MOLE CONCEPT

Preface

Mole concept is one of the fundamental concept to understand the quantitative aspects of chemical analysis.

Sincere attempt have been made to make it clear. Before studying the mole concept, students are advised to clear their concepts in symbol and formula of different substances.

It is necessary to remember the atomic numbers and atomic masses of some common elements to solve the problems of mole concept. It is to be remembered that mole concept is the heart of any chemical analysis. Without the sound knowledge in mole concept the word "chemistry", will surely sound as "Bigmistry".

This book consists of theoretical & practical explanations of all the concepts involved in the chapter. Each article is followed by a ladder of illustration. At the end of the theory part, there are miscellaneous solved examples which involve the application of multiple concepts of this chapter.

Students are advised to go through all these solved examples in order to develop better understanding of the chapter and to have better grasping level in the class.

Total number of Questions in **Mole concept** are :

In chapter Examples	....	11
Solved Examples	....	20
Total numbers of questions	....	**31**

1. SIGNIFICANT FIGURES

(A) Every scientific observation involves some degree of uncertainity depending upon the limitation of instrument. To represent scientific data, role of significant figures has its own importance.

(B) Significant figures are equal to the number of digits in numbers with last digit uncertain and rest all are certain digits i.e. all the digits of datum including the uncertain one , are called significant figures.

(C) **Rules for determination significant figure:**

(i) All non zero digits are significant.

Example : 3.14 has three significant figures

(ii) The zeros to the right of the decimal point are significant.

Example : 3.0 has two significant figures.

(iii) The zeros to the left of the first non zero digit in a number are not significant.

Example : 0.02 has one significant figure.

(iv) The zeros between two non zero digits are also significant.

Example : 6.01 has three significant figures.

(v) **Exponential form :** $N \times 10^n$. Where N show the significant figure.

Example : 1.86×10^4 has three significant figure.

(vi) **Rounding off the uncertain digit :**

(a) If the left most digit to be rounded off is more than 5 , the preceding number is increased by one.

Example : 2.16 is rounded to 2.2

(b) If the left most digit to be rounded off is less than 5 , the preceding number is retained.

Example : 2.14 is rounded off to 2.1

(c) If the left most digit to be rounded off is equal to 5 , the preceding number is not changed if it is even and increased by one if it is odd.

Example : 3.25 is rounded off to 3.2
2.35 is round off to 2.4

2. TYPES OF MIXTURE

2.1 Heterogenous mixture

A mixture in which the different constituents are not distributed uniformly is known as heterogenous mixture. eg Water

2.2 Homogenous mixture

A mixture in which the different constituents are uniformly distributed is known as homogenous mixture. eg. O_2, N_2 etc.

3. LAWS OF CHEMICAL COMBINATION

3.1 Law of conservation of mass-[Lavoisier, 1744]

(a) According to this law , matter is neither created nor destroyed in the course of chemical reaction although it may change from one form to other

(b) This law contradicts nuclear reactions where Einestein equation is applicable

(c) According to this law , sum of the masses of product formed is always equal to the sum of the masses of the reactant undergone change

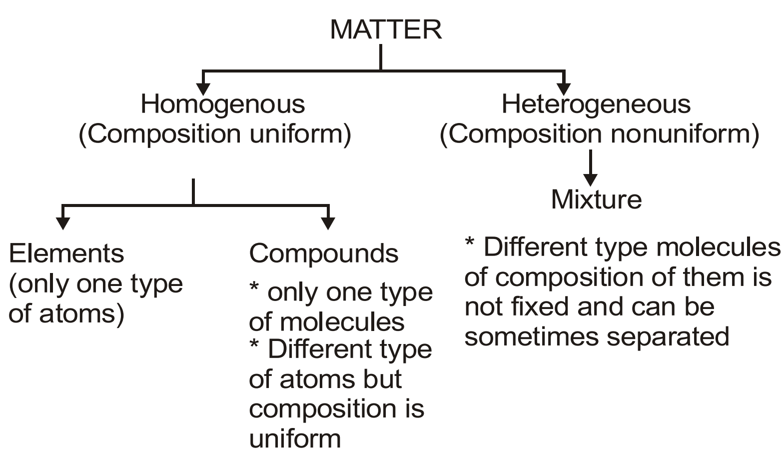
Example : $H_2 + Cl_2 \rightarrow 2HCl$

2gm 71 gm 2 [1 + 35.5]

$\Rightarrow$ 73 gm $\equiv$ 73 gm

3.2 Law of definite proportion [Proust, 1799]

(a) According to the law , the composition of a compound always remains a constant i.e. the ratio of weights of different elements in a



compound ; no matter by whatever method , it is prepared or obtained from different sources, remains always a constant

Example : In H_2O ratio of weight = 1 : 8

In CO_2 ratio of weight = 3 : 8

3.3 Law of multiple proportion [John Dalton, 1804]

According to this law, when two elements A and B combine to form more than one chemical compound then different weights of A , which combine with a fixed weight of B , are in a proportion of simple whole number

Example : CO & CO_2

12 : 16 & 12 : 32

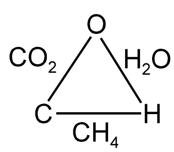
ratio = 16 : 32

= 1 : 2

3.4 Law of reciprocal proportions [Ritche, 1792-94]

When two elements combines separately with third element and form different types of molecules, their combining ratio is directly reciprocated if they combine directly

Example :



C with H form methane and with O form CO_2. In CH_4 , 12 grams of C reacts with 4 grams of H whereas in CO_2 12 gram of C reacts with 32 grams of O. Therefore when H combines with O they should combine in the ratio of 4 : 32 (i.e. = 1 : 8) or in simple multiple of it. The same is found to be true in H_2O molecule. The ratio of weights of H and O in Water is 1 : 8

3.5 Gay-Lussac's [1808] law of combining volumes :

This law states that under similar conditions of pressure and temperature, volume ratio of gases is always in terms of simple integers.

Ex.

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

vol. ratio 1 : 3 : 2

4. AVOGADRO'S HYPOTHESIS

According to this under similar conditions of pressure and temperature , equal volumes of gases contain equal number of molecules.

4.1 Salient features of Avogadro's hypothesis

(1) It has removed the anomaly between Dalton's atomic theory and Gay Lussac's law of volume by making a clear distinction in between atoms and molecules

(2) It reveals that common elementary gases like hydrogen , nitrogen , oxygen etc. are diatomic

(3) It provides a method to determine the atomic weights of gaseous elements

(4) It provides a relationship between vapour density and molecular weight of substances

Vapour density =

$$\frac{\text{Volume of definite amount of Gas}}{\text{Volume of same amount of Hydrogen}}$$

or Vapour density =

$$\frac{\text{Weight of n molecules of Gas}}{\text{Weight of n molecules of Hydrogen}}$$

or Vapour density =

$$\frac{\text{Weight of one molecule of Gas}}{\text{Weight of one atom of hydrogen} \times 2}$$

or Vapour density = $\frac{\text{Molecular weight}}{2}$

(5) It helps to determine molar volume

Molecular weight of the gas

= 2 × vapour density

$$= 2 \times \frac{\text{Weight of 1 litre of the Gas at S.T.P}}{\text{Weight of 1 litre of Hydrogen at S.T.P}}$$

$$= 2 \times \frac{\text{Weight of 1 litre of the Gas at S.T.P}}{0.089 \text{ gm}}$$

$$= \frac{2}{0.089} \times \text{Weight of 1 litre of the gas at S.T.P.}$$

= 22.4 × Weight of 1 litre of gas at S.T.P.

= Weight of 22.4 litre of the gas at S.T.P

5. ATOM, MOLECULES AND MOLECULAR FORMULA

Atom: It is the smallest particle of an element that takes part in a chemical reaction and not capable of independent existence.

Molecule : It is the smallest particle of matter which is capable of independent existence. A molecule is generaly an assembly of two or more tightly bonded atoms.

Homoatomic molecules : Molecules of an element contain one type of atoms. eg. O_2, Cl_2 etc.

Heteroatomic molecules : Molecules of compounds contain more that one type of atom. eg. H_2O, HCl etc

5.1 Atomic mass scale

(A) Oxygen as standard : The standard reference for atomic weight may be oxygen with an assigned value of 16.

Atomic weight of an element =

$$\frac{\text{Weight of 1 atom of element}}{1/16 \times \text{Weight of 1 atom of oxygen}}$$

(B) Carbon as standard : The modern reference standard for atomic weight is carbon isotope of mass number 12.

Atomic weight of an element =

$$\frac{\text{Weight of 1 atom of the element}}{1/12 \times \text{Weight of 1 atom of C} - 12}$$

IMPORTANT POINTS

(1) Atomic weight is not a weight but a number.

(2) Atomic weight is not absolute but relative to the weight of the standard reference element C-12

5.2 Molecular weight

It is the number of times a molecule is heavier than $1/12^{th}$ of an atom of C – 12.

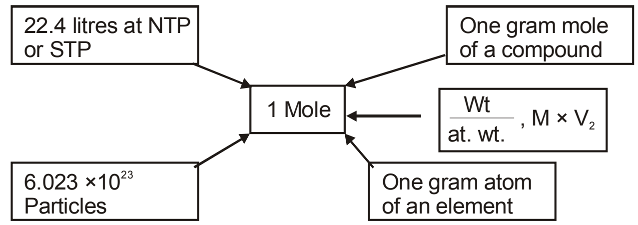
$$\text{Molecular weight} = \frac{\text{Weight of 1 molecule}}{1/12 \times \text{Weight of one C} - 12}$$

IMPORTANT POINTS

1. Molecular weight is not a weight but a number
2. Molecular weight is relative and not absolute
3. Molecular weight expressed in grams is called gram molecular weight
4. Molecular weight is calculated by adding all the atomic weights of all the atoms in a molecule

Example : $CO_2 = 12 + 2 \times 16 = 44$

6. MOLE CONCEPT



IMPORTANT :

⇒ 1 mole = 6.023×10^{23} particles

⇒ 1 mole atoms = 6.023×10^{23} atoms

⇒ One mole molecule = 6.023×10^{23} molecules

⇒ Mass of one mole of atoms = Gram atomic mass

⇒ Mass of one mole of molecules = Gram molecular mass

⇒ Moles of a compound = Mass of compound

⇒ Volume occupied by 1 mole of a gas at N.T.P = 22.4 litres.

Examples based on Mole Concept

Ex.1 Calculate the number of molecules of sulphur dioxide in 0.064 g of the gas

Sol. Gram molecular weight of sulphur dioxide (SO_2) = 64gm

Given mass = 0.064 gm

A gram molecular weight of any gas contain avogadro number of molecules $= 6.023 \times 10^{23}$

∴ 0.064 g of sulphur dioxide contain $\left(\frac{6.023 \times 10^{23}}{1000}\right)$ molecules $= 6.023 \times 10^{20}$

Ex.2 Which of the following contains the least number of molecules -

(A) 16g of CO_2

(B) 8g of O_2

(C) 4g of N_2

(D) 2g of H_2

Sol. [C]

(1) No. of moles of $CO_2 = \frac{\text{Weight}}{\text{Molecular weight}}$

$$= \frac{16}{44} = 0.36$$

(2) Number of moles of O_2 = $\frac{8}{32}$ = 0.25

(3) Number of moles of N_2 = $\frac{4}{28}$ = 0.14

(4) Number of moles of H_2 = $\frac{2}{2}$ = 1

Ex.3 Atomic weight of helium is 4. Calculate the number of atoms in 1g of helium -

Sol. 4g of Helium contains 6.023×10^{23} atoms

1g of Helium contains $\frac{6.023 \times 10^{23}}{4}$ = 1.506 $\times 10^{23}$ atoms

Ex.4 What is the mass of 1 molecule of CO -

Sol. Gram molecular weight of CO = 12 + 16 = 28g

6.023×10^{23} molecules of CO weighs 28gm

1 molecule of CO weighs

$= \frac{28}{6.02 \times 10^{23}} = 4.65 \times 10^{-23}$ g

Ex.5 Calculate the volume at STP occupied by 240gm of SO_2

Sol. Molecular weight of SO_2 = 32 + 2 × 16 = 64

64 gm of SO_2 occupies 22.4 litre at STP

240 gm of SO_2 occupies = $\frac{22.4}{64}$ × 240 = 84 litre at STP

Ex.6 Calculate the number of atoms in each of the following -

(a) 52 mole of He

(b) 52 amu

(c) 52 g of He

Sol. (a) 1mole of He contain 6.02×10^{23} atoms

∴ 52 mole of He contain

$= 52 \times 6.02 \times 10^{23} = 31.3 \times 10^{24}$ atoms

(b) Atomic weight of He = 4amu

∴ 52 amu of He contain = $\frac{52}{4}$

= 13 atoms of He

(c) Number of moles of He in 52g

$= \frac{52}{4}$ = 13 moles

∴ no. of atoms in 52g of He i.e. 13 moles

$= 13 \times 6.02 \times 10^{23}$ atoms

$= 78.26 \times 10^{23}$ atoms

7. CHEMICAL FORMULAS

It is of two types-

[A] Molecular formulae : Chemical formulae that indicate the actual numbers and type of atoms in a molecule are called molecular formulae

[B] Empirical formulae : The chemical formulae that give only the relative number of atoms of each type in a molecule are called empirical formulae

7.1 Determination of chemical formulae

[A] Determination of empirical formulae :

Step - I : Determination of percentage

Step - II : Determination of mole ratio

Step - III : Making it whole number ratio

Step - IV : Removal of fractions from mole ratio

[B] Determination of molecular formulae

(i) First of all find empirical formulae

(ii) Molecular formulae = (Empirical formulae) n

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

Examples based on Chemical Formulas

Ex.7 Phosgene , a poisonous gas used during World war-I, contains 12.1% C, 16.2% O and 71.7% Cl by mass. What is the empirical formula of phosgene.

Sol.

Element	%	Mole ratio	Simplest mole ratio
C	12.1	$\frac{12.1}{12}$ = 1.01	$\frac{1.01}{1.01}$ =1
O	16.2	$\frac{16.2}{16}$ = 1.01	$\frac{1.01}{1.01}$ = 1
Cl	71.7	$\frac{71.7}{35.5}$ = 2.02	$\frac{2.02}{1.01}$ = 2

Then empirical formula = $COCl_2$

Ex.8 5.325g sample of methyl benzoate, a compound used in the manufacture of perfumes is found to contain 3.758 g of carbon, 0.316g hydrogen and 1.251g of oxygen. What is empirical formulae, of compound. If mol. weight of methyl benzoate is 136.0, calculate its molecular formula.

Sol.

Element	%	Mole ratio	Simplest whole ratio
C	$\frac{3.758 \times 100}{5.325} = 70.57$	$\frac{70.57}{12} = 5.88$	$\frac{5.88}{1.47} = 4$
H	$\frac{0.316 \times 100}{5.325} = 5.93$	$\frac{5.93}{1} = 5.93$	$\frac{5.93}{1.47} = 4$
O	$\frac{1.251 \times 100}{5.325} = 23.50$	$\frac{23.50}{16} = 1.47$	$\frac{1.47}{1.47} = 1$

Empirical $= C_4H_4O$

$$n = \frac{\text{Mol. wt}}{\text{Empirical formula wt.}} = \frac{136}{68} = 2$$

$\Rightarrow$ Molecular formula $= C_8H_8O_2$

8. CHEMICAL EQUATION

Representation of the chemical change in terms of symbol and formulae of the reactants & products is called a chemical equation.

8.1 Information conveyed by a chemical equation

(1) Qualitatively , a chemical equation tells us the names of the various reactants

(2) Quantitatively , it express

(a) The relative no. of molecules of reactants and products

(b) The relative no. of moles of reactant and products

(c) The relative masses of reactants and products

(d) The relative volumes of gaseous reactants and products

8.2 Limitations of chemical equations

(1) The physical state of the reactants and products

(2) The dilution of solution of reactants and products are in soluble state

(3) The energy changes during chemical reaction

(4) The conditions of P, T etc at which reaction occurs.

(5) The rate of chemical reaction

8.3 Limiting Reagent

It may be defined as the reactant which is completely consumed during the reaction is called limiting reagent-

Example : $2H_2(g) + O_2(g) \rightarrow 2H_2O$

Here H_2 is known as limiting reagent.

Examples based on Chemical Reactions

Ex.9 Calculate the mass of oxygen required to burn 14g C_2H_4 completely-

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

Mole ratio 1 3 2 4

Moles of C_2H_4 to be burnt $= \frac{14}{28} = \frac{1}{2}$ mole.

$\because$ 1 mole C_2H_4 requires 3 mol O_2 for combustion

$\therefore$ $\frac{1}{2}$ mole C_2H_4 requires $3 \times \frac{1}{2}$ mole $O_2 = \frac{3}{2}$ mol O_2. & Mass of $O_2 = \frac{3}{2} \times 32 = 48$ gm

Ex.10 Calculate the weight and volume of H_2 at STP that will be displaced by 1 gram of Zn when it is completely dissolved in dilute sulphuric acid.

Sol. $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

1 atom 1 molecule 1 molecule 1 molecule

1 gram-atom 1 mole

65.4 g 2 gm or 22.4 dm^3

$\because$ 65.4 g of Zn displaces 2g of Hydrogen

$\therefore$ 1.0 g of Zn displaces $\frac{2}{65.4} \times 1$

$= 0.0306$ g of H_2

$\because$ 65.4 g of Zn displaces 22.4 dm^3 of H_2 at S.T.P.

$\therefore$ 1.0 g of Zn displaces $\frac{22.4}{65.4} \times 1.0$

$= 0.3425\ dm^3$

Ex.11 10 ml of liquid carbon disulphide (sp. gravity 2.63) is burnt in oxygen. Find the volume of the resulting gases measured at STP.

Sol. 1 ml of CS_2 Weighs 2.63 g

10 ml of CS_2 weighs 26.3 g

$CS_2 + 3O_2 \rightarrow CO_2 + 2SO_2$

$12 + (2 \times 32)$ 22.4 ℓ 44.8 ℓ

76 gm 67.2ℓ.

$\because$ 76g of CS_2 will yield 67.2 ℓ of a mixture of CO_2 and SO_2 at STP

$\therefore$ 26.3 g of CS_2 would yield $\frac{67.2}{76} \times 26.3$

$= 23.26$ lit.

SOLVED EXAMPLE

Ex.1 8 litre of H_2 and 6 litre of Cl_2 are allowed to react to maximum possible extent. Find out the final volume of reaction mixture. Suppose P and T remains constant throughout the course of reaction -

(A) 7 litre (B) 14 litre
(C) 2 litre (D) None of these.

Sol. **(B)**

	H_2	+ Cl_2	$\rightarrow$ 2 HCl
Volume before reaction	8 lit	6 lit	0
Volume after reaction	2	0	12

$\therefore$ Volume after reaction

= Volume of H_2 left + Volume of HCl formed

= 2 + 12 = 14 lit

Ex.2 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

$$= \frac{\%\text{ of I isotope} \times \text{its atomic mass} + \%\text{ of II isotope} \times \text{its atomic mass}}{100}$$

$$= \frac{75.53 \times 34.969 + 24.47 \times 36.96}{100}$$

= 35.5 amu.

Ex.3 Calculate the mass in gm of 2g atom of Mg-

(A) 12 gm (B) 24 gm
(C) 6 gm (D) None of these.

Sol. **(D)**

$\because$ 1 gm atom of Mg has mass = 24 gm

$\therefore$ 2 gm atom of Mg has mass

= 24 x 2 = 48 gm.

Ex.4 In 5 g atom of Ag (At. wt. of Ag = 108), calculate the weight of one atom of Ag -

(A) 17.93×10^{-23}gm (B) 16.93×10^{-23} gm
(C) 17.93×10^{23} gm (D) 36×10^{-23} gm

Sol. **(A)**

$\because$ N atoms of Ag weigh 108 gm

$\therefore$ 1 atom of Ag weigh $= \frac{108}{N}$

$$= \frac{108}{6.023 \times 10^{23}}$$

$= 17.93 \times 10^{-23}$ gm.

Ex.5 In 5g atom of Ag (at. wt. = 108), calculate the no. of atoms of Ag -

(A) 1 N (B) 3N
(C) 5 N (D) 7 N.

Sol. **(C)**

$\because$ 1 gm atom of Ag has atoms = N

$\therefore$ 5 gm atom of Ag has atoms = 5N.

Ex.6 Calculate the mass in gm of 2N molecules of CO_2 -

(A) 22 gm (B) 44 gm
(C) 88 gm (D) None of these.

Sol. **(C)**

$\because$ N molecules of CO_2 has molecular mass = 44.

$\therefore$ 2N molecules of CO_2 has molecular mass = 44 x 2 = 88 gm.

Ex.7 How many carbon atoms are present in 0.35 mol of $C_6H_{12}O_6$ -

(A) 6.023×10^{23} carbon atoms

(B) 1.26×10^{23} carbon atoms

(C) 1.26×10^{24} carbon atoms

(D) 6.023×10^{24} carbon atoms

Sol. **(C)**

$\therefore$ 1 mol of $C_6H_{12}O_6$ has = 6 N atoms of C

$\therefore$ 0.35 mol of $C_6H_{12}O_6$ has

= 6 × 0.35 N atoms of C

= 2.1 N atoms

$= 2.1 \times 6.023 \times 10^{23} = 1.26 \times 10^{24}$ carbon atoms

Ex.8 How many molecules are in 5.23 gm of glucose ($C_6H_{12}O_6$) -

(A) 1.65×10^{22} (B) 1.75×10^{22}
(C) 1.75×10^{21} (D) None of these

Sol. **(B)**

$\because$ 180 gm glucose has = N molecules

$\therefore$ 5.23 gm glucose has $= \frac{5.23 \times 6.023 \times 10^{23}}{180}$

$= 1.75 \times 10^{22}$ molecules

Ex.9 What is the weight of 3.01×10^{23} molecules of ammonia -

(A) 17 gm (B) 8.5 gm
(C) 34 gm (D) None of these

Sol. **(B)**

$\because$ 6.023×10^{23} molecules of NH_3 has weight = 17 gm

$\therefore$ 3.01×10^{23} molecules of NH_3 has weight

$= \dfrac{17 \times 3.01 \times 10^{23}}{6.023 \times 10^{23}}$

= 8.50 gm

Ex.10 How many significant figures are in each of the following numbers -

(a) 4.003 (b) 6.023×10^{23} (c) 5000

(A) 3, 4, 1 (B) 4, 3, 2
(C) 4, 4, 4 (D) 3, 4, 3

Sol. **(C)**

Ex.11 How many molecules are present in one mℓ of water vapours at STP -

(A) 1.69×10^{19} (B) 2.69×10^{-19}
(C) 1.69×10^{-19} (D) 2.69×10^{19}

Sol. **(D)**

$\because$ 22.4 litre water vapour at STP has

$= 6.023 \times 10^{23}$ molecules

$\therefore$ 1×10^{-3} litre water vapours at STP has

$= \dfrac{6.023 \times 10^{23}}{22.4} \times 10^{-3} = 2.69 \times 10^{+19}$

Ex.12 How many years it would take to spend Avogadro's number of rupees at the rate of 1 million rupees in one second -

(A) 19.098×10^{19} years
(B) 19.098 years
(C) 19.098×10^{9} years
(D) None of these

Sol. **(C)**

$\because$ 10^6 rupees are spent in 1sec.

$\therefore$ 6.023×10^{23} rupees are spent in

$= \dfrac{1 \times 6.023 \times 10^{23}}{10^6}$ sec

$= \dfrac{1 \times 6.023 \times 10^{23}}{10^6 \times 60 \times 60 \times 24 \times 365}$ years

$= 19.098 \times 10^{9}$ year

Ex.13 An atom of an element weighs 6.644×10^{-23} g. Calculate g atoms of element in 40 kg-

(A) 10 gm atom (B) 100 gm atom
(C) 1000 gm atom (D) 10^4 gm atom

Sol. **(C)**

$\because$ weight of 1 atom of element

$= 6.644 \times 10^{-23}$ gm

$\therefore$ weight of 'N' atoms of element

$= 6.644 \times 10^{-23} \times 6.023 \times 10^{23} = 40$ gm

$\because$ 40 gm of element has 1 gm atom.

$\therefore$ 40 x 10^3 gm of element has $\dfrac{40 \times 10^3}{40}$

$= 10^3$ gm atom.

Ex.14 Calculate the number of Cl^- and Ca^{+2} ions in 222 g anhydrous $CaCl_2$ -

(A) 2N ions of Ca^{+2} 4 N ions of Cl^-
(B) 2N ions of Cl^- & 4N ions of Ca^{+2}
(C) 1N ions of Ca^{+2} & 1N ions of Cl^-
(D) None of these.

Sol. **(A)**

$\because$ mol. wt. of $CaCl_2$ = 111 g

$\because$ 111 g $CaCl_2$ has = N ions of Ca^{+2}

$\therefore$ 222g of $CaCl_2$ has $\dfrac{N \times 222}{111}$

= 2N ions of Ca^{+2}

Also $\because$ 111 g $CaCl_2$ has = 2N ions of Cl^-

$\therefore$ 222 g $CaCl_2$ has = $\dfrac{2N \times 222}{111}$ ions of Cl^-

= 4N ions of Cl^- .

Ex.15 The density of O_2 at NTP is 1.429g / litre. Calculate the standard molar volume of gas-

(A) 22.4 lit. (B) 11.2 lit
(C) 33.6 lit (D) 5.6 lit.

Sol. **(A)**

$\because$ 1.429 gm of O_2 gas occupies volume = 1 litre.

$\therefore$ 32 gm of O_2 gas occupies = $\dfrac{32}{1.429}$

= 22.4 litre/mol.

Ex.16 Which of the following will weigh maximum amount-

(A) 40 g iron

(B) 1.2 g atom of N

(C) 1×10^{23} atoms of carbon

(D) 1.12 litre of O_2 at STP

Sol. **(A)**

(A) Mass of iron = 40 g

(B) Mass of 1.2 g atom of

N = 14 × 1.2 = 16.8 gm

(D) Mass of 1×10^{23} atoms of C

$$= \frac{12 \times 1 \times 10^{23}}{6.023 \times 10^{23}} = 1.99 \text{ gm.}$$

(D) Mass of 1.12 litre of O_2 at STP

$$= \frac{32 \times 1.2}{22.4} = 1.6 \text{ g}$$

Ex.17 How many moles of potassium chlorate to be heated to produce 11.2 litre oxygen -

(A) $\frac{1}{2}$ mol (B) $\frac{1}{3}$ mol

(C) $\frac{1}{4}$ mol (D) $\frac{2}{3}$ mol.

Sol. **(B)**

$$2\,KClO_3 \rightarrow 2KCl + 3O_2$$

Mole for reaction 2 2 3

$\because$ 3 × 22.4 litre O_2 is formed by 2 mol $KClO_3$

$\therefore$ 11.2 litre O_2 is formed by $\frac{2 \times 11.2}{3 \times 22.4}$

$= \frac{1}{3}$ mol $KClO_3$

Ex.18 Calculate the weight of lime (CaO) obtained by heating 200 kg of 95% pure lime stone ($CaCO_3$).

(A) 104.4 kg (B) 105.4 kg

(C) 212.8 kg (D) 106.4 kg

Sol. **(D)**

$\because$ 100 kg impure sample has pure $CaCO_3$ = 95 kg

$\therefore$ 200 kg impure sample has pure $CaCO_3$

$$= \frac{95 \times 200}{100} = 190 \text{ kg.}$$

$CaCO_3 \rightarrow CaO + CO_2$

$\because$ 100 kg $CaCO_3$ gives CaO = 56 kg.

$\therefore$ 190 kg $CaCO_3$ gives CaO = $\frac{56 \times 190}{100}$

= 106.4 kg.

Ex.19 The chloride of a metal has the formula MCl_3. The formula of its phosphate will be-

(A) M_2PO_4 (B) MPO_4

(C) M_3PO_4 (D) $M(PO_4)_2$

Sol. **(B)** $AlCl_3$ as it is $AlPO_4$

Ex.20 A silver coin weighing 11.34 g was dissolved in nitric acid. When sodium chloride was added to the solution all the silver (present as $AgNO_3$) was precipitated as silver chloride. The weight of the precipitated silver chloride was 14.35 g. Calculate the percentage of silver in the coin -

(A) 4.8 % (B) 95.2%

(C) 90 % (D) 80%

Sol. **(B)**

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

108

$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$

143.5

$\therefore$ 143.5 gm of silver chloride would be precipitated by 108 g of silver.

or 14.35 g of silver chloride would be precipitated 10.8 g of silver.

$\because$ 11.34 g of silver coin contain 10.8 g of pure silver.

$\therefore$ 100 g of silver coin contain $\frac{10.8}{11.34} \times 100$

= 95.2 %.