing the each atomic ratio by the lowest atomic ratio of the constituent element. Thus,

Pb	O
$\frac{0.4379}{0.4379} = 1$	$\frac{0.5843}{0.4379} = 1.33$

Now, atomic ratio of O cannot be fractional one. To eliminate the fractional part, we will search a simple fraction of decimal value equal to 1.33, of which numerator and denominator both are whole numbers. It has been found that $\frac{4}{3} \simeq 1.33$, so we can prepare the final table as.

	Pb	O
Atomic ratios of the elements present in red lead	1	4/3
simplest atomic ratio	3	4

$\therefore$ the empirical formula of red lead is Pb_3O_4.

Ex.15 A student heated 1.222 gm of hydrated $BaCl_2$ to dryness. 1.042 gm of the anhydrous salt remained. What is the formula of hydrate ? (Given at. wt. of Ba = 137)

Sol. Let x mol of water is present per mol of hydrated $BaCl_2$.

$\therefore$ $BaCl_2 + x\,H_2O \longrightarrow BaCl_2$

1.222 gm. 1.042 gm.

Hence wt. of water of crystallisation = 1.222 – 1.042 = 0.180 gm.

the mol. wt. of $BaCl_2$ = 137 + 2 × 35.5 = 208.

Per 1.042 gm of $BaCl_2$ water associated = 0.180 gm.

$\therefore$ Per 208 gm. of $BaCl_2$ water associated

$= \frac{0.180}{1.042} \times 208$ gm = 35.93 gm.

= 1.996 mole

$\simeq$ 2 mole.

since, in 1 mole of hydrated barium chloride, 1 mole of $BaCl_2$ is present.

$\therefore$ the formula of hydrated barium chlorides $BaCl_2 . 2H_2O$ **Ans.**

NOTES

10. GRAVIMETRIC ANALYSIS AND VOLUMETRIC ANALYSIS

The problems of stoichiometry have been classified broadly into two types :

(i) Gravimetric Analysis

(ii) Volumetric Analysis

(i) Gravimetric Analysis :

In gravimetric analysis we relate the weights of two substances or a weights of a substance with a volume of gas or volumes of two or more gases.

(ii) Volumetric Analysis :

The volumetric analysis is mainly based on the laws of equivalence which will be discussed in details in the equivalent concept, briefly, volumetric chemical analysis consists in experimentally finding the volume of standard solution (known strength) which will react completely with a given quantity of substance or a measured volume of a solution of unknown concentration. The completion of reaction is indicated by a substance, known as indicator and the entire process is known as titration. Based on the concept of law of equivalence, the strength of unknown solution is determined. From the calculation of strength, we easily calculate the amount of substance present in the given unknown solution.

10.1 Details about Gravimetric analysis :

Problems of Gravimetric analysis are of three types.

(i) Problems involving mass by mass relationship

(ii) Problems involving mass by volume relationship

(iii) Problems involving volume by volume relationship

(i) Problems involving mass by mass relationship :

To solve this types of problem you are to proceed as follows.

Step 1 : Find out the chemical reaction involved in the problem.

Step 2 : Write the chemical reaction as an equation and balance it by any suitable method.

Step 3 : Write the number of moles below the formula of the reactants and products. Also write the formula weights of reactants and products, below their respective formula.

Step 4 : Apply the unitary method to calculate the unknown factor (s). Let us start with a problem; it is, A mixture of NaCl and KCl weighed 5.4892 gm. The sample was dissolved in water and treated with an excess of silver nitrate solution. The resulting AgCl weighed 12.7052 gm. What is the percentage of NaCl in the mixture ? (Given atomic weight of Na, K, Ag, Cl are 23, 39, 108, 35.5 respectively)

The reactions involved are,

$NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$

58.5 143.5

$KCl + AgNO_3 \rightarrow AgCl + KNO_3$

74.5 143.5

Let us consider in the mixture x gm of NaCl present

$\therefore$ KCl present = (5.4892 – x) gm.

By, unitary methods, from

58.5 gm of NaCl, AgCl produced = 143.5 gm

1 gm of NaCl, AgCl produced = $\frac{143.5}{58.5}$ gm

x gm of NaCl, AgCl produced = $\frac{143.5\,x}{58.5}$ gm

similarly, from,

74.5 gm of KCl,

AgCl produced = $\frac{143.5}{74.5}$ (5.4892 – x) gm

$$\therefore \frac{143.5\,x}{58.5} + \frac{143.5}{74.5}(5.4892 - x) = 12.7052$$

$$\text{or } \frac{x}{58.5} + \frac{(5.4892 - x)}{74.5} = \frac{12.7052}{143.5} = 0.08854$$

$$\text{or } \frac{x}{58.5} - \frac{x}{74.5} = 0.08854 - \frac{5.4892}{74.5}$$

$$= 0.08854 - 0.07368 = 0.01486$$

or 16 x = 0.01486 × 58.5 × 74.5

or x = 4.0477

$\therefore$ % NaCl in the mixture = $\frac{4.0477}{5.4892} \times 100 = 73.74$ **Ans.**

Let us discuss another problem to clear your doubt further – it is, 25.4 gm of I_2 and 14.2 gm of Cl_2 are made to react completely to yield a mixture of ICl and ICl_3. Calculate moles of ICl and ICl_3 formed.

Step 1 : We will write the reaction in terms of chemical equation as.

$$I_2 + Cl_2 \longrightarrow ICl + ICl_3.$$

Step 2 : We will balance the chemical equation by any conveniable method as.

$$x\,I_2 + y\,Cl_2 \longrightarrow a\,ICl + b\,ICl_3.$$

by I balance by Cl balance

2x = a + b 2y = a + 3b.

2y – 2x = 2b.

or y – x = b. | a = 3x – y

if y = 2, x = 1, the a = 3 × 1 – 2 = 1.

b = 1

$\therefore$ $I_2 + 2Cl_2 \rightarrow ICl + ICl_3$.

This is the balanced equation.

Step 3 : At wt of I = 127, At wt of Cl = 35.5

$$\therefore 25.4 \text{ gm } I_2 = \frac{25.4}{127 \times 2} = 0.1 \text{ mole of } I_2.$$

$$14.2 \text{ gm } Cl_2 = \frac{14.2}{35.5 \times 2} = 0.2 \text{ mole } Cl_2.$$

NOTES

Step 4 : 1 mole I_2 reacts with 2 moles of Cl_2 as per equation

$\therefore$ 0.1 mole I_2 reacts with 0.2 moles of Cl_2

1 mole I_2 yields 1 mole of ICl & 1 mole of ICl_3

$\therefore$ 0.1 mole of I_2 yields 0.1 mole of ICl & 0.1 mole of ICl_3

$\therefore$ Moles of ICl formed = 0.1

$\therefore$ Moles of ICl_3 formed = 0.1 **Ans.**

(ii) Problems involving mass by volume relationship :

To solve the problems involving mass by volume relationship we are to proceed according to the following instructions :

Step 1 : Find out the chemical reaction involved in the problem and also find out the physical state of the given reactants and products.

Step 2 : Write the chemical reaction as an equation and balance it by any suitable method.

Step 3 : Gases are usually expressed in terms of volumes. Depending on your need, (i) convert the volume of the gas at NTP by applying the gas equation. or (ii) convert the volume of the gas into its wt. by applying the equation $PV = \frac{W}{M} RT$

Step 4 : Calculate the unknown factor by unitary method.

Let us solve one problem to clear your doubts. It is, a mixture of Na_2CO_3 and $NaHCO_3$ had a mass of 22 gm. Treatment with excess HCl liberates 6.00 lit CO_2 at 25ºC and 0.947 atm pressure. Determine the percent Na_2CO_3 in the mixture.

Step 1 : Among the reactants and products only CO_2 will be gaseous at the given condition. The Na_2CO_3 and $NaHCO_3$ both will react will HCl and yield CO_2 (g).

Step 2 : The balanced chemical reactions are

$Na_2CO_3 + 2\,HCl \rightarrow 2\,NaCl + H_2O + CO_2$

106 gm 44 gm

$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$

84 gm 44 gm

$\therefore \quad PV = \frac{W}{M} RT \text{ or } W = \frac{PVM}{RT} = \frac{0.947 \times 6 \times 44}{0.082 \times 298} = 10.23$

Let in the mixture x gm. of Na_2CO_3 present,

$\therefore$ from x gm. Na_2CO_3, CO_2 produced = $\frac{44x}{106}$ gm.

In the mixture the amount of $NaHCO_3$ present = (22 – x) gm.

From (22 – x) gm of $NaHCO_3$, CO_2 produced = $\frac{44}{84}$ (22 – x) gm

$\therefore \quad \frac{44x}{106} + \frac{44}{84}(22 - x) = 10.231$

on, solving x = 11.9 gm.

$\therefore$ % Na_2CO_3 in the 22 gm of mixture = $\frac{11.9}{22} \times 100 = 54.09$ **Ans.**

(iii) Problems involving volume by volume relationship :

(a) The type relationship are used correlate the volume of gaseous reactants with other gaseous reactants or gaseous products.

(b) Stoichiometric equation for the combustion of any hydrocarbon is, $C_xH_y + (x + y/4)\ O_2 \rightarrow xCO_2 + y/2\ H_2O$. Consider at temperature (T) and pressure (P) we have taken v_H volume of C_xH_y. Therefore for complete combustion of v_H volume of C_xH_y,
O_2 required = $v_H\ (x + y/4)$ vol. [At pressure P and temp. T]
CO_2 product = v_H x vol
[At pressure P and temp. T]
if H_2O is also gas at the given condition, then,
H_2O vapour produced = $v_H\ y/2$ [At pressure P and temp. T]
If the general formula of organic compound is
$C_xH_yO_z$, the stoichiometric equation for its combustion is
$C_xH_yO_z+(x+y/4 - z/2)\ O_2 \rightarrow xCO_2 + y/2\ H_2O$.

(c) To solve the problems related to volume by volume we are to follow the given instruction.

Step 1 : Find out the chemical reaction involved in the problem and also find out the physical state of the reactants and products at the given condition.

Step 2 : Write the chemical reaction and balance it by any suitable method.

Step 3 : Write down the volume of reactants and products below the formula to each reactant and product with help of the fact that 1 gm molecule of every gaseous substance occupies 22.4 litres at NTP.

Step 4 : In case volume of the gas is measured under particular (or room) temperature. Convert it to volume at NTP by using ideal gas equation
At the same temperature, it is seen that molar ratio of reactants and products is equal to that of ratio of volumes of reactants and products, therefore, we can conclude that mole % = volume %.
Let us apply the concept to solve a problem.
A mixture of ethane (C_2H_6) and ethene (C_2H_4) occupies 40 lit at 1 atm and 400 K. The mixture reacts completely with 130 gm of O_2 to produce CO_2 and H_2O. Assuming ideal gas behaviour, calculate the mole fraction of C_2H_6 and C_2H_4 in the mixture.
Sol. Let at the given condition volume of C_2H_6 = x lit
$\therefore$ at the given condition volume of $C_2H_4 = (40 - x)$ lit

NOTES

Step 1 : C_2H_6 and C_2H_4 will be completely oxidized into CO_2 and H_2O

Step 2 : Therefore equation for their combustion are :

$C_2H_6 + 7/2\ O_2 \rightarrow 2CO_2 + 3H_2O$ ...(1)

$C_2H_4 + 3O_2 \rightarrow 2CO_2 + 2H_2O$ ...(2)

Step 3 : For x lit of C_2H_6 at the given conditions O_2 required for combustion = 7/2 x lit

For (40 – x) lit of C_2H_4 at the given conditions O_2 required for combustion = 3 (40 – x) lit.

Now, total volume of O_2 required for the combustion of the 40 lit of given mixture

= 3.5 x + 3 (40 – x) lit

= (120 + 0.5 x) lit

Now P = 1 atm V = (120 + 0.5 x) lit

T = 400 K R = 0.082 lit atm/K/mole

$$PV = \frac{W}{M} RT \quad \text{or} \quad W = \frac{PVM}{RT}$$

$$= \frac{1(120 + 0.5x) \times 32}{0.082 \times 400} = \frac{(120 + 0.5x)}{1.025}$$

Now, wt of O_2 required = 130 gm.

$$\therefore \quad \frac{(120 + 0.5x)}{1.025} = 130 \text{ or } x = 26.5$$

$\therefore$ volume fraction of $C_2H_6 = \frac{26.5}{40} = 0.6625$

volume fraction of C_2H_4 = 0.3375

$\therefore$ mole fraction of C_2H_6 = 0.6625

$\therefore$ mole fraction of C_2H_4 = 0.3375 **Ans.**

SOME IMPORTANT REACTIONS

(A) Decomposition Reaction :

$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$

$MgCO_3(s) \xrightarrow{\Delta} MgO(s) + CO_2(g)$

$SrCO_3(s) \xrightarrow{\Delta} SrO(s) + CO_2(g)$

$2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + H_2O + CO_2$

$2KHCO_3 \xrightarrow{\Delta} K_2CO_3 + H_2O + CO_2$

$2HI \xrightarrow{\Delta} H_2 + I_2$

$2NH_3 \xrightarrow{\Delta} N_2 + 3H_2$

Note : arbonates of I[st] group elements i.e. Na, K, Rb, Cs, Fr do not decompose on heating

(B) Displacement Reactions :

$Zn(s) + CuSO_4 \rightarrow ZnSO_4 + Cu$

$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$

$Pb + CuCl_2 \rightarrow PbCl_2 + Cu$

$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

Note : Precipitation reactions discussed above are also example of displacement reactions.

(C) Double Displacement Reactions :

$NaCl + AgNO_3 \rightarrow \underset{\text{white precipitate}}{AgCl} \downarrow + NaNO_3$

$NaBr + AgNO_3 \rightarrow \underset{\text{Yellow}}{AgBr} \downarrow + NaNO_3$

$NaI + AgNO_3 \rightarrow \underset{\text{Yellow}}{AgI} \downarrow + NaNO_3$

$Na_2SO_4 + BaCl_2 \rightarrow \underset{\text{white}}{BaSO_4} \downarrow + 2NaCl$

Note : (i) Similarly sulphides as HgS (Black), PbS (Black), Bi_2S_3 (Black), CuS (Black), CdS (Yellow), As_2S_3 (Yellow), Sb_2S_3 (orange), SnS (Brown), SnS_2(Yellow) give precipitate.

(ii) Carbonates of 2^{nd} group elements also give precipitate.

(D) Neutralisation Reactions :

Reaction between acids (contain replaceable H^+ ion) and bases (containing replaceable OH^- ion) is known as neutralisation reaction.

Examples :

$NaOH + HCl \rightarrow NaCl + H_2O$

$2\,KOH + H_2SO_4 \rightarrow K_2SO_4 + 2H_2O$

➢ *Examples Based On* : **GRAVIMETRIC ANALYSIS**

Ex.16 Assuming that petrol is octane (C_8H_{18}) and has density 0.8 gm/ml, 1.425 litre of petrol on complete combustion will consume -

(A) 50 mole of O_2 (B) 100 mole of O_2

(C) 125 mole of O_2 (D) 200 mole of O_2

Sol. **(C)** wt of 1.425 litre of petrol = 1.425 × 10^3 × 0.8 gm

In 1.425 litre of petrol moles of C_8H_{18} present

$$= \frac{1.425 \times 10^3 \times 0.8}{114} = 10$$

$$C_8H_{18} + \left(8 + \frac{18}{4}\right) O_2 = 8\,CO_2 + \frac{18}{2}\,H_2O.$$

or $C_8H_{18} + 12.5\,O_2 = 8\,CO_2 + 9\,H_2O$.

$\therefore$ per 10 mole of C_8H_8, moles of O_2 required = 12.5 × 10 = 125

Ex.17 1gm dry green algae absorbs 0.10528 lit of CO_2 at STP per hour by photosynthesis. If the fixed carbon atoms were all stored after the photosynthesis as starch $(C_6H_{10}O_5)_n$, how long would it take for algae to double its own weight assuming photosynthesis taking place at a constant rate.

Sol. Let, time required to double the weight of algae = t hr.
So in t hr the amount of starch produced = 1 gm.

In 1 hr CO_2 uptake = 0.10528 lit at STP. $= \frac{0.10528}{22.4}$ mole

In 1 mole of CO_2, moles of carbon present = 1.

$\therefore$ in 1 hr the amount of carbon assimilated $= \frac{0.10528}{22.4}$ mole.

$\therefore$ in t hr the amount of carbon assimilated $= \frac{0.10528}{22.4} \times t$ moles.

The entire amount of the carbon assimilated will be converted into starch.
The molecular wt. of starch = 162 n
From, 6n moles of carbon starch produced = 162 n gm

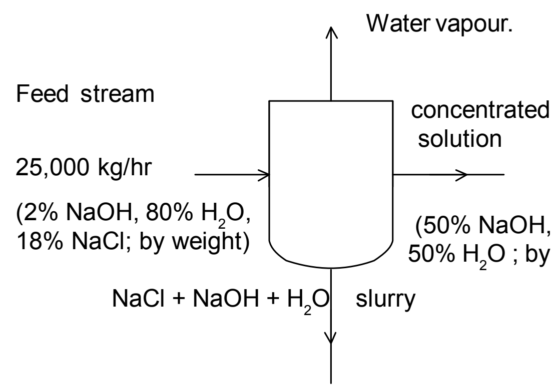
From, 1 mole of carbon starch produced $= \frac{162n}{6n}$ gm

$$\therefore \quad \frac{162}{6} \times \frac{0.10528t}{22.4} = 1$$

$$\text{or} \quad 1 = \frac{22.4 \times 6}{162 \times 0.10528} = 7.88$$

so, given green algae need 7.88 hr to double its weight **Ans.**

Ex.18 Figure shows a scheme, for concentrating a dilute solution of NaOH



Calculate :

1. How much water evaporate per hour
2. How much concentrated solution obtained per hour

Sol. Let per hour, slurry, concentrated solution and water vapour obtained x kg, y kg and z kg respectively.

By mass balance, we can write,

25, 000 = x + y + z ...(1)

By mass balance of NaOH, we can write,

25, 000 × 0.02 = 0.05 x + 0.5 y ...(2)

By mass balance of H_2O, we can write,

25, 000 × 0.8 = 0.05 x + 0.5 y + z ...(3)

By mass balance of NaCl, we can write,

25, 000 × 0.18 = x × 0.9 ...(4)

∴ x = 5000 kg/hr.

$$y = \frac{25{,}000 \times 0.02 - 0.05 \times 5000}{0.5} = 500 \text{ kg/hr}$$

∴ z = 25,000 × 0.8– 0.05 × 5000 – 0.5 × 500 = 19, 500 kg/hr

∴ Per hr water evaporate = 19,500 kg & per hr concentrated solution obtained = 500 kg **Ans.**

11. LIMITING REACTANTS OR LIMITING REAGENT

(a) In the single-reactant reactions, the calculations are carried out with only that amount of the reactant which has converted to the product.

(b) In the reactions where more than one reactants are involved, our first task prior to calculation is to determine the limiting reactant. Limiting reactant is such type of reactant which is completely consumed. All calculations has to be carried out with the amount of the limiting reactant only.

(c) How to find the limiting reactant !

$4\text{ A} + 9\text{ B} \longrightarrow 5$ moles of product

Now we have taken, 14 moles of A and 15 moles of B initially in the reaction mixture. Limiting reactant is the such type of reactant which is producing least number of moles of product. In the given example, from, 4 moles of A, product produced = 5 moles

from 14 moles of A, product produced

$$= \frac{5}{4} \times 14 \text{ moles} = 17.5 \text{ moles}$$

from, 9 moles of B, product produced

$$= \frac{5}{9} \times 15 = \frac{75}{9} \text{ moles}$$

Since reactant B is producing least number of moles of product, therefore, here B is the limiting reactant and it will be completely consumed and product will be produced on the basis of B.

➢ *Examples Based On* : **LIMITING REACTANTS**

Ex.19 A solution contains 11 gram Na_2CO_3, In this solution 11.1 gram $CaCl_2$ is added. Find out the mass of $CaCO_3$ formed.

Sol. $CaCl_2 + Na_2CO_3 \rightarrow CaCO_3 + 2NaCl$

from equation molar ratio of reactant which will react is 1:1

Now moles of Na_2CO_3 in solution $= \frac{11}{106} > 0.1$

Moles of $CaCl_2$ in solution $= \frac{11.1}{111} = 0.1$

NOTES

It means $CaCl_2$ is less in amount (In terms of moles) hence $CaCl_2$ is limiting reagent.

Now because $CaCl_2$ is limiting reagent hence it will convert completely to products because it has enough Na_2CO_3 to react.

Now 1 mole of $CaCl_2$ produces = 1 mole of $CaCO_3$

So 0.1 mole ———————— = 0.1 mole of $CaCO_3$

So mass of $CaCO_3$ produced = 0.1 × 100 = 10 grams

12. EUDIOMETRY

(a) To evaluate the composition of gases, they are allowed to react in a special type of tube known as Eudiometer tube. The tube is graduated in mm for volume measurement.

(b) The reacting gases taken in Eudiometer tube are exploded by sparks produced by passing electricity through the platinum terminals provided in the tube. The volumes of gaseous explosion products are determined by absorbing then in suitable reagents.

(c) For example,

(i) CO_2, SO_2 and Cl_2 are absorbed in the KOH solution,

(ii) O_2 is absorbed in alkaline pyrogallol, CO is absorbed in the solution of ammonical Cu_2Cl_2

(iii) O_3 is absorbed in turpentine oil.

(iv) NH_3 & HCl is absorbed in water.

(v) water vapour is absorbed in silicagel or anhydrous $CaCl_2$.

(d) The principle of eudioumetry is based on the principle of volume by volume relationship which is based on the Arogadro's law.

(e) We have already discussed the principle of volume by volume relationship, again we are reminding it. As per this principle, at the same temperature and pressure the volume of reacting gases and gaseous products maintain the same moles ratio as it was in the balanced chemical equation.

(f) Eudiometry is applied mainly to determine the composition of gas mixture and the molecular formula of gaseous reactants.

➢ *Examples Based On* **: EUDIOMETRY**

Ex.20 10 ml of gaseous hydrocarbon on combustion gives 40 ml. of CO_2 (g) and 50 ml of H_2O (vap). The hydrocarbon is -

(A) C_4H_5 (B) C_8H_{10}

(C) C_4H_8 (D) C_4H_{10}

Sol. **(D)** $C_x H_y + \left(x+\frac{y}{4}\right) O_2 = x\, CO_2 + \frac{y}{2} H_2O$

$10 \quad 10\left(x+\frac{y}{4}\right) \qquad 10x \quad 5y\ H_2O$

$\therefore$ 10 x = 40 or x = 4

5 y = 50 or y = 10

$\therefore$ molecular formula of hydrocarbon is C_4H_{10}

Ex.21 In gaseous reaction of the type aA + bB $\longrightarrow$ cC + dD, which of the following is wrong ?

(A) a litre of A combines with b litre of B to give C and D

(B) a mole of A combines with b mole of B to give C and D

(C) a gm of A combines with b gm of B to give C and D

(D) a molecule of A combines with b molecules of B to give C and D

Sol. **(C)** Atomic weights of A and B are unknown so, we cannot say, a gm of A combines with b gm of B to give C and D as per option (C).

Ex.22 50 ml of dry ammonia gas was sparked for a long time in an eudiometer tube over mercury. After sparkling, the volume becomes 97 ml. After washing the gas with water and drying, the volume becomes 94 ml. This was mixed with 60.5 ml of oxygen and the mixture was burnt. After the completion of the combustion of H_2, the volume of residual gas was 48.75 ml. Derive the molecular formula of ammonia. (Given, N and H are only present in ammonia molecule).

Sol. Let molecular formula of ammonia is N_xH_y.

$N_x H_y \longrightarrow x/2\ N_2 + y/2\ H_2$ (reaction due to sparking)

1 vol. x/2 vol y/2 vol

(50 – a) ml $\frac{ax}{2}$ ml $\frac{ay}{2}$ ml (After the sparking)

Here we have considered a ml of $N_x H_y$ has been decomposed due to sparkling.

$\therefore$ final volume of gas mixture = $(50 - a) + \frac{ax}{2} + \frac{ay}{2} = 97$

Again due to washing with water ammonia will removed, therefore volume reduction due to washing of the gas mixture with water = volume of unreacted ammonia.

$\therefore\ 50 - a = 3$ or $a = 47$

$\therefore\ \frac{ax}{2} + \frac{ay}{2} = 97 - (50 - a) = 97 - 3 = 94$

or $\frac{x}{2} + \frac{y}{2} = \frac{94}{a} = \frac{94}{47} = 2$ or $x + y = 4$ ...(1)

Among the gas remaining after washing H_2 will completely react with O_2.

$H_2(g) + \frac{1}{2}O_2\ (g) \rightarrow H_2O\ (\ell)$

$\frac{ay}{2}$ ml $\frac{ay}{4}$ ml

$\therefore$ residual $O_2 = \left(60.5 - \frac{ay}{4}\right)$ & unreacted $N_2 = \frac{ax}{2}$

$\therefore\ \frac{ax}{2} + 60.5 - \frac{ay}{4} = 48.75$

NOTES

or $a\left(\frac{x}{2}-\frac{y}{4}\right)$ = 48.75 – 60.5 = –11.75

or $a\left(\frac{y}{4}-\frac{x}{2}\right)$ = 11.75

or $\frac{y}{4}-\frac{x}{2}=\frac{11.75}{47}$ = 0.25

or y – 2x = 1 ...(2)

2x + 2y = 4 × 2 = 8

y – 2x = 1

3y = 9

∴ x = 4 – y = 4 – 3 = 1 or y = 3

∴ molecular formula of ammonia is NH_3 **Ans.**

13. CONCENTRATION TERMS

13.1 Weight by weight (w/w), weight by volume (w/v), volume by volume (v/v) percentage and mole percentage :

(a) Density (ρ) = $\frac{\text{Mass of the substance}}{\text{volume of the substance}}$

In C.G.S. and MKS units, density is expressed in gm/cm^3 or gm/ml and kg/m^3 respectively.

(b) specific gravity (d) = $\frac{\text{Mass of V volume of substance at t}^o\text{C}}{\text{Mass of V volume of water at t}^o\text{C}}$

(c) For simplification, we can conclude that the density and specific gravity of any substance is numerically same, but density has a definite unit, but specific gravity has no unit.

(d) Weight by weight percentage (%w/w)

or percentage by weight = $\frac{\text{weight of solute}}{\text{weight of solution}}\times 100$

(e) Weight by volume percentage (%w/v)

or percentage by volume = $\frac{\text{weight of solute}}{\text{volume of solution}}\times 100$

(f) Volume by volume percentage (%v/v)

or percentage by strength = $\frac{\text{volume of solute}}{\text{volume of solution}}\times 100$

(g) mole percentage (%mol/mol)

or percentage by moles = $\frac{\text{(Moles of solute)}}{\text{(Moles of solute + Moles of solvent)}}\times 100$

(h) Do remember, for the calculation of strength (%w/w, %w/v etc) the solute must be completely dissolved into the solution, otherwise, the given terminologies will be invalid.

For example, the specific gravity of gold = 19.3 gm/cm^3, if we add 193 gm gold powder in 1 litre of water,

its % w/w = $\frac{193}{1000+193}$ × 100 = 16.17 is appears to be correct, but gold is not dissolvable in water, its % w/w in water cannot be calculated.

(i) A 65% solution has the following meanings
65% by weight i.e. 100 gm solution contain 65 gm solute.
65% by volume i.e. 100 ml of solution contain 65 ml solute.
65% by strength i.e. 100 ml of solution contain 65 gm solute.
If, anything is not specified, 65% generally mean 65% by mass.

(j) For concentrated acids, like 98% H_2SO_4, 65% HNO_3 etc, if anything is not specified than percentage by mass/volume is usually considered.

➢ *Examples Based On* : **W/W, W/V, V/V, AND MOLE PERCENTAGE**

Ex.23 What is the mole fraction of ethanol in 20% by weight solution in water ?
(A) 0.095 (B) 0.089
(C) 0.9 (D) 1.2

Sol. **(B)** 100 gm of solution contain 20 gm C_2H_5OH and 80 gm of water

$\therefore$ moles of ethanol present $= \frac{20}{46} = 0.435$ (mol. wt of ethanol = 46)

$\therefore$ moles of water present $= \frac{80}{18} = 4.444$

Total no. of moles = 0.435 + 4.444 = 4.879

$\therefore$ mole fraction of $C_2H_5OH = \frac{0.435}{4.879} = 0.089$

Ex.24 When 10 ml of ethanol of density 0.7893 gm/ml is mixed with 20 ml of water of density 0.9971 gm/ml at 25ºC, the final solution has a density of 0.9571 gm/ml. Calculate the percentage change in total volume on mixing.

Sol. Mass of 10 ml of ethanol = 10 × 0.7893 = 7.893 gm.
Mass of 20 ml of water = 20 × 0.9971 = 19.94 gm.
Total mass of solution = 19.94 + 7.893= 27.83 gm.

$\therefore$ The actual volume of solution $= \frac{27.83}{0.9571} = 29.08$ ml.

Added volume of the two solution = 10 + 20 = 30 ml
$\therefore$ change in volume due to mixing = 30 – 29.08 = 0.92 ml.
$\therefore$ percentage change in volume due to mixing

$= \frac{0.92}{30} \times 100 = 3.067\%$ (contraction) **Ans.**

Ex.25 The specific gravity of a solution is 1.8, having 62% by weight of acid. It is to be diluted to specific gravity of 1.2. What volume of water should be added to 100 ml of this solution ?

Sol. Let, to 100 ml of given acid solution (Specific Gravity. 1.8) x ml. of water is added.
$\therefore$ the total volume of resulting solution = (100 + x) ml
$\therefore$ the total weight of resulting solution = (100 + x) 1.2 gm.
weight of acid present in the given acid solution (per 100 ml)
= 100 × 1.8 × 0.62

∴ the amount of water present in 100 ml of given acid solution
= 1.8 × 100 × 0.38

∴ total wt of acid present in the diluted solution
= (100 + x) 1.2 – x – 180 × 0.38
= 1.8 × 100 × 0.62

∴ 120 + 0. 2x = 180 or x = 300

∴ to lower sp. gravity of the given acid solution to 1.2, we are to add 300 ml of water per 100 ml of acid solution (sp gr. 1.2). **Ans.**

Ex.26 Fool's gold is so called because it bears a visual similarity to real gold. a block of fool's gold which measures 1.50 cm × 2.50 cm × 3.00 cm has a mass of 56.25 gm. The density of real gold is 19.3 gm/ml. What is the ratio of densities of Fool's gold to real gold ?

Sol. The volume of Fool's gold = 1.50 × 2.50 × 3.00 = 11.25 cm^3

the density of Fool's gold = $\frac{56.25}{11.25}$ gm/cm^3

∴ the ratio of densities Fool's gold to real gold is

$$= \frac{56.25}{11.25} \Big/ 19.3 = 0.259$$ **Ans.**

13.2 Parts per million (ppm), gm/litre, molarity, formality and molality :

(a) Parts per million (ppm)

It means the number of parts of solute per million parts of the weight of solution. It can be easily corelated with the % w/w or % w/v of the solution.

Let us consider m_1 gm. of solute 'x' is dissolved in m_2 gm of solvent.

∴ m_1 gm of 'X' present in (m_1 + m_2) gm of solution

∴ (m_1 + m_2) gm of solution contain m_1 gm solute

∴ 10^6 gm of solution contain $\frac{m_1}{m_1 + m_2} \times 10^6$ gm of solute

∴ ppm concentration of 'X' in the given solution = $\frac{m_1}{m_1 + m_2} \times 10^6$.

(b) Gram per litre (gm/lit) : It is the amount of solute in gm dissolved in 1 litre (1000 ml) of solution. It is one of higher unit of % mass by volume.

(c) Formality : It is the number of formula units of solute dissolved in 1 litre (1000 ml) of solution. It is similar to that of molarity.

$$\text{Formality} = \frac{\text{Number of formula units of solute}}{\text{volume of solution (in litres)}}$$

For ionic substances, since there is no actual existence of molecule or mole, so instead of molecular weight we are using the formula weight.

For example, in 20 ml of solution 7.45 gm. of KCl is present, what is the formality of the given solution ?

The formula weight of KCl = 39 + 35.5 = 74.5

∴ 7.45 gm KCl = $\frac{7.45}{74.5}$ = 0.1 formula unit of KCl

In 20 ml of solution the number of formula units of KCl present = 0.1

In 10^3 ml of solution the number of formula units of KCl present

$= \frac{0.1}{20} \times 10^3 = 0.1 \times 50 = 5$

Hence, the strength of the solution = 5(F)

where, F stands for formality.

(d) Molality : It is the number of moles of solute present per 1000 gm (1 kg) of solvent in a given solution.

$$\text{molality} = \frac{\text{Number of moles of solute}}{\text{weight of solvent (in kg.)}}$$

Let us consider, m_1 gm. of solute (mol. wt M_1) is dissolved in m_2 gm of solvent.

$\therefore$ per m_2 gm of solvent moles of solute present = $\frac{m_1}{M_1}$

$\therefore$ per 10^3 gm of solvent moles of solute present

$$= \frac{m_1}{M_1} \times \frac{1}{m_2} \times 10^3$$

Hence the molality of solution (symbol m) $= \frac{m_1}{M_1} \times \frac{1}{m_2} \times 10^3$

If $m_1 = 10$, $M_1 = 250$ and $m_2 = 100$, the strength

$= \frac{10}{250} \times \frac{1}{100} \times 10^3 = 0.4$ (m). Where, m stands for molality.

(e) Molarity : It is the number of moles of solute present per litre of solution.

$$\text{molarity} = \frac{\text{Number of moles of solute}}{\text{volume of solution (in litres)}}$$

If m_1 gm. of solute (mol. wt M_1) is dissolved in m_2 gm of solvent and if the density of resulting solution is ρ_s gm/ml then, molarity

(M) = $\frac{m_1 \times \rho_s}{M_1(m_1 + m_2)}$

1.80 gm. of glucose ($C_6H_{12}O_6$) is dissolved in 50 ml of solution, what will be the molarity of the resulting solution ?

The molecular weight of glucose = 6 × 12 + 12 × 1 + 16 × 6 = 180

$\therefore$ 1.80 gm glucose = $\frac{1.80}{180}$ = 0.01 mole of glucose

In 50 ml of solution the number of moles of $C_6H_{12}O_6$ present = 0.01 mole

$\therefore$ In 10^3 ml of solution the number of moles of $C_6H_{12}O_6$ present

$= \frac{0.01}{50} \times 10^3 = 0.2$ mole

Hence, the strength of the solution = 0.2 (M). Here, symbol (M) stands for molarity.

Now to calculate the number of moles of urea present in 150 ml of 0.05 (M) solution.
The word 0.05 (M) solution implies that in 1000 ml of the given solution 0.05 moles of urea present

In 150 ml of the given solution $= \frac{0.05}{1000} \times 150$ moles of urea present

$= 0.75 \times 10^{-3}$ moles of urea present

(Now, 1 mole = 10^3 milli moles) = 0.75 milli moles of urea present

(f) Let us consider, the molarity, molality and density (gm/ml) of a solution are M, m and ρ_s respectively. We are to find out a relation among M, m and ρ_s. Given that the molecular weight of solute is M.
From the given data we can write,

10^3 ml of solution contain M moles of solute

$\Rightarrow$ 10^3 ml of solution contain MM^1 gm of solute

$\Rightarrow$ $10^3 \times \rho_s$ gm of solution contain MM^1 gm of solute

$\therefore$ the weight of solvent present in 10^3 ml of solution $= (10^3 \times \rho_s - MM')$ gm

$\therefore$ molality of solution (m) $= \dfrac{M \times 10^3}{(10^3 \rho_s - MM')}$

or $\dfrac{1}{m} = \dfrac{10^3 \rho_s - MM'}{M \times 10^3}$

or $\dfrac{1}{m} = \dfrac{\rho_s}{M} - \dfrac{M'}{10^3}$

or $\dfrac{\rho_s}{M} = \dfrac{1}{m} + \dfrac{M'}{10^3}$

or $\dfrac{1}{M} = \dfrac{1}{\rho_s}\left(\dfrac{1}{m} + \dfrac{M'}{10^3}\right) = \dfrac{1}{\rho_s}\left(\dfrac{1}{m} + 10^{-3} \times M'\right)$

➢ *Examples Based On* : **FORMALITY, MOLARITY, MOLALITY AND PPM CONCENTRATION**

Ex.27 In a solution the concentration of $MgCl_2$ is 5 (m), the sp gravity of the solution is 1.05 the concentration of Cl^- in the solution is -

(A) 10 (M) (B) 20 (M)
(C) 18.5 (M) (D) 7.12 (M)

Sol. **(D)** $\dfrac{1}{M} = \dfrac{1}{\rho_s}\left(\dfrac{1}{m} + 10^{-3} \times M'\right)$

$M' = 24 + 71 = 95$ $\rho_s = 1.05$

$\therefore$ $\dfrac{1}{M} = \dfrac{1}{1.05}\left(\dfrac{1}{5} + 0.095\right)$

or M = 3.56

or $\dfrac{1}{M} = 0.28095$

$\therefore$ molarity of $MgCl_2$ = 3.56

$\therefore$ concentration of Cl^- = 3.56 × 2 = 7.12 (M)