1 Some Basic Concepts of Chemistry

LESSON AT A GLANCE

- All substances contain matter which can exist in three states—solid, liquid or gas.
- Matter can also be classified into elements, compounds and mixtures.
- **Element:** An element contains particles of only one type which may be atoms or molecules.
- **Compounds:** They are formed where atoms or two or more elements combine in a fixed ratio to each other.
- **Mixtures:** Many of the substances present around us are mixtures.
- **Scientific notation:** The measurement of quantities in chemistry are spread over a wide range of 10^{-31} to 10^{23}. Hence, a convenient system of expressing the number in scientific notation is used.
- **Scientific figures:** The uncertainty is taken care of by specifying the number of significant figures in which the observations are reported.
- **Dimensional analysis:** It helps to express the measured quantities in different systems of units.
- **Laws of Chemical Combinations are:**
  (*i*) Law of Conservation of mass
  (*ii*) Law of Definite proportions
  (*iii*) Law of Multiple proportions
  (*iv*) Gay Lussac's law of Gaseous volumes
  (*v*) Avogadro's law.
- **Atomic mass:** The atomic mass of an element is expressed relative to ^{12}C isotope of carbon which has an exact value of 12u.

- **Average atomic mass:** Obtained by taking into account the natural aboundance of different isotopes of that element.
- **Molecular mass:** The molecular mass of a molecule is obtained by taking sum of atomic masses of different atoms present in a molecule.
- **Avogadro number:** The number of atoms, molecules or any other particles present in a given system are expressed in terms of Avogadro constant.

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

- **Balanced chemical equation:** A balanced equation has the same number of atoms of each element on both sides of the equation.
- **Stoichiometry:** The quantitative study of the reactants required or the products formed is called stoichiometry. Using stoichiometric calculations, the amounts of one or more reactants required to produce a particular amount of product can be determined and vice-versa.

TEXTBOOK QUESTIONS SOLVED

Q1. *Calculate the molecular mass of the following:*

(*i*) H_2O (*ii*) CO_2

(*iii*) CH_4

Ans. Molecular masses of:

(*i*) $H_2O = (2 \times 1.008) + (1 \times 16.00) = 18.016$ amu.

(*ii*) $CO_2 = (1 \times 12.01) + (2 \times 16.00) = 44.01$ amu.

(*iii*) $CH_4 = (1 \times 12.01) + (4 \times 1.008) = 16.04$ amu.

Q2. *Calculate the mass per cent of different elements present in sodium sulphate (Na_2SO_4).*

Ans. Mass per cent of an element

$$= \frac{\text{Mass of the element in 1 mole of substance}}{\text{Molar mass of the substance}} \times 100$$

Molar mass of Na_2SO_4

$$= (2 \times 23) + (1 \times 32) + (4 \times 16) \text{ g mol}^{-1}$$
$$= 142 \text{ g mol}^{-1}$$

Mass of sodium in 1 mole of $Na_2SO_4 = 2 \times 23 = 46$ g mol^{-1}

Mass of sulphur in 1 mole of $Na_2SO_4 = 32$ g mol^{-1}

Mass of oxygen in 1 mole of $Na_2SO_4 = 4 \times 16 = 64$ g mol^{-1}

Mass % of Na

$$= \frac{\text{Mass of sodium in 1 mol } Na_2SO_4}{\text{Molar mass of } Na_2SO_4} \times 100$$

$$= \frac{46 \text{ g mol}^{-1}}{142 \text{ g mol}^{-1}} \times 100 = \mathbf{32.39\%}$$

Mass % of S

$$= \frac{\text{Mass of sulphur in 1 mol } Na_2SO_4}{\text{Molar mass of } Na_2SO_4} \times 100$$

$$= \frac{32 \text{ g mol}^{-1}}{142 \text{ g mol}^{-1}} \times 100 = \mathbf{22.54\%}$$

Mass % of O

$$= \frac{\text{Mass of oxygen in 1 mol } Na_2SO_4}{\text{Molar mass of } Na_2SO_4} \times 100$$

$$= \frac{64 \text{ g mol}^{-1}}{142 \text{ g mol}^{-1}} \times 100 = \mathbf{45.07\%}$$

Q3. *Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% oxygen by mass.*

Ans. Given: Mass per cent of iron = 69.9% and that of oxygen = 30.1%

∴ Masses of elements per 100 g of compound:

iron = 69.9 g

oxygen = 30.1 g

Element	*Mass per 100 g compound /g*	*Molar mass /g mol⁻¹*	*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	31.1	16.00	$\frac{30.1}{16.00} = 1.88$	$\frac{1.88}{1.25} = 1.5$	3

∴ Empirical formula = Fe_2O_3

Q4. *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.

Ans. The reaction is:

$$\underset{\substack{1\ \text{mol} \\ 12\ \text{g}}}{C(s)} + \underset{\substack{1\ \text{mol} \\ 32\ \text{g}}}{O_2(g)} \longrightarrow \underset{\substack{1\ \text{mol} \\ 44\ \text{g}}}{CO_2(g)}$$

According to this equation:

(*i*) When 1 mole of carbon is burnt in air, one mole of CO_2 (44 g) would be produced.

(*ii*) 1 mole of carbon is burnt in 16 g of dioxygen. Since 1 mole of carbon needs 32 g of oxygen to burn and form CO_2 the reaction would not be completed and dioxygen is the limiting reagent.

By 32 g of dioxygen, the amount of CO_2 formed = 1 mole

∴ By 16 g of dioxygen, the amount of

$$CO_2 \text{ formed} = \frac{1\ \text{mol} \times 16\ \text{g}}{32\ \text{g}} = \mathbf{0.5\ mole\ or\ 22\ g}$$

(*iii*) 2 moles of carbon are burnt in 16 g of dioxygen. 2 moles of carbon would need 2 moles (64 g) of dioxygen for complete burning but the mass of dioxygen available is only 16 g, it is the limiting reagent and the amount of carbon dioxide formed would be governed by it.

By 32 g of dioxygen, the amount of carbon dioxide formed = 1 mol

∴ By 16 g of dioxygen, the amount of carbon dioxide formed

$$= \frac{1\ \text{mol} \times 16\ \text{g}}{32\text{g}} = 0.5\ \text{mole} = \mathbf{22\ g.}$$

Q5. *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}.*

Ans. To make 1 L (1000 mL) of 0.375 molar solution, the moles of sodium acetate required = 0.375 mol

∴ To make 500 mL of 0.375 molar solution, the moles of sodium acetate required

$$= \frac{0.375\ \text{mol} \times 500\ \text{mL}}{1000\ \text{mL}} = (0.5 \times 0.375)\ \text{mol}$$

Since molar mass of sodium acetate

$= 82.0245\ \text{g mol}^{-1}$

$\therefore$ Mass of sodium acetate required

$= (0.5 \times 0.375)$ mol $\times$ 82.0245 g mol^{-1}

= **15.3796 g**

Q6. *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 per cent of nitric acid in it being 69%.*

Ans. Density of nitric acid solution

= 1.41 g mL^{-1}

Mass of 1 L (1000 mL) solution

= 1000 mL $\times$ 1.41 g mL^{-1} = 1410 g

Since mass % of nitric acid in the solution is 69%.

Mass of nitric acid in 1 L solution

$$= \frac{69 \times 1410}{100} = 972.9 \text{ g}$$

Molar mass of nitric acid (HNO_3)

$= (1 + 14 + 3 \times 16)$ g mol^{-1} = 63 g mol^{-1}

$\therefore$ Moles of nitric acid in 1 L solution

= Molarity of solution

$$= \frac{972.9 \text{ g}}{63 \text{g mol}^{-1}} = \mathbf{15.44\ M.}$$

Q7. *How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$)?*

Ans. Molar mass of $CuSO_4$

$= (63.5 + 32 + 4 \times 16)$ g mol^{-1} = 159.5 g mol^{-1}.

1 mole of $CuSO_4$ contains 1 mole of copper

$\therefore$ Mass of copper in 159.5 g $CuSO_4$ = 63.5 g

and mass of copper in 100 g $CuSO_4$ = $\frac{63.5 \times 100}{159.5}$ g

= **39.81 g**

Thus, 39.81 g copper can be obtained from 100 g $CuSO_4$.

Q8. *Determine the molecular formula of an oxide of iron in which the mass per cent of iron and oxygen are 69.9 and 30.1 respectively.*

Ans. Same as question number 1.3

$\therefore$ Empirical formula = Fe_2O_3

Since molar mass of the oxide is not given, molecular formula cannot be determined.

Q9. *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

Ans. From the data:

Fractional abundance of ^{35}Cl = 0.7577

Atomic mass = 34.9689 u

Fractional abundance of ^{37}Cl = 0.2423

Atomic mass = 36.9659 u

Average atomic mass

$= (0.7577) \times (34.9689 \text{ u}) + (0.2423) \times (36.9659 \text{ u})$

$= 26.4959 + 8.9568 = \mathbf{35.4897.}$

Q10. *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.

Ans. (i) 1 mole of ethane contains 2 moles of carbon atoms

∴ 3 moles of ethane would contain 2 × 3

= **6 moles of carbon atoms**

(ii) 1 mole of ethane contains 6 moles of hydrogen atoms

∴ 3 moles of ethane would contain 6 × 3

= **18 moles of hydrogen atoms**

(iii) 1 mole of ethane contains 6.022×10^{23} molecules of ethane

∴ 3 moles of ethane would contain $3 \times 6.022 \times 10^{23}$

= **18.066×10^{23} molecules of ethane**

Q11. *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 upto 2 L?*

Ans. Given: Mass of sugar dissolved = 20 g

Volume of solution prepared = 2 L

Molar mass of sugar, $C_{12}H_{22}O_{11}$

$= (12 \times 12) + (22 \times 1) + (16 \times 11) \text{ g mol}^{-1}$

$= 144 + 22 + 176 = 342 \text{ g mol}^{-1}$

Number of moles of sugar

$$= \frac{\text{Mass}}{\text{Molar mass}} = \frac{20 \text{ g}}{342 \text{ g mol}^{-1}}$$

$$= 0.0585 \text{ mol}$$

$$\text{Molarity} = \frac{\text{No. of moles of solute}}{\text{Volume of solution in litres}}$$

$$= \frac{0.0585 \text{ mol}}{2 \text{ L}} = \mathbf{0.0293 \text{ mol L}^{-1}}.$$

Q12. *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?*

Ans. Given: Molarity of solution = 0.25 mol L^{-1}

Volume of solution = 2.5 L

∴ Number of moles of methanol needed

$$= M \times V = (0.25 \text{ mol L}^{-1}) \times (2.5 \text{ L})$$

$$= 0.625 \text{ mol}$$

Molar mass of methanol, CH_3OH

$$= (12) + (3 \times 1) + (16) + 1 = 32 \text{ g mol}^{-1}$$

$$= 32 \times 10^{-3} \text{ kg mol}^{-1} = 0.032 \text{ kg mol}^{-1}$$

∴ Mass of methanol needed

$$= (0.625 \text{ mol}) \times (0.032 \text{ kg mol}^{-1}) = 0.02 \text{ kg}$$

Density of methanol = 0.793 kg L^{-1}

∴ Volume of methanol needed

$$= \frac{\text{Mass}}{\text{Density}} = \frac{0.02 \text{ kg}}{0.793 \text{ kg L}^{-1}}$$

$$= 0.02522 \text{ L} = \mathbf{25.22 \text{ mL}}$$

Q13. *Pressure is determined as force per unit area of the surface. The SI unit of pressure is pascal as shown below:*

$$1 \text{ Pa} = 1 \text{ Nm}^{-2}$$

If mass of air at sea level is 1034 g cm^{-2}, calculate the pressure in pascal.

Ans. $$\text{Pressure} = \frac{\text{Force}}{\text{Area}} = \frac{\text{Mass} \times \text{Acceleration}}{\text{Area}}$$

Given: Mass of air over 1 cm^2 area at sea level

$$= 1034 \text{ g} = 1034 \times 10^{-3} \text{ kg}$$

$$\text{Area} = 1 \text{ cm}^2 = 1 \times (10^{-3}\text{m})^2 = 1 \times 10^{-6} \text{ m}^2$$

Acceleration due to gravity = g = 9.8 ms^{-2}

$$\therefore \quad \text{Pressure} = \frac{(1034 \times 10^{-3}\ \text{kg})(9.8\ \text{m s}^{-2})}{1 \times 10^{-6}\ \text{m}^2}$$

$$= 1.01332 \times 10^5\ \text{kg m}^{-1}\text{s}^{-2} = \mathbf{1.01332\ Pa}$$

Q14. *What is the SI unit of mass?*

Ans. SI unit of mass = kg.

Q15. *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

Ans. (*i*) micro = 10^{-6} (*ii*) deca = 10

(*iii*) mega = 10^6 (*iv*) giga = 10^9

(*v*) femto = 10^{-15}

Q16. *What do you mean by significant figures?*

Ans. Significant figures is the number of certain (meaningful) digits plus one (last digit with uncertainty of ± 1).

Q17. *A sample of drinking water was found to be severely contaminated with chloroform, $CHCl_3$, supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).*

(*i*) *Express this in per cent by mass.*

(*ii*) *Determine the molality of chloroform in the water sample.*

Ans. Concentration of chloroform = 15 ppm, *i.e.*, it contains 15 parts of chloroform in one million (10^6) parts of drinking water by mass.

(*i*) ∴ Mass percentage

$$= \frac{\text{Mass of contaminant}}{\text{Total mass}} \times 100$$

$$= \frac{15 \times 100}{10^6} = \mathbf{1.5 \times 10^{-3}\%}$$

(*ii*) $$\text{Molality} = \frac{\text{Moles of contaminant}}{\text{Mass of drinking water (in kg)}}$$

From mass per cent:

Mass of chloroform in 100 g drinking water

$= 1.5 \times 10^{-3}$ g

$\therefore$ Mass of chloroform in contaminated 1000 g drinking water

$$= \frac{1.5 \times 10^{-3}\,g \times 1000\,g}{100\,g}$$

$$= 1.5 \times 10^{-2}\,g = 0.015\,g$$

Mass of uncontaminated water

$$= (1000 - 0.015)\,g = 999.985\,g = 1000\,g$$

Molar mass of chloroform, $CHCl_3$

$$= 12 + 1 + 3 \times 35.5\,g\,mol^{-1} = 119.5\,g\,mol^{-1}$$

$\therefore$ Mole of chloroform in 1000 g drinking water

$$= \text{molality} = \frac{0.015\,g\,kg^{-1}}{119.5\,g\,mol^{-1}}$$

$$= \mathbf{1.255 \times 10^{-4}\,mol\,kg^{-1}}$$

Q18. *Express the following in the scientific notation:*

(*i*) *0.0048* (*ii*) *234,000*

(*iii*) *8008* (*iv*) *500.0*

(*v*) *6.0012*

Ans. (*i*) $0.0048 = 4.8 \times 10^{-3}$ (*ii*) $234000 = 2.34 \times 10^5$

(*iii*) $8008 = 8.008 \times 10^3$ (*iv*) $500.0 = 5.000 \times 10^2$

(*v*) $6.0012 = 6.0012 \times 10^0$

Q19. *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*

Ans. No. of significant figures in

(*i*) 0.0025 = 2 (*ii*) 208 = 3

(*iii*) 5005 = 4 (*iv*) 126,000 = 3

(*v*) 500.0 = 4 (*vi*) 2.0034 = 5

Q20. *Round up the following upto three significant figures:*

(*i*) *34.216* (*ii*) *10.4107*

(*iii*) *0.04597* (*iv*) *2808*

Ans. (*i*) 34.2 (*ii*) 10.4

(*iii*) 0.0460 (*iv*) 2810

Q21. *The following data are obtained when dinitrogen and dioxygen react together to form different 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 = ... dm³*

Ans. (*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*) $1\ \text{km} = 10^3\ \text{m km}^{-1} \times 10^3\ \text{mm m}^{-1} = 10^6\ \text{mm}.$

$1\ \text{km} = 10^3\ \text{m km}^{-1} \times 10^{12}\ \text{pm m}^{-1} = \mathbf{10^{15}\ pm}$

(*ii*) $1\ \text{mg} = 10^{-3}\ \text{g mg}^{-1} \times 10^{-3}\ \text{kg g}^{-1} = 10^{-6}\ \text{kg}$

$1\ \text{mg} = 10^{-3}\ \text{g mg}^{-1} \times 10^{9}\ \text{ng g}^{-1} = \mathbf{10^{6}\ ng}$

(*iii*) $1\ \text{mL} = 10^{-3}\ \text{L mL}^{-1} = 10^{-3}\ \text{L}$

$1\ \text{mL} = 10^{-3}\ \text{L mL}^{-1} \times 1\ \text{dm}^3\ \text{L}^{-1} = \mathbf{10^{-3}\ dm^3}.$

Q22. *If the speed of light is $3.0 \times 10^8\ m\ s^{-1}$, calculate the distance covered by light in 2.00 ns?*

Ans. Distance = Speed × Time

$= 3.0 \times 10^8\ \text{m s}^{-1} \times 2 \times 10^{-9}\text{s} = 6 \times 10^{-1}\ \text{m} = \mathbf{0.6\ m}$

Q23. *In a reaction*

$$A + B_2 \longrightarrow AB_2$$

Identify the limiting reagent, if any, in the following reaction mixtures.

(*i*) *300 atoms of A + 200 molecules of B*

(*ii*) *2 mol A + 3 mol B*

(*iii*) *100 atoms of A + 100 molecules of B*

(*iv*) *5 mol A + 2.5 mol B*

(*v*) *2.5 mol A + 5 mol B*

Ans. The reaction is:

A	+	B_2	$\longrightarrow$	AB_2
1 atom		1 molecule 2 atoms		1 molecule
1 mol		1 mol		1 mol

(*i*) 300 atoms of A would require 300 molecules of B_2. Since only 200 molecules of B are available, *it is limiting reagent.*

(*ii*) 2 mol of A would require 2 mol of B. Since B is present in excess (3 mol), *A is the limiting reagent.*

(*iii*) 100 atoms of A would react completely with 100 molecules of B. Hence *no limiting reagent.*

(*iv*) 5 mol of A would require 5 mol of B, but only 2.5 mol are available, hence *B is the limiting reagent.*

(*v*) 2.5 mol of A would require 2.5 mol of B, but it is present in excess (5 mol), *A is the limiting reagent.*

Q24. *Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation:*

$$N_2(g) \;+\; 3H_2(g) \longrightarrow 2NH_3(g)$$

(*i*) *Calculate the mass of ammonia produced if 2.00×10^3 g dinitrogen reacts with 1.00×10^3 g of dihydrogen.*

(*ii*) *Will any of the two reactants remain unreacted?*

(*iii*) *If yes, which one and what would be its mass?*

Ans. Given reaction is:

$N_2(g)$	+	$3H_2(g)$	$\longrightarrow$	$2NH_3(g)$
1 mol		3 mol		2 mol
$1 \times 2 \times 14$		$3 \times 2 \times 1$		$2 \times (14 + 3)$
= 28 g		= 6 g		= 34 g

Starting masses: $N_2 = 2.00 \times 10^3$ g

$H_2 = 1.00 \times 10^3$ g

Thus, 28 g of dinitrogen would react with mass of dihydrogen = 6 g

$\therefore$ 2.00×10^3 g of dinitrogen would react with mass of dihydrogen

$$= \frac{6}{28} \times 2.00 \times 10^3$$

$$= 0.4286 \times 10^3 \text{ g} = 428.6 \text{ g}$$

Thus, H_2 is present in excess and N_2 is the limiting reagent.

(*i*) Mass of NH_3 produced would be governed by mass of N_2 reacting.

From equation:

28 g N_2 produces 34 g NH_3

$\therefore$ 2.00×10^3 g N_2 would produce $\frac{34}{28} \times 2.00 \times 10^3$ g NH_3

$= 2.429 \times 10^3$ g NH_3

= 2429 g NH_3

(*ii*) Yes, H_2 will remain unreacted since it is present in excess.

(*iii*) As calculated earlier, the mass of H_2 reacted = 428.6 g

$\therefore$ Mass of unreacted $H_2 = 1.0 \times 10^3 - 428.6$ g

$= 1000 - 428.6 =$ **571.4 g.**

Q25. *How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?*

Ans. *0.50 mol Na_2CO_3* : It means $0.5 \times 6.022 \times 10^{23}$

$= 3.011 \times 10^{23}$ formula units.

Also, molar mass of Na_2CO_3

$= 2 \times 23 + 12 + 3 \times 16$

$= 46 + 12 + 48 = 106$ g mol^{-1}

$\therefore$ 0.50 mol means $0.5 \times 106 = 53$ g Na_2CO_3

0.50 M Na_2CO_3: It means a solution containing 0.5 mol or 53 g of Na_2CO_3 dissolved per litre of solution.

Q26. *If ten volumes of dihydrogen gas react with five volumes of dioxygen gas, how many volumes of water vapour would be produced?*

Ans. Given reaction is

$$\underset{2\text{ volumes}}{2H_2(g)} + \underset{1\text{ volume}}{O_2(g)} \longrightarrow \underset{2\text{ volumes}}{2H_2O(g)}$$

Thus 2 volumes H_2 reacts with 1 volume O_2 to produce 2 volumes $H_2O(g)$.

$\therefore$ 10 volumes H_2 would react with 5 volumes O_2 to produce 10 volumes $H_2O(g)$.

Thus, the given volumes of H_2 (10 volumes) and O_2 (5 volumes) would react completely and produce 10 volumes of water vapour.

Q27. *Convert the following into basic units:*

(*i*) *28.7 pm* (*ii*) *15.15 pm*

(*iii*) *25365 mg*

Ans. (*i*) 28.7 pm $= 28.7 \times 10^{-12}$ m $= 2.87 \times 10^{-11}$ m

(*ii*) 15.15 pm $= 15.15 \times 10^{-12}$ m $= 1.515 \times 10^{-11}$ m

(*iii*) 25365 mg $= 25365 \times 10^{-6}$ kg $= 2.5365 \times 10^{-2}$ kg

Q28. *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)*

Ans. Let us calculate the number of atoms in each case.

(*i*) *1 g Au*(*s*): Molar mass of Au = 196.97 g mol^{-1}

$$\therefore \text{ Moles of Au(s)} = \frac{1}{196.97}\text{ mol}$$

$$\text{No. of atoms} = \frac{1}{196.97} \times 6.022 \times 10^{23}\text{ atoms}$$

(*ii*) *1 g (Na*(*s*): Molar mass of Na

$$= 22.99 \text{ g mol}^{-1}$$

$$\therefore \quad 1\text{ g Na(s)} = \frac{1}{22.99}\text{ mol}$$

$$\frac{1}{22.99} = 6.022 \times 10^{23}\text{ atoms}$$

(*iii*) *1 g Li*(s): Molar mass of Li

$$= 6.94 \text{ g mol}^{-1}$$

$$\therefore \quad 1\text{(g) Li(s)} = \frac{1}{6.94}\text{ mol}$$

$$= \frac{1}{6.94} \times 6.022 \times 10^{23}\text{ atoms}$$

(*iv*) *1 g Cl_2(g)*:

Molar mass of Cl_2 = 2 × 35.45 = 70.90 mol^{-1}

$$\therefore \quad 1\text{ g } Cl_2(g) = \frac{1}{70.90}\text{ mol}$$

$$= \frac{1}{70.9} \times 6.022 \times 10^{23}\text{ molecules of } Cl_2$$

$$= \frac{2}{70.9} \times 6.022 \times 10^{23}$$

$$= \frac{1}{35.45} \times 10^{23}\textbf{ atoms of Cl}$$

Comparing the number of atoms: The largest number of atoms will be present in 1 g Li(s). *Alternately*: Same conclusion can also be reached quickly by comparing the molar masses of elements. Since the mass is same in each case, the element with lowest molar mass would have the largest number of atom. Molar masses are:

Au = 196.97 g mol^{-1}, Na = 22.99 g mol^{-1},

Li = 6.94 g mol^{-1} and Cl = 35.45 g mol^{-1}

∴ Li with lowest molar mass would have the largest number of atoms.

Q29. *Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040.*

Ans. Mole fraction of ethanol in the solution = 0.040.

Since it is dilute solution, volume of solution would be the same as that of water.

For molarity, moles of ethanol present in 1 L solution (1 L water) have to be calculated

$$\therefore \quad x_{\text{ethanol}} = \frac{n_{\text{ethanol}}}{n_{\text{ethanol}} + n_{\text{water}}}$$

$$n_{\text{water}} = \text{no. of moles in 1 L water}$$

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

$$\therefore \quad 0.040 = \frac{n_{\text{ethanol}}}{n_{\text{ethanol}} + 55.55 \text{ mol}}$$

$$0.040\, n_{\text{ethanol}} + 0.040 \times 55.55 \text{ mol} = n_{\text{ethanol}}.$$

$$\therefore \quad (1 - 0.040)\, n_{\text{ethanol}} = 0.040 \times 55.55 \text{ mol}$$

$$0.96\, n_{\text{ethanol}} = 0.040 \times 55.55 \text{ mol}$$

$$\text{or} \quad n_{\text{ethanol}} = \frac{0.040 \times 55.55}{0.96} = 2.32 \text{ mol}$$

Thus, solution contains 2.32 mol of ethanol in 1 L solution.

$$\therefore \quad \text{Molarity} = \mathbf{2.32\ M.}$$

Q30. *What will be the mass of one ^{12}C atom in g?*

Ans. Molar mass of ^{12}C = 12.00 g mol^{-1}

$\therefore$ Mass of 6.022×10^{23} carbon atoms = 12.00 g

$$\therefore \quad \text{Mass of 1 carbon atom} = \frac{12.00}{6.023 \times 10^{23}} \text{g}$$

$$= \mathbf{1.99269 \times 10^{-23}\ g}$$

Q31. *How many significant figures should be present in the answer of the following calculations?*

(*i*) $\dfrac{0.02856 \times 298.15 \times 0.112}{0.5785}$

(*ii*) 5×5.364 (*iii*) $0.0125 + 0.7864 + 0.0215$

Ans. Number of significant figures present in answer to:

(*i*) $\dfrac{0.02856 \times 298.15 \times 0.112}{0.5785}$

Since the least number of significant figures (3) is present in 0.112, the answer should have 3 significant figures.

(*ii*) 5×5.364

Here 5 is an integer with infinite significant figures, the term 5.364 has the least number of significant figures (4). Therefore, the answer should have 4 significant figures.

(*iii*) $0.0125 + 0.7864 + 0.0215$

Here all the terms has 4 digits after decimal. Therefore, the answer should be reported upto 4 decimal places.

Q32. *Use the data given in the following table to calculate the molar massess of naturally occurring argon isotopes:*

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%

Ans. From the data, fractional abundance of:

(*i*) $^{36}Ar = 0.00337$ (*ii*) $^{38}Ar = 0.00067$

(*iii*) $^{40}Ar = 0.99600$

Molar mass of argon = $(0.00337 \times 35.96755)$
$+ (0.00067 \times 37.96272) + (0.99600 \times 39.9624)$
$= 0.1212 + 0.02544 + 39.8026 =$ **39.949.**

Q33. *Calculate the number of atoms in each of the following (i) 52 moles of Ar (ii) 52 u of He (iii) 52 g of He.*

Ans. (*i*) Number of atoms in 1 mol of Ar = 6.022×10^{23}

$\therefore$ Number of atom in 52 mol of Ar
$= 52 \times 6.022 \times 10^{23} = \mathbf{3.131 \times 10^{25}}$ **atoms**

(*ii*) Since atomic mass of He = 4 u

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

(*iii*) Molar mass of He = $4\ g\ mol^{-1}$

$\therefore$ 52 g He $= \frac{52}{4} = 13$ mol
$= 13 \times 6.022 \times 10^{23}$ atoms
$= \mathbf{7.8286 \times 10^{24}}$ **atoms.**

Q34. *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 STP) 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.*

Ans. (*i*) Mass of carbon in 3.38 g $CO_2 = \frac{3.38 \times 12}{44} = 0.9218$ g

$\therefore$ Mass of carbon in the sample of fuel gas burnt = **0.9218 g**

(*ii*) Mass of hydrogen in 0.690 g $H_2O = \frac{0.690 \times 2}{18}$

$= 0.0767$ g

$\therefore$ Mass of hydrogen in the sample of fuel gas burnt = 0.0767 g

Since the fuel gas contains only carbon and hydrogen, the mass of the gas burnt

$= 0.9218 + 0.0767 = 0.9985$ g

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

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

Table for empirical formula

Element	*Mass per 100 g comp. / g*	*Molar mass / g mol⁻¹*	*Number of moles*	*Simplest molar ratio*	*Simplest whole number ratio*
Carbon, C	92.32	12	$\frac{92.32}{12} = 7.69$	$\frac{7.69}{7.68} = 1$	1
Hydrogen, H	7.68	1	$\frac{7.68}{1} = 7.68$	$\frac{7.68}{7.68} = 1$	1

$\therefore$ Empirical formula of fuel gas = CH

and empirical formula mass = 12 + 1 = 13

For molar mass

Given: 10.0 L volume at STP weighs 11.6 g

$\therefore$ Mass of 1 mole, *i.e.*, 22.7 L of gas at

STP (273.15 K, 1 bar) $= \frac{11.6 \times 22.7}{10.0} = 26.3$ g

$\therefore$ Molecular mass of gas = 26.3

$$n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{26.3}{13} \simeq 2$$

$\therefore$ Molecular formula = 2(CH) = $\mathbf{C_2H_2}$.

Q35. *Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction,*

$$CaCO_3(s) + 2\,HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l).$$

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

Ans. Mass of HCl in 25 mL (0.025 L) of 0.75 M solution

Number of moles of HCl = $M \times V$ (in litres)

$= 0.75 \times 0.025 = 0.01875$ mol

Molar mass of HCl = 1 + 35.5 = 36.5 g mol^{-1}

$\therefore$ Mass of HCl in 25 mL of 0.75 M solution

$= (0.01875 \text{ mol}) \times (36.5 \text{ g mol}^{-1}) = 0.6844$ g

The given reaction is:

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

1 mol	2 mol
$1 \times (40 + 12 + 3 \times 16)$	$2 \times (1 + 35.5)$
= 100 g	= 73 g

$\therefore$ 73 g HCl reacts with 100 g $CaCO_3$

and 0.6844 g HCl would react with $\frac{100 \times 0.6844}{73}$ g

$= \mathbf{0.9375\ g}$.

Q36. *Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction*

$$4\,HCl(aq) + MnO_2(s) \longrightarrow 2H_2O(l) + MnCl_2(aq) + Cl_2(g)$$

How many grams of HCl react with 5.0 g of manganese dioxide?

Ans. The given reaction is:

$$4HCl(aq) + MnO_2(s) \longrightarrow 2H_2O(l) + MnCl_2(aq) + Cl_2(g)$$

4 mol	1 mol
$4 \times (1 + 35.5)$	$(54.94 + 2 \times 16)$
= 146 g	= 86.94 g

$\therefore$ 86.94 g MnO_2 reacts with 146 g HCl

5.0 g MnO_2 would react with

$$\frac{146 \times 5.0}{86.94} \text{ g} = \mathbf{8.40\ g\ HCl.}$$

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