11th Chemistry – Lesson 1

Creative – Study Material

Question 1. Calculate the molecular mass of the following:
(i) H_20(ii) $C0_2$(iii) CH_4

Answer: (i) Molecular mass of H_2O = 2(1.008 amu) + 16.00 amu=18.016 amu
(ii) Molecular mass of CO_2= 12.01 amu + 2 x 16.00 amu = 44.01 amu
(iii) Molecular mass of CH_4= 12.01 amu + 4 (1.008 amu) = 16.042 amu

Question 2. Calculate the mass percent of different elements present in sodium sulphate ($Na_2\ SO_4$).
Answer:

$$\text{Mass \% of an element} = \frac{\text{Mass of that element in the compound}}{\text{Molar mass of the compound}} \times 100$$

Now, Molar mass of Na_2SO_4 $= 2\ (23.0) + 32.0 + 4 \times 16.0$
$= 142\ \text{g mol}^{-1}$

$$\text{Mass percent of sodium} = \frac{46}{142} \times 100 = 32.39\ \%$$

$$\text{Mass percent of sulphur} = \frac{32}{142} \times 100 = 22.54\ \%$$

$$\text{Mass percent of oxygen} = \frac{64}{142} \times 100 = 45.07\ \%$$

Question 3. Determine the empirical formula of an oxide of Iron which has 69.9 % iron and 30.1 % dioxygen by mass.

Answer:

Element	Symbol	% by mass	Atomic mass	Moles of the element (Relative no. of moles)	Simplest molar ratio	Simplest whole number molar ratio
Iron	Fe	69.9	55.85	$\frac{69.9}{55.85} = 1.25$	$\frac{1.25}{1.25} = 1$	2
Oxygen	O	30.1	16.00	$\frac{30.1}{16.00} = 1.88$	$\frac{1.88}{1.25} = 1.5$	3

$\therefore$ Empirical formula = $\mathbf{Fe_2O_3}$.

Question 4. Calculate the amount of carbon dioxide that could be produced when
(i) 1 mole of carbon is burnt in air.
(ii) 1 mole of carbon is burnt in 16 g of dioxygen.
(iii) 2 moles of carbon are burnt in 16 g of dioxygen.

Answer: The balanced equation for the combustion of carbon in dioxygen/air is

$$\underset{1\text{ mole}}{C(s)} + \underset{\substack{1\text{ mole}\\(32\ g)}}{O_2(g)} \longrightarrow \underset{\substack{1\text{ mole}\\(44\ g)}}{CO_2(g)}$$

(i) In air, combustion is complete. Therefore,CO_2 produced from the combustion of 1 mole of carbon = 44 g.(ii) As only 16 g of dioxygen is available, it can combine only with 0.5 mole of carbon, i.e., dioxygen is the limiting reactant. Hence,CO_2 produced = 22 g.(iii) Here again, dioxygen is the limiting reactant. 16 g of dioxygen can combine only with 0.5 mole of carbon.CO_2 produced again is equal to 22 g.

Question 5. Calculate the mass of sodium acetate (CH_3COONa) required to make 500 mL of 0.375 molar aqueous solution. Molar mass of sodium acetate is 82.0245 g mol^{-1}

Answer: 0.375 M aqueous solution means that 1000 mL of the solution contain sodium acetate = 0.375 mole

$\therefore$ 500 mL of the solution should contain sodium acetate = $\frac{0.375}{2}$ mole

Molar mass of sodium acetate = 82.0245 g mol^{-1}

$\therefore$ Mass of sodium acetate required = $\frac{0.375}{2}$ mole,× 82.0245 g mol^{-1} = **15.380 g.**

Question 6. Calculate the concentration of nitric acid in moles per litre in a sample which has a density 1.41 g mL^{-1} and the mass percent of nitric acid in it is being 69%.

Answer: Mass percent of 69% means that 100 g of nitric acid solution contain 69 g of nitric acid by mass.

Molar mass of nitric acid HNO_3= 1 + 14 + 48 = 63 gmol^{-1}

$$\therefore \quad \text{Moles in 69 g } HNO_3 = \frac{69\text{ g}}{63\text{ g mol}^{-1}} = 1.095\text{ mole}$$

$$\text{Volume of 100 g nitric acid solution} = \frac{100\text{ g}}{1.41\text{ g mL}^{-1}} = 70.92\text{ mL} = 0.07092\text{ L}$$

$$\therefore \quad \text{Conc. of } HNO_3 \text{ in moles per litre} = \frac{1.095\text{ mole}}{0.07092\text{ L}} = \mathbf{15.44\ M.}$$

Question 7. How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$)? (Atomic mass of Cu= 63.5 amu)

Answer: 1 mole of $CuS0_4$ contains 1 mole (1 g atom) of Cu

Molar mass of $CuS0_4$= 63.5 + 32 + 4 x 16 = 159.5 g mol^{-1}

Thus, Cu that can be obtained from 159.5 g of $CuS0_4$ = 63.5 g

$$\therefore \quad \text{Cu that can be obtained from 100 g of } CuSO_4 = \frac{63.5}{159.5} \times 100 \text{ g} = \mathbf{39.81\ g.}$$

Question 8. Determine the molecular formula of an oxide of iron in which the mass percent of iron and oxygen are 69.9 and 30.1 respectively. Given that the molar mass of the oxide is 159.8 g mol^{-1}(Atomic mass: Fe = 55.85, O = 16.00 amu)Calculation of Empirical Formula. See Q3.

Answer: Empirical formula mass of Fe_20_3 = 2 x 55.85 + 3 x 16.00 = 159.7 g mol^{-1}

$$n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{159.8}{159.7} = 1$$

Hence, molecular formula is same as empirical formula, viz.,Fe_20_3.

Question 9.Calculate the atomic mass (average) of chlorine using the following data:

	% Natural Abundance	*Molar Mass*
^{35}Cl	*75.77*	*34.9689*
^{37}Cl	*24.23*	*36.9659*

Answer:

Fractional abundance of ^{35}Cl = 0.7577, Molar mass = 34.9689

Fractional abundance of ^{37}Cl = 0.2423, Molar mass = 36.9659

$$\therefore \quad \text{Average atomic mass} = (0.7577)(34.9689\text{ amu}) + (0.2423)(36.9659\text{ amu})$$
$$= 26.4959 + 8.9568 = \mathbf{35.4527}$$

Question 10.In three moles of ethane (C_2H_6), calculate the following:
(i) Number of moles of carbon atoms (ii) Number of moles of hydrogen atoms
(iii) Number of molecules of ethane

Answer: (i) 1 mole of C_2H_6 contains 2 moles of carbon atoms
∴ 3 moles of C_2H_6 will C-atoms = 6 moles
(ii) 1 mole of C_2H_6 contains 6 moles of hydrogen atoms
∴ 3 moles of C_2H_6 will contain H-atoms = 18 moles

(*iii*) 1 mole of C_2H_6 contains Avogadro's no., *i.e.*, 6.02×10^{23} molecules
∴ 3 moles of C_2H_6 will contain ethane molecules $= 3 \times 6.02 \times 10^{23}$
$= \mathbf{18.06 \times 10^{23}}$ **molecules**

Question 11. What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a final volume up to 2 L?

Answer:
Molar mass of sugar $(C_{12}H_{22}O_{11}) = 12 \times 12 + 22 \times 1 + 11 \times 16 = 342 \text{ g mol}^{-1}$

$$\text{No. of moles in 20 g of sugar} = \frac{20 \text{ g}}{342 \text{g mol}^{-1}} = 0.0585 \text{ mole}$$

$$\text{Molar concentration} = \frac{\text{Moles of solute}}{\text{Volume of sol in L}} = \frac{0.0585}{2 \text{ L}} = 0.0293 \text{ mol L}^{-1} = \mathbf{0.0293\ M.}$$

Question 12. If the density of methanol is 0.793 kg L^{-1}, what is its volume needed for making 2.5 L of its 0.25 M solution?

Answer:
Molar mass of methanol $(CH_3OH) = 32 \text{ g mol}^{-1} = 0.032 \text{ kg mol}^{-1}$

$$\text{Molarity of the given solution} = \frac{0.793 \text{ kg L}^{-1}}{0.032 \text{ kg mol}^{-1}} = 24.78 \text{ mol L}^{-1}$$

Applying $M_1 \times V_1 = M_2V_2$
(Given solution) (Solution to be prepared)

$$24.78 \times V_1 = 0.25 \times 2.5 \text{ L or } V_1 = 0.02522 \text{ L} = \mathbf{25.22\ mL}$$

Question 13. Pressure is determined as force per unit area of the surface. The S.I. unit of pressure, pascal, is as shown below:1 Pa = 1 Nm^{-2}.If mass of air at sea level is 1034 g cm^{-2},calculate the pressure in pascal.

Answer: Pressure is the force (i.e., weight) acting per unit area But weight = mg

$$\therefore \quad \text{Pressure} = \text{Weight per unit area} = \frac{1034 \text{ g} \times 9.8 \text{ m s}^{-2}}{\text{cm}^2}$$

$$= \frac{1034 \text{ g} \times 9.8 \text{ ms}^{-2}}{\text{cm}^2} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \frac{100 \text{ cm}}{1 \text{ m}} \times \frac{100 \text{ cm}}{1 \text{ m}} \times 1 \times \frac{1 \text{ N}}{\text{kg ms}^{-2}} \times \frac{1 \text{ Pa}}{1 \text{ Nm}^{-2}}$$

$= \mathbf{1.01332 \times 10^5}$ **Pa.**

Question 14. What is the S.I. unit of mass?
Answer: S.I. unit of mass is kilogram (kg).

Question 15. Match the following prefixes with their multiples:

Prefixes	*Multiples*
(i) micro	10^6
(ii) deca	10^9
(iii) mega	10^{-6}
(iv) giga	10^{-15}
(v) femto	10

Answer:

micro = 10^{-6}, deca = 10, mega = 10^6, giga = 10^9, femto = 10^{-15}.

Question 16. What do you mean by significant figures?
Answer: The digits in a properly recorded measurement are known as significant figures. It is also defined as follows. The total numbers of figures in a number including the last digit whose value is uncertain is called number of significant figures.

Question 17. A sample of drinking water was found to be severely contaminated with chloroform, CHCly supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).
(i) Express this in percent by mass
(ii) Determine the molality of chloroform in the water sample.
Answer:

(*i*) 15 ppm means 15 parts in million (10^6) parts

$$\therefore \quad \% \text{ by mass} = \frac{15}{10^6} \times 100 = 15 \times 10^{-4} = \mathbf{1.5 \times 10^{-3}\ \%}$$

(*ii*) Molar mass of chloroform ($CHCl_3$) = 12 + 1 + 3 × 35.5 = 119.5 g mol^{-1}
100 g of the sample contain chloroform = 1.5×10^{-3} g
$\therefore$ 1000 g (1 kg) of the sample will contain chloroform = 1.5×10^{-2} g

$$= \frac{1.5 \times 10^{-2}}{119.5} = 1.26 \times 10^{-4} \text{ mole}$$

$\therefore$ Molality = $\mathbf{1.266 \times 10^{-4}}$ **m.**

Question 18. Express the following in scientific notation:
(i) 0.0048 (v) 6.0012 (ii) 234,000 (iii) 8008 (iv) 500.0
Answer:
(i) 4.8×10^{-3} (ii) 2.34×10^{5} (iii) 8.008×10^{3} (iv) 5.000×10^{2}
(v) 6.0012×10^{0}

Question 19. How many significant figures are present in the following?
(i) 0.0025 (ii) 208 (iii) 5005 (iv) 126,000
(v) 500.0 (vi) 2.0034
Answer: (i) 2 (ii) 3 (iii) 4 (iv) 3 (v) 4 (vi) 5.

Question 20. Round up the following upto three significant figures:
(i) 34.216 (ii) 10.4107 (iii) 0.04597 (iv) 2808
Answer: (i) 34.2 (ii) 10.4 (iii) 0.0460 (iv) 2810

Question 21. The following data were obtained when dinitrogen and dioxygen react together to form compounds:

	Mass of dinitrogen	Mass of dioxygen
(i)	14 g	16 g
(ii)	14 g	32 g
(iii)	28 g	32 g
(iv)	28 g	80 g

(a) Which law of chemical combination is obeyed by the above experimental data? Give its statement.
(b) Fill in the blanks in the following conversions:
(i) 1 km = ... mm = ... pm (ii) 1 mg = ... kg = ...ng
(iii) 1 mL = ... L = ... dm^3

Answer: (a) Fixing the mass of dinitrogen as 28 g, masses of dioxygen combined will be 32,64, 32 and 80 g in the given four oxides. These'are in the ratio 1 : 2 : 1 : 5 which is a simple whole number ratio. Hence, the given data obey the law of multiple proportions.

(b) (i) $1\text{ km} = 1\text{ km} \times \frac{1000\text{ m}}{1\text{ km}} \times \frac{100\text{ cm}}{1\text{ m}} \times \frac{10\text{ mm}}{1\text{ cm}} = \mathbf{10^6\text{ mm}}$

$$1\text{ km} = 1\text{ km} \times \frac{1000\text{ m}}{1\text{ km}} \times \frac{1\text{ pm}}{10^{-12}\text{ m}} = \mathbf{10^{15}\text{ pm}}$$

(ii) $1\text{ mg} = 1\text{ mg} \times \frac{1\text{ g}}{1000\text{ mg}} \times \frac{1\text{ kg}}{1000\text{ g}} = \mathbf{10^{-6}\text{ kg}}$

$$1\text{ mg} = 1\text{ mg} \times \frac{1\text{ g}}{1000\text{ mg}} \times \frac{1\text{ ng}}{10^{-9}\text{ g}} = \mathbf{10^{6}\text{ ng}}$$

(iii) $1\text{ mL} = 1\text{ mL} \times \frac{1\text{ L}}{1000\text{ mL}} = \mathbf{10^{-3}\text{ L}}$

$$1\text{ mL} = 1\text{ cm}^3 = 1\text{ cm}^3 \times \frac{1\text{ dm} \times 1\text{ dm} \times 1\text{ dm}}{10\text{ cm} \times 10\text{ cm} \times 10\text{ cm}} = 10^{-3}\text{ dm}^3.$$

Question 22.

If the speed of light is $3.0 \times 10^8\text{ ms}^{-1}$, calculate the distance covered by light in 2.00 ns.

Answer:

Distance covered = Speed × Time = $3.0 \times 10^8\text{ ms}^{-1} \times 2.00\text{ ns}$

$$= 3.0 \times 10^8\text{ ms}^{-1} \times 2.00\text{ ns} \times \frac{10^{-9}\text{ s}}{1\text{ ns}} = 6.00 \times 10^{-1}\text{ m} = \mathbf{0.600\text{ m}}$$

Question 23. In the reaction, A + B2——> AB2, identify the limiting reagent, if any, in the following mixtures
(i) 300 atoms of A + 200 molecules ofB
(ii) 2 mol A + 3 mol B
(iii) 100 atoms of A + 100 molecules ofB
(iv) 5 mol A + 2.5 mol B
(v) 2.5 mol A + 5 mol B

Answer: (i) According to the given reaction, 1 atom of A reacts with 1 molecule of B
.•. 200 molecules of B will react with 200 atoms of A and 100 atoms of A will be left unreacted. Hence, B is the limiting reagent while A is the excess reagent.
(ii) According to the given reaction, 1 mol of A reacts with 1 mol of B
.•. 2 mol of A will react with 2 mol of B. Hence, A is the limiting reactant.
(iii) No limiting reagent.
(iv) 2.5 mol of B will react with 2.5 mol of A. Hence, B is the limiting reagent.
(v) 2.5 mol of A will react with 2.5 mol of B. Hence, A is the limiting reagent.

Question 24. Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation:(i) N_2 (g) + $3H_2$(g) ——> 2NH3 (g)
(ii) Will any of the two reactants remain unreacted?
(iii) If yes, which one and what would be its mass?
Answer:

(*i*) 1 mol of N_2 *i.e.,* 28 g react with 3 mol of H_2, *i.e.,* 6 g of H_2

$\therefore$ 2000 g of N_2 will react with $H_2 = \frac{6}{28} \times 200$ g = 428.6 g. Thus, N_2 is the limiting reagent while H_2 is the excess reagent.

2 mol of N_2, *i.e.,* 28 g of N_2 produce NH_3 = 2 mol = 34 g

$\therefore$ 2000 g of N_2 will produce $NH_3 = \frac{34}{28} \times 2000$ g = **2428.57 g**

(*ii*) H_2 will remain unreacted.

(*iii*) Mass left unreacted = 1000 g – 428.6 g = 571.4 g

Question 25. How are 0.50 mol Na_2C0_3 and 0.50 M Na_2C0_3 different?
Answer: Molar mass of Na_2C0_3= 2 x 23 + 12 + 3 x 16 = 106g mol^{-1} 0.50 mol Na_2C0_3 means 0.50 x 105 g = 53 g 0. 50 M Na2C03 means 0.50 mol, i.e., 53 g Na_2C0_3 are present in 1 litre of the solution.

Question 26. If ten volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour could be produced?
Answer: H_2 and 0_2 react according to the equation

$2H_2$(g) + 0_2 (g) ——>$2H_2O$ (g)

Thus, 2 volumes of H_2 react with 1 volume of 0_2 to produce 2 volumes of water vapour. Hence, 10 volumes of H_2 will react completely with 5 volumes of 0_2 to produce 10 volumes of water vapour.

Question 27. Convert the following into basic units:
(i) 28.7 pm (ii) 15.15 μs (iii) 25365 mg
Answer:

$$(i)\ 28.7\text{ pm} = 28.7\text{ pm} \times \frac{10^{-12}\text{ m}}{1\text{ pm}} = \mathbf{2.87 \times 10^{-11}\ m}$$

$$(ii)\ 15.15\ \mu s = 15.15\ \mu s \times \frac{10^{-6}\text{ s}}{1\ \mu s} = \mathbf{1.515 \times 10^{-5}\ s}$$

$$(iii)\ 25365\text{ mg} = 25365\text{ mg} \times \frac{1\text{ g}}{1000\text{ mg}} \times \frac{1\text{ kg}}{1000\text{ g}} = \mathbf{2.5365 \times 10^{-2}\ kg}$$

Question 28. Which one of the following will have largest number of atoms?
(i) 1 g Au (s) (ii) 1 g Na (s) (iii) 1 g Li (s) (iv) 1 g of Cl_2(g) (Atomic masses: Au = 197, Na = 23, Li = 7, Cl = 35.5 amu)

Answer:

(*i*) 1 g Au $= \frac{1}{197}$ mol $= \frac{1}{197} \times 6.02 \times 10^{23}$ atoms

(*ii*) 1 g Na $= \frac{1}{23}$ mol $= \frac{1}{23} \times 6.02 \times 10^{23}$ atoms

(*iii*) 1 g Li $= \frac{1}{7}$ mol $= \frac{1}{7} \times 6.02 \times 10^{23}$ atoms

(*iv*) 1 g $Cl_2 = \frac{1}{71}$ mol $= \frac{1}{71} \times 6.02 \times 10^{23}$ molecules $= \frac{2}{71} \times 6.02 \times 10^{23}$ atoms

Thus, **1 g of Li** has the largest number of atoms.

Question 29. Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040.

Answer:

$$x_{C_2H_5OH} = \frac{n(C_2H_5OH)}{n(C_2H_5OH) + n(H_2O)} = 0.040 \text{ (Given)} \quad ...(i)$$

The aim is to find number of moles of ethanol in 1 L of the solution which is nearly = 1 L of water (because solution is dilute)

No. of moles in 1 L of water $= \frac{1000 \text{ g}}{18 \text{ g mol}^{-1}} = 55.55$ moles

Substituting $n(H_2O) = 55.55$ in eqn (*i*), we get

$$\frac{n(C_2H_5OH)}{n(C_2H_5OH) + 55.55} = 0.040$$

or $0.96\, n(C_2H_5OH) = 55.55 \times 0.040$ or $n(C_2H_5OH) = 2.31$ mol

Hence, molarity of the solution = **2.31 M.**

Question 30.

What will be the mass of one ^{12}C atom in g?

Answer:

1 mol of ^{12}C atoms $= 6.022 \times 10^{23}$ atoms = 12 g

Thus, 6.022×10^{23} atoms of ^{12}C have mass = 12g

$\therefore$ 1 atom of ^{12}C will have mass $= \frac{12}{6.022 \times 10^{23}}$ g $= \mathbf{1.9927 \times 10^{-23}}$ **g**

Question 31. How many significant figures should be present in the answer of the following?

(i) $\frac{0.02856 \times 298.15 \times 0.112}{0.5785}$ (ii) 5×5.364 (iii) $0.0125 + 0.7864 + 0.0215$

Answer: (i) The least precise term has 3 significant figures (i.e., in 0.112). Hence, the answer should have 3 significant figures.
(ii) Leaving the exact number (5), the second term has 4 significant figures. Hence, the answer should have 4 significant figures.
(iii) In the given addition, the least number of decimal places in the term is 4. Hence, the answer should have 4 significant.

Question 32. Use the data given in the following table to calculate the molar mass of naturally occurring argon.

Isotope	*Isotopic molar mass*	*Abundance*
^{36}Ar	$35.96755\ g\ mol^{-1}$	0.337
^{38}Ar	$37.96272\ g\ mol^{-1}$	0.063
^{40}Ar	$39.9624\ g\ mol^{-1}$	99.600

Answer: Molar mass of Ar = 35.96755 x 0.00337 + 37.96272 x 0.00063 + 39.96924 x 0.99600 = 39.948 g mol^{-1}

Question 33. Calculate the number of atoms in each of the following:
(i) 52 moles of He (ii) 52 u of He (iii) 52 g of He

Answer:

(i) 1 mol of He = 6.022×10^{23} atoms

$\therefore$ 52 mol of He = $52 \times 6.022 \times 10^{23}$ atoms = $\mathbf{3.131 \times 10^{25}}$ **atoms**

(ii) 1 atom of He = 4 u of He

4 u of He = 1 atom of He

$\therefore$ 52 u of He = $\frac{1}{4} \times 52$ atoms = **13 atoms**

(iii) 1 mole of He = 4 g = 6.022×10^{23} atoms

$$\therefore \quad 52\text{ g of He} = \frac{6.022 \times 10^{23}}{4} \times 52 \text{ atoms}$$

$$= \mathbf{7.8286 \times 10^{24}\ atoms.}$$

Question 34. A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at S.T.P.) of this welding gas is found to weigh 11.6 g. Calculate (i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.

Answer:

Amount of carbon in 3.38 g CO_2 = $\frac{12}{44} \times 3.38\text{ g} = 0.9218\text{ g}$

Amount of hydrogen in 0.690 g H_2O = $\frac{2}{18} \times 0.690\text{ g} = 0.0767\text{ g}$

As compound contains only C and H, therefore, total mass of the compound = 0.9218 + 0.0767 g = 0.9985 g

% of C in the compound = $\frac{0.9218}{0.9985} \times 100 = 92.32$

% of H in the cmpound = $\frac{0.0767}{0.9985} \times 100 = 7.68$

Calculation of Empirical Formula

Element	*% by mass*	*Atomic mass*	*Moles of the element*	*Simplest molar ratio*	*Simplest whole no. molar ratio*
C	92.32	12	$\frac{92.32}{12} = 7.69$	1	1
H	7.68	1	$\frac{7.68}{1} = 7.68$	1	1

∴ Empirical formula = **CH**

10.0 L of the gas at STP weight = 11.6 g

∴ 22.4 L of the gas at S.T.P will weight = $\frac{11.6}{10.0} \times 22.4 = 25.984\text{ g} \approx 26\text{ g}$

∴ Molar mass = **26 g mol^{-1}**

Empirical formula mass of CH = 12 + 1 = 13

∴ $n = \frac{\text{Molecular mass}}{\text{E.F. mass}} = \frac{26}{13} = 2$ ∴ Molecular formula = 2 × CH = $\mathbf{C_2H_2}$

Question 35. Calcium carbonate reacts with aqueous HCl according to the reaction

$CaCO_3$ (s) + 2HCl (aq) ———>$CaCl_2$ (aq) +CO_2(g) +H_2O(l).

What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl?

Answer: **Step 1.** To calculate mass of HCl in 25 mL of 0.75 m HCl

1000 mL of 0.75 M HCl contain HCl = 0.75 mol = 0.75 x 36.5 g = 24.375 g

∴ 25 mL of 0.75 HCl will contain HCl = $\frac{24.375}{1000} \times 25\text{ g} = 0.6844\text{ g}$.

Step 2. To calculate mass of $CaCO_3$ reacting completely with 0.9125 g of HCl

$CaCO_3$ (s) + 2HC1 (aq)———>$CaCl_2$(aq) +CO_2(g) + H_2O

2 mol of HCl, i.e., 2 x 36.5 g = 73 g HCl react completely with $CaC0_3$ = 1 mol = 100 g

$\therefore$ 0.6844 g HCl will react completely with $CaCO_3 = \frac{100}{73} \times 0.6844$ g = **0.938 g.**

Question 36. Chlorine is prepared in the laboratory by treating manganese dioxide ($Mn0_2$) with aqueous hydrochloric acid according to the reaction.
4 HCl (aq) + $Mn0_2$ (s) ———> 2 H_2O (l) + $MnCl_2$(aq) +Cl_2(g)

How many grams of HCl react with 5.0 g of manganese dioxide? (Atomic mass of Mn = 55 u)
Answer: 1 mole of $Mn0_2$, i.e., 55 + 32 = 87 g $Mn0_2$ react with 4 moles of HCl, i.e., 4 x 36.5 g = 146 g of HCl.

$\therefore$ 5.0 g of MnO_2 will react with HCl = $\frac{146}{87} \times 5.0$ g = **8.40 g**

MORE QUESTIONS SOLVED

I.Very Short Answer Type Questions

Question 1. What is the SI unit of molarity?
Answer: SI unit of molarity = mol dm^{-3}

Question 2. What do you understand by stoichiometric coefficients in a chemical equation?
Answer: The coefficients of reactant and product involved in a chemical equation represented by the balanced form, are known as stoichiometric coefficients.
For example, N_2(g) + $3H_2$(g) ———> 2 NH_3(g)
The stoichiometric coefficients are 1, 3 and 2 respectively.

Question 3. Give an example of a molecule in which the ratio of the molecular formula is six times the empirical formula.
Answer: The compound is glucose. Its molecular formula is $C_6H_{12}O_6$, while empirical formula is CH_2O.

Question 4. What is an atom according to Dalton's atomic theory?
Answer: According to Dalton's atomic theory, an atom is the ultimate particle of matter which cannot be further divided.

Question 5. Why air is not always regarded as homogeneous mixture?
Answer: This is due to the presence of dust particles.

Question 6. Define the term 'unit' of measurement.
Answer: It is defined as the standard of reference chosen to measure a physical quantity.

Question 7. Define law of conservation of mass.
Answer: It states that matter can neither be created nor destroyed.

Question 8. How is empirical formula of a compound related to its molecular formula?
Answer: Molecular formula = (Empirical formula)n where n is positive integer.

Question 9. How many oxygen atoms are there in 18 g of water?
Answer: Molar mass of water is 18 g/mol.
Number of oxygen atoms is 18 g of water = 6.02×10^{23}

Question 10. Name two factors that introduce uncertainty into measured figures.
Answer: (i) Reliability of measuring instrument.
(ii) Skill of the person making the measurement.

Question 11. State Avogadro's law.
Answer: Equal volumes of all gases under the conditions of same temperature and pressure contain the same number of molecules.

Question 12. How are 0.5 ml of NaOH differents from 0.5 M of NaOH?
Answer: 0.5 ml of NaOH means 0.5 mole (20.0 g) of NaOH, 0.5M of NaOH means that 0.5 mole (20.0g) of NaOH are dissolved in 1L of its solution.

Question 13. What is one a.m.u. or one 'u'?
Answer: 1 a.m.u. or 1 u = 1/12 th mass of an atom of carbon 12.

Question 14. What is the number of significant figures in 1.050×10^4?
Answer: Four.

II. Short Answer Type Questions

Question 1. Define molality. How does molality depend on temperature?
Answer: Molality is defined as the moles of solute per kilogram of solvent.

$$\text{Molality} = m = \frac{\text{Moles of solute}}{\text{Mass of solvent (in kg)}}$$

Molality of a solution does not depend on temperature.

Question 2. Convert 2.6 minutes in seconds.
Answer: We know that, 1 min = 60 s
Conversion factor =60 s/(1min)
2.6 min = 2.6 min x conversion factor = 2.6 x 60s/1min= 156 s.

Question 3. Express the following up to four significant figures.
(i) 6.5089 (ii) 32.3928 (iii) 8.721×10^4 (iv) 2000

Answer:
(i) 6.509 (ii) 32.39 (iii) 8.721×10^4 (iv) 2.000×10^3

Question 4. Calculate the number of moles in each of the following.

Answer:

(i) 392 g of sulphuric acid
(ii) 44.8 litres of sulphur dioxide at N.T.P.
(iii) 6.022×10^{22} molecules of oxygen
(iv) 8g of calcium

(i) 392 g of sulphuric acid

Molar mass of $H_2SO_4 = 2 \times 1 + 32 + 4 \times 16 = 98$ g

98 g of sulphuric acid = 1 mol

$$392 \text{ g of sulphuric acid} = 1 \text{ mol} \times \frac{392 \text{ g}}{(98 \text{ g})} = \mathbf{4\ mol}$$

(ii) 44.8 litres of sulphur dioxide at N.T.P.

22.4 litres of sulphur dioxide at N.T.P. = 1 mol

$$44.8 \text{ litres of sulphur dioxide at N.T.P.} = \frac{1 \text{ mol}}{(22.4\text{L})} \times (44.8 \text{ L}) = \mathbf{2.0\ mol}$$

(iii) 6.022×10^{22} molecules of oxygen

6.022×10^{23} molecules of oxygen = 1 mol

$$6.022 \times 10^{22} \text{ molecules of oxygen} = 1 \text{ mol} \times \frac{6.022 \times 10^{22}}{6.022 \times 10^{23}} = \mathbf{0.1\ mol}$$

(iv) 8g of calcium

Gram atomic mass of Ca = 40 g

40 g of calcium = 1 mol

$$8.0 \text{ g of calcium} = 1 \text{ mol} \times \frac{(8.0 \text{ g})}{(40 \text{ g})} = \mathbf{0.2\ mol.}$$

Question 5. A compound on analysis was found to contain C = 34.6%, H = 3.85% and O = 61.55%. Calculate the empirical formula.

Answer: **Step I.** Calculation of simplest whole number ratios of the elements.

Element	Percentage	Atomic Mass	Gram atoms (Moles)	Atomic ratio (Molar ratio)	Simplest whole no. ratio
C	34.6	12	$\frac{34.6}{12} = 2.88$	$\frac{2.88}{2.88} = 1$	3
H	3.85	1	$\frac{3.85}{1} = 3.85$	$\frac{3.85}{2.88} = 1.337 \text{ or } \frac{4}{3}$	4
O	61.55	16	$\frac{61.55}{16} = 3.85$	$\frac{3.85}{2.88} = 1.337 \text{ or } \frac{4}{3}$	4

The simplest whole number ratios of the different elements are: C:H:O::3:4:4

Step II. Writing the empirical formula of the compound.
The empirical formula of the compound = $C_3H_4O_4$.

Question 6.Calculate:
(a) Mass of 2.5 gram atoms of magnesium,
(b) Gram atom in 1.4 grams of nitrogen (Atomic mass Mg = 24, N = 14)
Answer: (a) 1 gram atom of Mg = 24g
2.5 gram atoms of Mg = 24 x 2.5 = 60g
(b) 1 gram atom of N = 14g;
14g of N = 1 gram atom 1
1.4g of N = 1/14 x 1.4 = 0.1 gram atom.

Question 7. The density of water at room temperature is 1.0 g/mL. How many molecules are there in a drop of water if its volume is 0.05 mL?
Answer:
Volume of a drop of water = 0.05 mL
Mass of a drop of water = Volume × density
= (0.05 mL) × (1.0 g/mL) = 0.05 g
Gram molecular mass of water (H_2O) = 2 × 1 + 16 = 18 g
18 g of water = 1 mol

$$0.05 \text{ g of water} = \frac{1 \text{ mol}}{(18 \text{ g})} \times (0.05 \text{ g}) = 0.0028 \text{ mol}$$

No. of molecules present
1 mole of water contain molecules = 6.022×10^{23}
0.0028 mole of water contain molecules = $6.022 \times 10^{23} \times 0.0028 = \mathbf{1.68 \times 10^{21}}$ **molecules**.

Question 8.What is the molecular mass of a substance each molecule of which contains 9 atoms of carbon, 13 atoms of hydrogen and 2.33×10^{-23} g other component?

Answer:
Mass of 9 atoms of carbon = 9 × 12 amu = 108 u.
Mass of 13 atoms of hydrogen = 13 × 1 amu = 13 u

$$\text{Mass of } 2.33 \times 10^{-23}\text{g of other component} = (1\text{u}) \times \frac{(2.33 \times 10^{-23}\text{ g})}{(1.66 \times 10^{-24}\text{ g})} = 14.04 \text{ u}$$

Molecular mass of the substance = (108 + 13 + 14.04) u = **135.04 u.**

III. Long Answer Type Questions

Question 1. Calculate no. of carbon and oxygen atoms present in 11.2 litres of CO_2 at N.T.P.

Answer: **Step I.** Number of CO_2 molecules in 11.2 litres

22.4 litres of CO_2 at N.T.P. = 1 gram mol

$$11.2 \text{ litres of } CO_2 \text{ at N.T.P.} = \frac{(1 \text{ gram mol})}{(22.4 \text{ litres})} \times (11.2 \text{ litres}) = 0.5 \text{ gram mol}$$

Now 1 gram mole of CO_2 contain molecules = 6.022×10^{23}

$\therefore$ 0.5 gram mole of CO_2 contain molecules = $6.022 \times 10^{23} \times 0.5 = 3.011 \times 10^{23}$

Step II. *Number of carbon and oxygen atoms in 3.011×10^{23} molecules of CO_2*

1 molecule of CO_2 contains carbon atoms = 1

$\therefore$ 3.011×10^{23} molecules of CO_2 will contain carbon atoms = $\mathbf{3.011 \times 10^{23}}$

Similarly, 1 molecule of CO_2 contains oxygen atoms = 2

$\therefore$ 3.011×10^{23} molecules of CO_2 will contain oxygen atoms = $2 \times 3.011 \times 10^{23}$ = $\mathbf{6.022 \times 10^{23}}$ **atoms.**

Question 2. $KClO_3$ on heating decomposes to give KCl and O_2. What is the volume of O_2 at N.T.P liberated by 0.1 mole of $KClO_3$?

Answer: The chemical equation for the decomposition of $KClO_3$ is

$$\underset{2\text{ mol}}{2KClO_3} \xrightarrow{\text{Heat}} 2KCl + \underset{\substack{3\text{ mol} \\ (3 \times 22.4\text{ L} = 67.2\text{L})}}{3O_2}$$

2 moles of $KClO_3$ evolve O_2 at N.T.P. = 67.2 L

$$1 \text{ mole of } KClO_3 \text{ evolve } O_2 \text{ at N.T.P.} = \frac{67.2}{2} \text{ L}$$

$$0.1 \text{ mole of } KClO_3 \text{ evolve } O_2 \text{ at N.T.P.} = \frac{67.2}{2} \times 0.1\text{L} = \mathbf{3.36 \text{ L}}$$

Question 3. 10 ml of a solution of NaCl containing KCl gave on evaporation 0.93 g of the mixed salt which gave 1.865 g of AgCl by reacting with $AgNO_3$ solution. Calculate the quantity of NaCl in 10 mL of the solution.

Answer: The chemical equation for the reaction is:

$$\underset{23 + 35.5\ (= 58.5\ g)}{AgNO_3 + NaCl} \longrightarrow AgCl + \underset{108 + 35.5\ (= 143.5\ g)}{NaNO_3}$$

Let the mass of NaCl and KCl in the mixture be respectively a g and b g.

$\therefore \quad a + b = 0.93$ (given)

Let us find AgCl formed on reacting NaCl and KCl with $AgNO_3$ solution.

58.5 g of NaCl give AgCl = 143.5 g

$$\therefore \quad a \text{ g of NaCl will give AgCl} = \frac{(143.5\ g)}{(58.5\ g)} \times (a\ g)$$

Similarly, 74.5 g of KCl give AgCl = 143.5 g

$$b \text{ g of KCl will give AgCl} = \frac{(143.5\ g)}{(74.5\ g)} \times (b\ g)$$

But mass of AgCl actually formed = 1.865 g (given)

$$\therefore \quad \frac{143.5 \times a}{58.5} + \frac{143.5 \times b}{74.5} = 1.865; \quad \frac{143.5 \times a}{58.5} + \frac{143.5(0.93 - a)}{74.5} = 1.865$$

$$2.453\ a + 1.93(0.93 - a) = 1.865; \ 2.453\ a + 1.795 - 1.93\ a = 1.865$$

$$0.523\ a = 0.07 \quad \text{or} \quad a = \frac{0.07}{0.523} = 0.14$$

Mass of NaCl in the mixture = **0.14 g**

Mass of KCl in the mixture = (0.93 – 0.14) = **0.79 g.**

Question 4. The cost of table salt (NaCl) and table sugar ($C_{12}H_{22}O_{11}$) are Rs 1 per kg and Rs 6 per kg respectively.Calculate their cost per mole.

Answer: (a) Cost of table salt (NaCl) per mole

Gram molecular mass of NaCl = 23 + 35.5 = 58.5 g Now, 1000 g of NaCl cost = Rs 2

$$\therefore \quad 58.5 \text{ g of NaCl will cost} = \frac{2}{(1000\ g)} \times (58.5\ g) = 0.117 \text{ Rupee}$$

$$= 0.117 \times 100 = \textbf{12 paise (approx.)}$$

(b) Cost of table sugar ($C_{12}H_{22}O_{11}$) per mole

Gram molecular mass of ($C_{12}H_{22}O_{11}$) = 12 x 12 + 22 x 1 = 16 x 1= 144 + 22 + 176 = 342 g

Now, 1000 g of sugar cost = Rs 6

$$\therefore \quad 342 \text{ g of sugar will cost} = \frac{6}{(1000\ g)} \times (342\ g) = 2.052$$

$$= \textbf{2.0 Rupees (approx.)}$$

Question 5. A flask P contains 0.5 mole of oxygen gas. Another flask Q contains 0.4 mole of ozone gas. Which of the two flasks contains greater number of oxygen atoms?

Answer: 1 molecule of oxygen (O_2) = 2 atoms of oxygen

1 molecule of ozone (O_3) = 3 atoms of oxygen

In flask P: 1 mole of oxygen gas = 6.022×10^{23} molecules

0.5 mole of oxygen gas = $6.022 \times 10^{23} \times 0.5$ molecules

$= 6.022 \times 10^{23} \times 0.5 \times 2$ atoms = $\mathbf{6.022 \times 10^{23}}$ **atoms**

In flask Q: 1 mole of ozone gas = 6.022×10^{23} molecules

0.4 mole of ozone gas = $6.022 \times 10^{23} \times 0.4$ molecules

$= 6.022 \times 10^{23} \times 0.4 \times 3$ atoms = $\mathbf{7.23 \times 10^{22}}$ **atoms**

$\therefore$ Flask Q has a greater number of oxygen atoms as compared to the flask P.

Question 6. Calculate the total number of electrons present in 1.6 g of methane.

Answer:

(*i*) Molar mass of methane (CH_4) = 12 + 4 × 1 = 16 g

16 g of methane contain molecules = 6.022×10^{23}

$$1.6 \text{ g of methane contain molecule} = \frac{6.022 \times 10^{23}}{(16\text{ g})} \times (1.6\text{ g}) = 6.022 \times 10^{22}$$

(*ii*) *Number of electrons in 6.022×10^{22} molecules of methane*

1 molecule of methane contains electrons = 6 + 4 = 10

6.022×10^{22} molecules of methane contain electrons

$= 6.022 \times 10^{22} \times 10 = \mathbf{6.022 \times 10^{23}}$.

Question 7. The vapour density of a mixture of NO_2 and N_2O_4 is 38.3 at 27°C. Calculate the number of moles of NO_2 in 100 g of the mixture.

Answer:

Vapour density of the mixture of NO_2 and N_2O_4 = 38.3

Molecular mass of the mixture = 2 × Vapour density = 2 × 38.3 = 76.6 u = 76.6 g

Mass of the mixture = 100 g

$$\text{No. of moles of the mixture} = \frac{100}{76.6}$$

Let the mass of NO_2 in the mixture = x g

∴ Mass of N_2O_4 in the mixture = $(100 - x)$ g

Molar mass of NO_2 = 14 + 32 = 46 u = 46 g

Molar mass of N_2O_4 = 28 + 64 = 92 u = 92 g

$$\text{No. of moles of } NO_2 = \frac{x}{46}$$

$$\text{No. of moles of } N_2O_4 = \frac{(100 - x)}{92}$$

$$\text{Total no. of moles in the mixture} = \frac{x}{46} + \frac{(100 - x)}{92}$$

$$\text{Equating } (i) \text{ and } (ii),\ \frac{x}{46} + \frac{(100 - x)}{92} = \frac{100}{76.6}$$

$$92x + 46(100 - x) = \frac{100}{76.6} \times 46 \times 92 = 5524.8$$

$$92x - 46x = 5524.8 - 4600 = 924.8.$$

Question 8. The Vapour Density of a gaseous element is 5 times that of oxygen under similar conditions. If the molecule is triatomic, what will be its atomic mass?

Answer: Molecular mass of oxygen = 32 u

$$\text{Density of oxygen} = \frac{32}{2} = 16 \text{ u}$$

Density of gaseous element = 16 × 5 = 80 u

Molecular mass of gaseous element = 80 × 2 = 160 u

Atomicity of the element = 3

$$\text{Atomic mass of the element} = \frac{\text{Molecular Mass}}{\text{Atomicity}} = \frac{160}{3} = \mathbf{53.33\ u.}$$

1/6th of its volume. Since equal volumes of gases have equal number of moles according to Avogadro's Law,

$$\therefore \quad \frac{\text{Moles of } CO_2}{\text{Moles of both gases}} = \frac{x}{(2x+y)} = \frac{1}{6}$$

or $6x = 2x + y$ or $4x = y$ or $\frac{y}{x} = 4$

∴ **Molar ratio of formic acid : oxalic acid = 4 : 1.**

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