Plus One Chemistry Chapter Wise Important Questions Chapter 1 Some Basic Concepts of Chemistry

Question 1.
Calculate the number of moles of o2 required to produce 240g of MgO by burning Mg metal. (Atomic mass: Mg = 24, o = 16) (March – 2019)
Answer:
$Mg + 1/2\ O_2 \rightarrow MgO$
Molecular mass of MgO = 24 + 16 = 40
No. of moles of MgO present in 240g of MgO = 6 moles
As per the balanced equation, we have to use 6 moles of Mg and 3 moles of O_2 in order to get 240g of MgO.

Question 2.
If the mass percent of the various elements of a compound is known, its empirical formula can be calculated. (March – 2010)
a) What is mass percent? Give its mathematical expression.
b) A compound contains 4.07% hydrogen, 24.27% carbon and 71.65% chlorine. Its molecular mass is 98.96. What are the empirical and molecular formulae?
Answer:
a) Mass percentage of a component in a solution is the weight of that component present in 100gm of the solution.

b)

$$\text{Mass percentage} = \frac{\text{Mass of solutex100}}{\text{Mass of solution}}$$

Element	%	Relative no. of atoms	Dividing by smallest factor	Rate of no. atoms
C	24.27%	24.2712=2.02	2.012.01=2.01	1
H	4.07%	4.071=4.07	4.072.01=2.02	2
Cl	71.65%	71.6535.5=2.018	2.012.01=1	1

Empirical formula = CH_2CI
Molecular mass = 98.96

Question 3.
One mole is the amount of substance that contains as many elementary particles as 12g of ^{12}C isotope of carbon. (Say – 2010)
a) What do you mean by molar mass of a compound?
b) Calculate the number of moles in 1 litre of water (Density of water is 1g/mL). Also calculate the number of water molecules in 1 litre of water.

Answer:
a) The mass of one mole of any substance is called the molar mass.

Empirical formula mass = 12x1+1x2+1x35.5
= 12+2+35.5 = 49.5

$$n = \frac{98.96}{49.5} = 2$$

Molecular formula of the compound= $(CH_2Cl) \times 2$
= $C_2H_4Cl_2$

b) $$\text{Number of moles} = \frac{\text{Mass of the substance}}{\text{Gram molecular mass}}$$

Number of moles in 1 litre of water

$$= \frac{1000\ gL^{-1}}{18\ g\ mol^{-1}} = 55.55\ mol\ L^{-1}$$

Number of water molecules in 1 litre

$$= 55.55 \times 6.022 \times 10^{23}$$

Question 4.
The laws of chemical combination are the basis of the atomic theory. (March – 2011)
a) Name the law of chemical combination illustrated by the pair of compounds, CO and CO_2.
b) State and explain the law of conservation of mass.
c) Calculate the molarity of a solution containing 8g of NaOH in 500 mLofwater.
Answer:
a) Law of multiple proportions.
b) Matter can neither be created nor be destroyed.

$$CaCO_3 \xrightarrow{heat} CaO + CO_2$$

100g of $CaCO_3$ on heating given 56 grm of CaO and 44 gram of CO_2.

c) $$\text{Molarity} = \frac{\text{wt of the solute in grams}}{\text{M.wt of solute x volume of solution in litre}}$$

$$\frac{8}{40 \times 0.5} = \frac{1}{2.5} = \underline{0.4M}$$

Question 5.
The laws of chemical combination govern the formation of compounds from elements. (Say – 2011)
a) State the law of conservation of mass. Who put forward this law?
b) The following data are obtained when dinitrogen and dioxygen react together to from different compounds.

Sl. No.	Mass of dinitrogen (g)	Mass of dioxygen (g)
1	14	16
2	14	32
3	28	48
4	28	80

Which law of chemical combination is illustrated by the above experimental data? Explain.
Answer:
a) It states that matter can neither be created nor destroyed.

OR
The total mass of the reactants in a chemical reaction is equal to the total mass of the products formed. This law was proposed by Antoine Lavoisier,

b) Law of multiple proportions proposed by John Dalton.

According to this law if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers.

Here, the oxides of nitrogen are,

NO	NO_2	N_2O_3 and	N_2O_5
14g:16g	14g:32g	28g:48g	28g:80g

The different masses of oxygen which combine with a fixed mass (28g) of nitrogen are in the ratio 32:64:48:80 = 2:4:3:5, which is a simple whole-number ratio. Hence, the law is verified.

Question 6.
The combination of elements ot form compounds is governed by the laws of chemical combination. (March – 2012)
a) Hydrogen combines with oxygen to form compounds, namely water and hydrogen peroxide. State and illustrate the related law of chemical combination.
b) What is meant by 'limiting reagent' in a chemical reaction?
c) 28 g of nitrogen is mixed with 12 g of hydrogen to form ammonia as per the reaction, $N_2 + 3H_2 \rightarrow 2NH_3$. Which is the 'limiting reagent' in this reaction. [Atomic masses : N = 14, H = 1]
Answer:
a) Law of multiple proportions

Hydrogen + Oxygen → Water
2g + 16g - 18g
Hydrogen + Oxygen → Hydrogen Peroxide
2g 32g 34g

Here, the masses of oxygen (ie, 16g and 32g)which combine with a fixed mass of hydorgen (2g) bear a simple ratio ie 16:32 or 1:2.

Law of multiple proportions states that if two elements combine to form two or more compound, the weights of one of elements which combine with a fixed weight of the other in these compound. bear simple whole-number ratio by weight.

b) The reactant which is completely used up first in a reaction is called the limiting reactant.
c) $N_2 - 3H_2 \rightarrow 2NH_3$
28g 6g → 34g
28g 12g → (given data)
∴ N_2 is the limiting reagent. (N_2 is completely used up in a reaction)

Question 7.

a) Mole is a very large number to indicate the number of atoms, molecules, etc. Write another name for one mole. (Say – 2012)

b) i) How the molecular formula is different from that of the Empirical formula?

ii) An organic compound on analysis gave the following composition.

Carbon=40%, Hydrogen=6.66% and oxy- gen=53.34%. Calculate its molecular formula if its molecular mass is 90.

Answer:

a) 1 mole = gram atomic mass or 1 gram atom

1 mole = gram molecular mass or 1 gram molecule

b) 1) Empirical formula of a compound is defined as the simplest formula that gives the ratio of the various elements in a molecules.

Eg: Empirical formula of benzene is CH. Molecular formula of a compound gives actual number of atoms of each element present in a molecular of the compound, eg: molecular formula of benzene is C_6H_6.

Element	%	At.mass	%/ At.mass	simplest at.ratio	simplest whole number ratio
C	40	12	$\frac{40}{12} = 3.33$	$\frac{3.33}{3.33} = 1$	1
H	6.66	1	6.66	$\frac{6.66}{3.33} = 2$	2
O	53.34	16	$\frac{53.34}{16} = 3.33$	$\frac{3.33}{3.33} = 1$	1

$\therefore$ Empirical formula = CH_2O

M.F = n x E.F

$$n = \frac{\text{mol.mass}}{\text{E.F.mass}} = \frac{90(\text{given})}{(12+2+16)} = \frac{90}{30} = 3$$

$\therefore$ M.F. = $3x(CH_2O) = C_3H_6O_3$

Question 8.

The mole concept helps in handling a large number of atoms and molecules in stoichiometric calculations. (March – 2013)

a) Define 1 mol.

b) What is the number of hydrogen atoms in 1 mole of methane (CH_4)?

c) Calculate the amount of carbon dioxide formed by the complete combustion of 80g of methane as per the reaction :

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

(Atomic masses: C = 12.01 u.

H = 1.008u,

O = 16u)

Answer:

a) I mole is defined as the amount of any substance which contains Avogadro number of particles (ie atoms, ions or molecules)

b) I mole methane contain 4 Flydrogen atoms (ie n = 4)

No. of hydrogen atoms

= No. of mole x N_A x n

= 1 x 6.023 x 10^{23} x 4

= 24.092 x 10^{23} atoms

Question 9.

a) Atoms have very very small mass and so usually the masses of atoms are given relative to a standard called atomic mass unit. What is the Atomic Mass Unit (AMU)? (Say –

2013)

b) In a reaction $A + B_2 \rightarrow AB_2$, identify the limiting reagent in the reaction mixture containing 5 mol A and 2.5 mol B.

c) Calculate the mass of NaOH required to make 500 mL of 0.5 M aqueous solution (Molecular mass of NaOH = 40).

Answer:

c) $CH_{4(g)} + 2O_2 \rightarrow CO_2 + 2H_2O$

16g

16g of CH_4 gives 44 g of CO_2

$\therefore$ 1 g of CH_4 gives $\frac{44}{16}$ g of CO_2

$\therefore$ 80 g of CH_4 gives $\frac{44}{16}$ x 80 of CO_2

ie 44 x 5 = 220g of CO_2

b) As per the reaction 1 mol of A reacts completely with 1 mole of B_2 to form 1 mole of AB_2. Thus, 5 mole of A requires 5 mole of B_2. Flence, B_2 is the limiting reagent.

a) The mass equal to $1/12^{th}$ the mass of a ^{12}C atom is called one atomic mass unit (amu).

$$1 \text{ atomic mass unit} = \frac{\text{Mass of a } ^{12}C \text{ atom}}{12}$$

$$\frac{1.9924 \times 10^{-23}}{12} g = 1.66 \times 10^{-24} g$$

$$= 1.66 \times 10^{-27} kg$$

c) $Molarity\ (M)\ of\ NaOH\ solution = \dfrac{Mass\ of\ NaOH \times 1000}{Molar\ mass\ of\ NaOH \times Volume\ of\ NaOH\ in\ mL}$

$\therefore Mass\ of\ NaOH =$

$$\frac{Molarity \times M.\ mass\ of\ NaOH \times Vol.\ of\ NaOH\ in\ mL}{1000}$$

$$= \frac{0.5\ M \times 40\ g\ mol^{-1} \times 500\ mL}{1000}$$

= 10 g

Question 10.

a) How many moles of dioxygen are present in 64 g of dioxygen? (Molecular mass of dioxygen is 32).

b) The following data were obtained when dinitrogen (N_2) and dioxygen (O_2) react together to form different compounds.

Mass of N_2	Mass of 0_2
14 g	16g
14 g	32 g
28 g	32 g
28 g	80 g

Name the law of chemical combination obeyed by the above experimental data.

c) Define empirical formula. How is it related to the molecular formula of a compound?

Answer:

b) Law of multiple proportions.

c) Empirical formula is the simplest formula which represents the simplest whole number ratio of various atoms present in a compound.

a) Number of moles = $\dfrac{\text{Mass}}{\text{Molar mass}}$

Number of moles of dioxygen in 64 g of dioxygen

$$= \frac{64\ g}{32\ g\ mol^{-1}} = 2\ mol$$

Molecular formula = n x (Empirical formula) where ‘n’ = 1, 2, 3, ...

Question 11.
Hydrogen combines with oxygen to form two different compounds, namely, water (H_2O) and hydrogen peroxide (H_2O_2) (August – 2014)
a) Which law is obeyed by this combination?
b) State the law
c) How may significant figures are present in the following?
i) 0.0025
ii) 285
Answer:
a) Law of Multiple Proportions
b) It states that if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers.
c) i) 0.0025 – 2 significant figures
ii) 285 – 3 significant figures

Question 12.
'A given compound always contains exactly the same proportion of elements by weight'. (March – 2015)
a) i) Name the above law.
ii) Write the name of the scientist who proposed this law.
b) Calculate the number of molecules in each of the following:
i) 1g N_2
ii) 1g CO_2
(Given that Na is 6.02 x 10^{23}, molecular mass of N_2 is 28 and CO_2 is 44).
Answer:
a) i) Law of Definite Proportions Or Law of Definite Composition
ii) Joseph Proust

b) i) Number of molecules in 1 g N_2

$$= \frac{\text{Mass}}{\text{Gram molecular mass}} \times N_A$$

$$= \frac{1}{28} \times 6.022 \times 10^{23} = 2.15 \times 10^{22}$$

ii) Number of molecules in 1 g CO_2

$$= \frac{\text{Mass}}{\text{Gram molecular mass}} \times N_A$$

$$= \frac{1}{44} \times 6.022 \times 10^{23} = 1.37 \times 10^{22}$$

Question 13.
12g of 12C contains Avogadro's number of carbon atms. (Say – 2015)
a) Give the Avagadro's number.
b) The mass of 2 moles of ammonia gas is
i) 2g
ii) 1.2 x 10^{22}g
iii) 17g
iv) 34g

c) Calculate the volume of ammonia gas produced at STP when 140g of nitrogen gas reacts with 30g of hydrogen gas. (Atomic masses : N = 14u, H = 1u)
Answer:
a) 6.022 x 10^{23}
b) iv) 34 g

Question 14.

a) When nitrogen and hydrogen combines to form ammonia, the ratio between the volumes of gaseous reactants and products is 1:3:2. Name the law of chemical combination illustrated here. (March – 2016)

b) A compound is made up of two elements A and B, hasA= 70%, B = 30%. The relative number of moles of A and B in the compound are 1.25 and 1.88 respectively. If the molecular mass of the compound is 160, find the molecular formula of the compound.

c) N_2 reacts with H_2 as per the equation

$N_2(g)$	+ $3H_2(g)$	$\rightleftharpoons$ $2NH_3(g)$
1 mol	3 mol	2 mol
28 g	6 g	34 g
22.4 mL	67.2 mL	44.8 mL (at STP)

$$\text{Volume of } NH_3(g) = \frac{44.8L \times 140g}{28g} = 224L$$

Answer:

a) Gay Lussac's law of gaseous volumes

b) The empiricial formula of the compound is A_2B_3

$$\text{Atomic mass} = \frac{\text{Mass percentage}}{\text{Relative no. of moles}}$$

$$\text{Atomic mass of A} = \frac{70}{1.25} = 56$$

$$\text{Atomic mass of B} = \frac{30}{1.88} = 15.96 \approx 16$$

$$\text{Empirical formula mass} = (56 \times 2) + (16 \times 3)$$
$$= 112 + 48 = 160$$

$$\text{Molecular mass} = 160 \text{ (given)} \quad \therefore \quad n = \frac{160}{160} = 1$$

$$\text{Molecular formula} = (\text{Empirical formula}) \times n$$
$$= (A_2B_3) \times 1 = A_2B_3$$

Question 15.

The empirical formula represents the simplest whole number ratio of various atoms present in a compound. (Say – 2016)

a) Give the relationship between empirical formula and molecular formula.

b) An organic compound has the following percentage composition C = 12.36%, H = 2.13%, Br = 85%. Its vapour density is 94. Find its molecular formula.

c) What is mole fraction?

Answer:

a) Molecular formula is a whole number multiple of the empirical formula.

i.e., Molecular formula = nx empirical formula where, n = 1, 2, 3...

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

c) Mole fraction of a component in a solution is the ratio of the number of moles of that component to the total number of moles of all the components in solution.

If a substance 'A' dissolves in substance 'B' and their number of moles are nA and nB respectively; then the mole fractions of A and B are given as Mole fraction of A.

Question 16.

a) Determine the number of moles present in 0.55mg of electrons. (March – 2017)

i) 1 mole

ii) 2 moles
iii) 1.5 moles
iv) 0.5 mole

b) Give the empirical formula of the following.
$C_6H_{12}O_6$, C_6H_6, CH_3COOH, $C_6H_6Cl_6$
c) Two elements, carbon and hydrogen combine to form C_2H_6, C_2H_4 and C_2H_2. Identify the law illustrated here.
Answer:
a) i) 1 mol
[Explanation: Mass of 1 electron
= 9.1094 x 10^{-31} kg
= 9.1094 x 10^{-25} mg

Mass of 1 mole of electrons
= 6.022 x 10^{23} X 9.1094 x 10^{-25} mg
= 0.55 mg
∴ Number of moles present in 0.55 mg of

c) Law of multiple propotions [Explanation:
In C_6H_6:24g Carbon + 6g Hydrogen
In C_2H_4:24g Carbon + 4g Hydrogen
In C_2H_2:24g Carbon + 2g Hydrogen
Here, the masses of hydrogen (6g, 4g and 2g) which combine with a fixed mass of carbon (24g) bear a simple ratio, i.e., 6 : 4 : 2 = 3 : 2 : 1]

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b)

Element	%	At. mass	Rel.no. of moles	Simple ratio	Sim. whole no. ratio
Carbon	12.36	12.01	$\frac{12.36}{12.01}=1.03$	$\frac{1.03}{1.03}=1$	1
Hydrogen	2.13	1.008	$\frac{2.13}{1.008}=2.11$	$\frac{2.11}{1.03}=2$	2
Bromine	85	79.90	$\frac{85}{79.9}=1.06$	$\frac{1.06}{1.03}=1$	1

Empirical formula = CH_2Br
Given, Vapour density = 94
Molar mass = 2 × Vapour density = 2 × 94 = 188
Empirical formula mass
= 12.01 + 2 × 1.008 + 79.9
= 93.926

$$n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{188}{93.926} = 2$$

∴ Molecular formula = 2 × (CH_2Br) = $C_2H_4Br_2$

$$= \frac{\text{No. of moles of A}}{\text{No. of moles of solution}}$$

$$= \frac{n_A}{n_A + n_B}$$

Mole fraction of B

$$= \frac{\text{No. of moles of B}}{\text{No. of moles of solution}}$$

$$= \frac{n_B}{n_A + n_B}$$

$$\text{electrons} = \frac{0.55 \text{ mg}}{0.55 \text{ mg mol}^{-1}} = 1 \text{ mol }]$$

b)

MF	EF	
$C_6H_{12}O_6$	CH_2O	(dividing MF by 6)
C_6H_6	CH	(dividing MF by 6)
CH_3COOH	$(C_2H_4O_2)CH_2O$	(dividing MF by 2)
$C_6H_6Cl_6$	CHCl	(dividing MF by 6)