

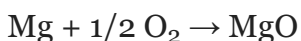
Plus One Chemistry Chapter Wise Important Questions

Chapter 1 Some Basic Concepts of Chemistry

Question 1.

Calculate the number of moles of O_2 required to produce 240g of MgO by burning Mg metal. (Atomic mass: Mg = 24, O = 16) (March – 2019)

Answer:



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 100g of the solution.

b)

$$\text{Mass percentage} = \frac{\text{Mass of solute} \times 100}{\text{Mass of solution}}$$

Element	%	Relative no. of atoms	Dividing by smallest factor	Rate of no. atoms
C	24.27%	$24.27/12=2.02$	$2.02/2.02=1.01$	1
H	4.07%	$4.07/1=4.07$	$4.07/2.02=2.02$	2
Cl	71.65%	$71.65/35.5=2.018$	$2.018/2.018=1$	1

Empirical formula = CH_2Cl

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.

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 CO₂.

b) State and explain the law of conservation of mass.

c) Calculate the molarity of a solution containing 8g of NaOH in 500 mL of water.

Answer:

a) Law of multiple proportions.

b) Matter can neither be created nor be destroyed.

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 form 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.

$$\begin{aligned}\text{Empirical formula mass} &= 12 \times 1 + 1 \times 2 + 1 \times 35.5 \\ &= 12 + 2 + 35.5 = 49.5\end{aligned}$$

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

$$\begin{aligned}\text{Molecular formula of the compound} &= (\text{CH}_2\text{Cl}) \times 2 \\ &= \text{C}_2\text{H}_4\text{Cl}_2\end{aligned}$$

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

Number of moles in 1 litre of water

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

Number of water molecules in 1 litre

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



100g of CaCO₃ on heating given 56 gm of CaO and 44 gram of CO₂.

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

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

OR

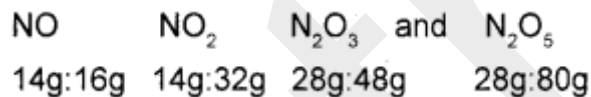
The total mass of the reactants in a chemical reaction is equal to the total mass of the products formed. This law was proposed by Antoine Lavoisier,

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,

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 to 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.

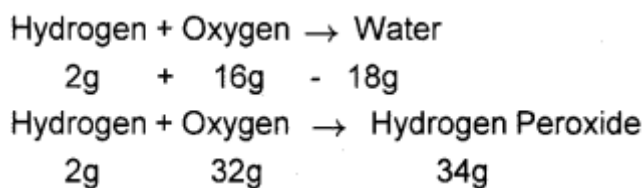
b) What is meant by 'limiting reagent' in a chemical reaction?

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:

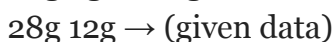
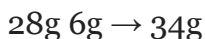
a) Law of multiple proportions

Here, the masses of oxygen (ie, 16g and 32g) which combine with a fixed mass of hydrogen (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 compounds, the weights of one of the elements which combine with a fixed weight of the other in these compounds bear a simple whole-number ratio by weight.

b) The reactant which is completely used up first in a reaction is called the limiting reactant.



∴ N₂ is the limiting reagent. (N₂ 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 oxygen=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

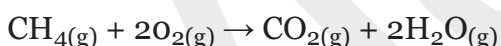
b) 1) Empirical formula of a compound is defined as the simplest formula that gives the ratio of the various elements in a molecules.

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 C₆H₆.

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 (CH₄)?
 c) Calculate the amount of carbon dioxide formed by the complete combustion of 80g of methane as per the reaction :



(Atomic masses: C = 12.01 u,

H = 1.008u,

O = 16u)

Answer:

- a) 1 mole is defined as the amount of any substance which contains Avogadro number of particles (ie atoms, ions or molecules)
 b) 1 mole methane contain 4 Hydrogen atoms (ie n = 4)

$$\begin{aligned} \text{No. of hydrogen atoms} &= \text{No. of mole} \times N_A \times n \\ &= 1 \times 6.023 \times 10^{23} \times 4 \\ &= 24.092 \times 10^{23} \text{ atoms} \end{aligned}$$

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 –

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

∴ Empirical formula = CH₂O

$$\text{M.F} = n \times \text{E.F}$$

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

∴ M.F. = 3x(CH₂O) = C₃H₆O₃

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:

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 . Hence, B_2 is the limiting reagent.

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 O_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.

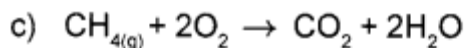
c) Define empirical formula. How is it related to the molecular formula of a compound?

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.

Molecular formula = n x (Empirical formula) where 'n' = 1, 2, 3,



16g

16g of CH_4 gives 44 g of CO_2

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

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

$$\text{ie } 44 \times 5 = 220\text{g of } CO_2$$

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

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

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

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

c) Molarity (M) of NaOH solution =

$$\frac{\text{Mass of NaOH} \times 1000}{\text{Molar mass of NaOH} \times \text{Volume of NaOH in mL}}$$

$$\therefore \text{Mass of NaOH} =$$

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

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

$$= 10 \text{ g}$$

a) Number of moles = $\frac{\text{Mass}}{\text{Molar mass}}$
Number of moles of dioxygen in 64 g of dioxygen

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

Question 11.

Hydrogen combines with oxygen to form two different compounds, namely, water (H₂O) and hydrogen peroxide (H₂O₂) (August – 2014)

- a) Which law is obeyed by this combination?
- b) State the law
- c) How many 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 N₂
 - ii) 1g CO₂(Given that N_A is 6.02 × 10²³, molecular mass of N₂ is 28 and CO₂ is 44).

Answer:

- a) i) Law of Definite Proportions Or Law of Definite Composition
ii) Joseph Proust

Question 13.

12g of ¹²C contains Avogadro’s number of carbon atoms. (Say – 2015)

- a) Give the Avogadro’s number.
- b) The mass of 2 moles of ammonia gas is
 - i) 2g
 - ii) 1.2 × 10²²g
 - 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 × 10²³
- b) iv) 34 g

b) i) Number of molecules in 1 g N₂

$$= \frac{\text{Mass}}{\text{Gram molecular mass}} \times N_A$$
$$= \frac{1}{28} \times 6.022 \times 10^{23} = 2.15 \times 10^{22}$$

ii) Number of molecules in 1 g CO₂

$$= \frac{\text{Mass}}{\text{Gram molecular mass}} \times N_A$$
$$= \frac{1}{44} \times 6.022 \times 10^{23} = 1.37 \times 10^{22}$$

Question 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, has A = 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.

Answer:

- a) Gay Lussac's law of gaseous volumes
b) The empirical formula of the compound is 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 = $n \times$ empirical formula where, $n = 1, 2, 3, \dots$

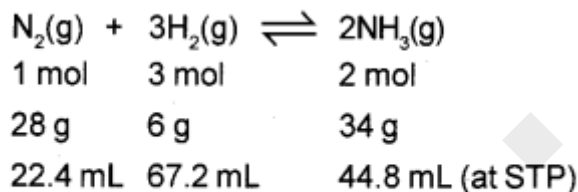
c) Mole fraction of a component in a solution is the ratio of the number of moles of that component to the total number of moles of all the components in solution.

If a substance 'A' dissolves in substance 'B' and their number of moles are n_A and n_B 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

c) N_2 reacts with H_2 as per the equation



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

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

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

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

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

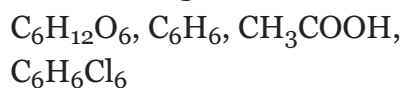
$$\text{Molecular mass} = 160 \text{ (given)} \therefore n = \frac{160}{160} = 1$$

$$\begin{aligned} \text{Molecular formula} &= (\text{Empirical formula}) \times n \\ &= (A_2B_3) \times 1 = A_2B_3 \end{aligned}$$

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

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

b) Give the empirical formula of the following.



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
 $= 9.1094 \times 10^{-31}$ kg
 $= 9.1094 \times 10^{-25}$ mg

Mass of 1 mole of electrons
 $= 6.022 \times 10^{23} \times 9.1094 \times 10^{-25}$ mg
 $= 0.55$ mg
 \therefore Number of moles present in 0.55 mg of

c) Law of multiple proportions [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 = $2 \times$ Vapour density = $2 \times 94 = 188$

Empirical formula mass
 $= 12.01 + 2 \times 1.008 + 79.9$
 $= 93.926$

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

\therefore Molecular formula = $2 \times (CH_2Br) = C_2H_4Br_2$

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

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

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)