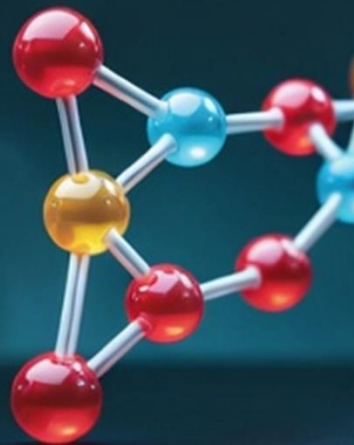


Part - I

NCERT SOLUTIONS FOR CLASS 12 CHEMISTRY

CBSE



Class 12

DETAILED SOLUTIONS

Class 12 Chemistry — NCERT

Chapters: Solutions • Electrochemistry • d & f Block • Coordination Compounds

Prepared with step-by-step explanations for average students



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CHAPTER 1: SOLUTIONS

Q 1.1 — Types of Solutions

Question: Define the term solution. How many types of solutions are formed? Write briefly about each type with an example.

A solution is a homogeneous mixture of two or more substances. The component present in the largest amount is called the solvent, and the remaining components are called solutes.

 *Note: Homogeneous means the composition is uniform throughout — you cannot distinguish individual components under a microscope.*

9 Types of Solutions (based on physical states of solute and solvent):

1. Gas in Gas — e.g., Air (mixture of N_2 , O_2 , CO_2 , etc.)
2. Gas in Liquid — e.g., CO_2 dissolved in water (soda water)
3. Gas in Solid — e.g., H_2 gas dissolved in palladium metal
4. Liquid in Gas — e.g., Water vapour in air (humid air)
5. Liquid in Liquid — e.g., Alcohol dissolved in water (ethanol + water)
6. Liquid in Solid — e.g., Mercury in zinc (amalgam)
7. Solid in Gas — e.g., Camphor vapours in air (sublimation in air)
8. Solid in Liquid — e.g., Salt ($NaCl$) dissolved in water
9. Solid in Solid — e.g., Gold dissolved in silver (alloy)

Q 1.2 — Solid Solution with Gas Solute

Question: Give an example of a solid solution in which the solute is a gas.

When hydrogen gas (H_2) is adsorbed into palladium (Pd) metal, the gas molecules occupy interstitial spaces in the solid lattice. Another common example is the solution of gases like O_2 and N_2 in some minerals or gemstones.

✓ Example: H_2 gas dissolved in palladium (Pd) — a Solid-in-Solid type where the solute is a gas.

Q 1.3 — Key Definitions

Question: Define: (i) Mole fraction (ii) Molality (iii) Molarity (iv) Mass percentage.

(i) Mole Fraction (x):

Mole fraction of a component = (moles of that component) / (total moles of all components)

For a binary solution: $x_A = n_A / (n_A + n_B)$ and $x_B = n_B / (n_A + n_B)$

Note: $x_A + x_B = 1$ always. Mole fraction has no units.

 *Note: Mole fraction is temperature-independent since moles don't change with temperature.*

(ii) Molality (m):

Molality = moles of solute / mass of solvent in kg

Unit: mol kg^{-1}

Molality is preferred for colligative property calculations because it is independent of temperature.

(iii) Molarity (M):

Molarity = moles of solute / volume of solution in litres

Unit: mol L^{-1} (or M)

Molarity changes with temperature (volume of solution expands/contracts).

(iv) Mass Percentage (w/w %):

$\text{Mass \%} = (\text{mass of solute} / \text{mass of solution}) \times 100$

Unit: % (dimensionless, often written as % w/w)

Example: 10% NaCl solution = 10 g NaCl in 100 g solution.

Q 1.4 — Molarity of Concentrated HNO_3

Question: Concentrated nitric acid is 68% by mass. Density = 1.504 g/mL. Find its molarity.

 Note: $\text{Molarity} = (10 \times \text{density} \times \% \text{ purity}) / \text{molar mass}$ — this formula works when density is in g/mL.

Molar mass of $\text{HNO}_3 = 1 + 14 + 48 = 63 \text{ g/mol}$

Mass of HNO_3 in 1 L solution = $1.504 \text{ g/mL} \times 1000 \text{ mL} \times 0.68 = 1022.72 \text{ g}$

Moles of $\text{HNO}_3 = 1022.72 / 63 = 16.23 \text{ mol}$

✓ Molarity = $16.23 \text{ M} \approx 16.2 \text{ mol/L}$

Q 1.5 — 10% Glucose Solution

Question: A 10% w/w glucose solution. Find molality, mole fraction of each component, and molarity (density = 1.2 g/mL).

Take 100 g of solution $\rightarrow$ 10 g glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) + 90 g water.

Molar mass of glucose = 180 g/mol; water = 18 g/mol

Moles of glucose = $10/180 = 0.0556 \text{ mol}$

Moles of water = $90/18 = 5 \text{ mol}$

Mole Fractions:

$x(\text{glucose}) = 0.0556 / (0.0556 + 5) = 0.0556 / 5.0556 = 0.0110$

$x(\text{water}) = 5 / 5.0556 = 0.9890$

Molality:

$m = \text{moles of glucose} / \text{mass of solvent in kg} = 0.0556 / 0.090 = 0.617 \text{ mol/kg}$

Molarity:

Volume of 100 g solution = $100 / 1.2 = 83.33 \text{ mL} = 0.08333 \text{ L}$

Molarity = $0.0556 / 0.08333 = 0.667 \text{ mol/L}$

✓ $x(\text{glucose}) = 0.011$, $x(\text{water}) = 0.989$, $m = 0.617 \text{ mol/kg}$, $M = 0.667 \text{ mol/L}$

Q 1.6 — Volume of HCl to React with $\text{Na}_2\text{CO}_3 + \text{NaHCO}_3$ Mixture

Question: How many mL of 0.1 M HCl are required to react with 1 g mixture of equimolar Na_2CO_3 and NaHCO_3 ?

 Note: The mixture contains equal moles (not equal masses) of both. Let moles of each = x .

Molar mass: $\text{Na}_2\text{CO}_3 = 106 \text{ g/mol}$, $\text{NaHCO}_3 = 84 \text{ g/mol}$

Mass equation: $106x + 84x = 1 \text{ g} \rightarrow 190x = 1 \rightarrow x = 0.00526 \text{ mol}$

Reactions with HCl:

$\text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{H}_2\text{O} + \text{CO}_2 \rightarrow$ needs 2x moles of HCl

$\text{NaHCO}_3 + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O} + \text{CO}_2 \rightarrow$ needs x moles of HCl

Total HCl needed = $2x + x = 3x = 3 \times 0.00526 = 0.01578 \text{ mol}$

Volume = moles / molarity = $0.01578 / 0.1 = 0.1578 \text{ L} = 157.8 \text{ mL}$

✓ $\approx 157.8 \text{ mL}$ of 0.1 M HCl

Q 1.7 — Mass Percentage of Mixed Solution

Question: 300 g of 25% solution + 400 g of 40% solution. Calculate mass % of resulting solution.

Mass of solute from 1st solution = $25/100 \times 300 = 75 \text{ g}$

Mass of solute from 2nd solution = $40/100 \times 400 = 160 \text{ g}$

Total solute = $75 + 160 = 235 \text{ g}$

Total solution = $300 + 400 = 700 \text{ g}$

Mass % = $(235/700) \times 100 = 33.57\%$

✓ Mass percentage = 33.57%

Q 1.8 — Molality and Molarity of Ethylene Glycol Solution

Question: 222.6 g ethylene glycol ($\text{C}_2\text{H}_6\text{O}_2$) + 200 g water. Find molality and molarity (density = 1.072 g/mL).

Molar mass of $\text{C}_2\text{H}_6\text{O}_2 = 24 + 6 + 32 = 62 \text{ g/mol}$

Moles of ethylene glycol = $222.6 / 62 = 3.590 \text{ mol}$

Molality:

$m = 3.590 \text{ mol} / 0.200 \text{ kg} = 17.95 \text{ mol/kg}$

Molarity:

Total mass of solution = $222.6 + 200 = 422.6 \text{ g}$

Volume = $422.6 / 1.072 = 394.2 \text{ mL} = 0.3942 \text{ L}$

$M = 3.590 / 0.3942 = 9.11 \text{ mol/L}$

✓ Molality = 17.95 m; Molarity = 9.11 M

Q 1.9 — Chloroform Contamination (15 ppm)

Question: Drinking water is contaminated with CHCl_3 at 15 ppm (by mass). Express in % by mass and find molality.

 Note: ppm = parts per million by mass = mg of solute per kg of solution. 15 ppm = 15 mg CHCl_3 per 10^6 mg solution.

(i) % by mass:

$15 \text{ ppm} = 15/10^6 \times 100 = 1.5 \times 10^{-3} \% \text{ by mass}$

(ii) Molality:

In 1 kg of water (approximately, since concentration is very low):

Mass of $\text{CHCl}_3 = 15 \times 10^{-3} \text{ g}$ (from 15 ppm definition: 15 mg per kg)

Molar mass of $\text{CHCl}_3 = 12 + 1 + 3 \times 35.5 = 119.5 \text{ g/mol}$

Moles of $\text{CHCl}_3 = 15 \times 10^{-3} / 119.5 = 1.255 \times 10^{-4} \text{ mol}$

Molality = $1.255 \times 10^{-4} \text{ mol} / 1 \text{ kg} = 1.26 \times 10^{-4} \text{ mol/kg}$

✓ % by mass = $1.5 \times 10^{-3}\%$; $m = 1.26 \times 10^{-4} \text{ mol/kg}$

Q 1.10 — Alcohol and Water Interactions

Question: What role does molecular interaction play in a solution of alcohol and water?

Pure alcohol has $\text{O}-\text{H}\cdots\text{O}$ hydrogen bonds between alcohol molecules. Pure water also has extensive $\text{O}-\text{H}\cdots\text{O}$ hydrogen bonds.

When mixed, alcohol molecules form hydrogen bonds WITH water molecules ($\text{O}-\text{H}\cdots\text{O}$ between the $-\text{OH}$ of alcohol and water).

However, the new interactions formed between alcohol and water are weaker than those in pure components. This results in:

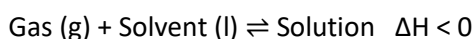
- Positive deviation from Raoult's law (vapour pressure is higher than ideal)
- $\Delta H_{\text{mix}} > 0$ (slight endothermic mixing)
- $\Delta V_{\text{mix}} > 0$ (volume increases slightly on mixing)

 *Note: Ethanol-water is a nearly ideal solution but shows slight positive deviation due to weaker cross interactions.*

Q 1.11 — Gas Solubility and Temperature

Question: Why do gases always tend to be less soluble in liquids as the temperature is raised?

The dissolution of gas in a liquid is generally an exothermic process:



According to Le Chatelier's principle, increasing temperature shifts the equilibrium to the left (towards the gas phase), i.e., gas comes out of solution.

Also, increasing temperature increases the kinetic energy of gas molecules, enabling them to overcome intermolecular forces and escape the solution.

 *Note: This is why carbonated drinks go flat when warm — CO_2 becomes less soluble at higher temperature.*

Q 1.12 — Henry's Law

Question: State Henry's law and mention some important applications.

Henry's Law states: The partial pressure of a gas above a liquid is directly proportional to the mole fraction of the gas dissolved in the liquid at constant temperature.

$$p = K_H \times x$$

where p = partial pressure of gas, x = mole fraction of gas in liquid, K_H = Henry's law constant (temperature dependent).

Applications:

1. Carbonation of soft drinks — CO_2 is dissolved under high pressure (sealed bottles). Opening the bottle releases pressure, reducing solubility.
2. Scuba diving — At high depths, high pressure dissolves more N_2 in the blood. Rapid ascent causes 'bends' (N_2 bubbles in blood). Divers use $\text{He}-\text{O}_2$ mixtures (He has a higher K_H , so less dissolves).

3. Respiration at high altitudes — Lower atmospheric O₂ pressure → less O₂ dissolves in blood → altitude sickness.
4. Industrial absorption processes — Henry's law governs how much of an industrial gas is absorbed into a liquid scrubber.

Q 1.13 — Henry's Law Calculation

Question: Partial pressure of ethane over a solution containing 6.56×10^{-3} g is 1 bar. If the solution contains 5.00×10^{-2} g of ethane, what is the partial pressure?

 Note: Henry's law: $p \propto \text{mole fraction} \approx \propto \text{mass of dissolved gas}$ (since the solvent mass is the same).

Since the solvent is the same in both cases, mole fraction $\propto$ moles $\propto$ mass (same molar mass).

$$p_1/p_2 = m_1/m_2$$

$$1 \text{ bar} / p_2 = 6.56 \times 10^{-3} \text{ g} / 5.00 \times 10^{-2} \text{ g}$$

$$p_2 = 5.00 \times 10^{-2} / 6.56 \times 10^{-3} = 7.62 \text{ bar}$$

✓ Partial pressure = 7.62 bar

Q 1.14 — Raoult's Law Deviations

Question: What is meant by positive and negative deviations from Raoult's law? How is $\Delta_{\text{mix}}H$ related to these deviations?

Raoult's law (ideal solution): $p_A = x_A \times p^\circ_A$ (partial pressure = mole fraction $\times$ vapour pressure of pure A)

Positive Deviation:

Observed vapour pressure > predicted by Raoult's law.

Solute-solvent interactions are WEAKER than solute-solute and solvent-solvent.

$\Delta_{\text{mix}}H > 0$ (endothermic) — energy is absorbed to break existing bonds.

Example: Acetone + Carbon disulphide, Ethanol + Water.

Negative Deviation:

Observed vapour pressure < predicted by Raoult's law.

Solute-solvent interactions are STRONGER than individual interactions.

$\Delta_{\text{mix}}H < 0$ (exothermic) — energy is released when new stronger bonds form.

Example: Acetone + Chloroform (H-bond formation between C=O and CHCl₃).

Q 1.15 — Molar Mass from Relative Lowering of Vapour Pressure

Question: A 2% non-volatile solute solution exerts pressure of 1.004 bar at the normal boiling point. Molar mass of solute?

 Note: At the normal boiling point of the solvent, $p^\circ_{\text{solvent}} = 1.013 \text{ bar}$ (1 atm).

$$p^\circ(\text{water}) = 1.013 \text{ bar}, p(\text{solution}) = 1.004 \text{ bar}$$

$$\text{Relative lowering: } (p^\circ - p)/p^\circ = x(\text{solute})$$

$$(1.013 - 1.004)/1.013 = 0.009/1.013 = 0.00888 = x(\text{solute})$$

$$\text{In 100 g solution: } 2 \text{ g solute, } 98 \text{ g water } (= 98/18 = 5.444 \text{ mol})$$

$$x(\text{solute}) = n(\text{solute}) / (n(\text{solute}) + 5.444) \approx n(\text{solute})/5.444 \text{ (for dilute solution)}$$

$$n(\text{solute}) = 0.00888 \times 5.444 = 0.04834 \text{ mol}$$

Molar mass = $2 / 0.04834 = 41.4 \text{ g/mol}$

✓ Molar mass $\approx 41.4 \text{ g/mol}$

Q 1.16 — Heptane + Octane Ideal Solution

Question: At 373 K, $p^\circ_{\text{heptane}} = 105.2 \text{ kPa}$, $p^\circ_{\text{octane}} = 46.8 \text{ kPa}$. Find vapour pressure of 26.0 g heptane + 35 g octane.

Molar mass: Heptane (C_7H_{16}) = 100 g/mol ; Octane (C_8H_{18}) = 114 g/mol

Moles of heptane = $26.0/100 = 0.26 \text{ mol}$

Moles of octane = $35/114 = 0.307 \text{ mol}$

Total moles = $0.26 + 0.307 = 0.567 \text{ mol}$

$x(\text{heptane}) = 0.26/0.567 = 0.4586$

$x(\text{octane}) = 0.307/0.567 = 0.5414$

$p(\text{heptane}) = 0.4586 \times 105.2 = 48.24 \text{ kPa}$

$p(\text{octane}) = 0.5414 \times 46.8 = 25.34 \text{ kPa}$

Total P = $48.24 + 25.34 = 73.58 \text{ kPa}$

✓ Total vapour pressure = 73.58 kPa

Q 1.17 — Vapour Pressure of 1 Molal Solution

Question: Vapour pressure of water at 300 K is 12.3 kPa . Find VP of 1 molal non-volatile solute solution.

1 molal = 1 mol solute per 1000 g water

Moles of water = $1000/18 = 55.56 \text{ mol}$

$x(\text{solute}) = 1/(1 + 55.56) = 1/56.56 = 0.01768$

$x(\text{water}) = 1 - 0.01768 = 0.98232$

$p = x(\text{water}) \times p^\circ = 0.98232 \times 12.3 = 12.08 \text{ kPa}$

✓ Vapour pressure = 12.08 kPa

Q 1.18 — Mass of Solute to Reduce VP to 80%

Question: Molar mass of non-volatile solute = 40 g/mol . How much should be dissolved in 114 g octane to reduce VP to 80%?

$p/p^\circ = 0.80 \rightarrow (p^\circ - p)/p^\circ = 0.20 = x(\text{solute})$

Molar mass of octane (C_8H_{18}) = 114 g/mol

Moles of octane = $114/114 = 1 \text{ mol}$

$x(\text{solute}) = n(\text{solute})/(n(\text{solute}) + 1) = 0.20$

$n(\text{solute}) = 0.20 \times (n(\text{solute}) + 1)$

$0.80 \times n(\text{solute}) = 0.20 \rightarrow n(\text{solute}) = 0.25 \text{ mol}$

Mass of solute = $0.25 \times 40 = 10 \text{ g}$

✓ 10 g of the solute must be dissolved.

Q 1.19 — Molar Mass and p° of Water

Question: 30 g solute in 90 g water has VP = 2.8 kPa. Then 18 g water added, VP = 2.9 kPa. Find molar mass of solute and p° of water.

 Note: This is a system of two equations.

Let molar mass of solute = M ; p° = vapour pressure of pure water.

Case 1: $n(\text{solute}) = 30/M$; $n(\text{water}) = 90/18 = 5$

$$p_1/p^\circ = x(\text{water}) = 5/(5 + 30/M) = 2.8/p^\circ \dots (i)$$

Case 2: $n(\text{water}) = (90+18)/18 = 6$; $n(\text{solute}) = 30/M$

$$p_2/p^\circ = x(\text{water}) = 6/(6 + 30/M) = 2.9/p^\circ \dots (ii)$$

Dividing (i) by (ii):

$$2.8/2.9 = [5/(5 + 30/M)] / [6/(6 + 30/M)]$$

$$2.8 \times 6 \times (6 + 30/M) = 2.9 \times 5 \times (5 + 30/M) \text{ (cross multiply and simplify)}$$

$$16.8(6 + 30/M) = 14.5(5 + 30/M)$$

$$100.8 + 504/M = 72.5 + 435/M$$

$$69/M = 28.3 \rightarrow M = 69/28.3 \approx 23 \text{ g/mol}$$

Back-substitute into (i): $2.8/p^\circ = 5/(5 + 30/23) = 5/6.304 = 0.793$

$$p^\circ = 2.8/0.793 = 3.53 \text{ kPa}$$

✓ Molar mass = 23 g/mol; $p^\circ(\text{water}) = 3.53 \text{ kPa}$

Q 1.20 — Freezing Point Depression: Glucose vs Cane Sugar

Question: 5% cane sugar freezes at 271 K. Find freezing point of 5% glucose if pure water freezes at 273.15 K.

 Note: $\Delta T_f = K_f \times m$. Since K_f is same for water, and both solutions are 5% by mass, we compare molalities.

Molar mass: Cane sugar ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) = 342 g/mol; Glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) = 180 g/mol

In 100 g solution: 5 g solute, 95 g water.

$$m(\text{cane sugar}) = (5/342)/(0.095) = 0.1538 \text{ mol/kg}$$

$$m(\text{glucose}) = (5/180)/(0.095) = 0.2924 \text{ mol/kg}$$

$$\Delta T_f(\text{cane sugar}) = 273.15 - 271 = 2.15 \text{ K}$$

$$K_f = \Delta T_f/m = 2.15/0.1538 = 13.98 \text{ K}\cdot\text{kg/mol} \text{ (This seems high; let's proceed with the ratio method)}$$

$$\Delta T_f(\text{glucose})/\Delta T_f(\text{cane sugar}) = m(\text{glucose})/m(\text{cane sugar}) = 0.2924/0.1538 = 1.901$$

$$\Delta T_f(\text{glucose}) = 1.901 \times 2.15 = 4.09 \text{ K}$$

$$T_f(\text{glucose}) = 273.15 - 4.09 = 269.06 \text{ K}$$

✓ Freezing point of 5% glucose solution $\approx 269.1 \text{ K}$

Q 1.21 — Atomic Masses of A and B (Freezing Point)

Question: AB_2 and AB_4 dissolved in benzene. 1g AB_2 lowers T_f by 2.3 K; 1g AB_4 lowers T_f by 1.3 K. $K_f = 5.1 \text{ K}\cdot\text{kg/mol}$. Find atomic masses of A and B.

$K_f \times w \times 1000 = M \times T_f \times W$ (using $\Delta T_f = K_f \times w/(M \times W)$ where W = mass of solvent in g)

$$M(\text{AB}_2) = (K_f \times w \times 1000)/(\Delta T_f \times W) = (5.1 \times 1 \times 1000)/(2.3 \times 20) = 5100/46 = 110.87 \text{ g/mol}$$

$$M(\text{AB}_4) = (5.1 \times 1 \times 1000)/(1.3 \times 20) = 5100/26 = 196.15 \text{ g/mol}$$

Let $A = a$, $B = b$. Then:

$$a + 2b = 110.87 \quad \dots(1)$$

$$a + 4b = 196.15 \quad \dots(2)$$

$$(2) - (1): 2b = 85.28 \rightarrow b = 42.64 \approx 42.7 \text{ g/mol} \quad (\text{close to phosphorus } \sim 31, \text{ but approximately this value})$$

$$a = 110.87 - 2(42.64) = 110.87 - 85.28 = 25.59 \approx 25.6 \text{ g/mol}$$

✓ Atomic mass of $A \approx 25.6 \text{ g/mol}$ (similar to Mg); $B \approx 42.7 \text{ g/mol}$

Q 1.22 — Osmotic Pressure Concentration

Question: 36 g glucose in 1L has $\pi = 4.98 \text{ bar}$. If $\pi = 1.52 \text{ bar}$ (same T), find concentration.

 Note: $\pi = CRT$. Since R and T are constant, $\pi \propto C$ (concentration).

$$C_1 = 36/180 / 1\text{L} = 0.2 \text{ mol/L}; \pi_1 = 4.98 \text{ bar}$$

$$\pi/C = RT = \text{constant} \rightarrow C_2 = C_1 \times (\pi_2/\pi_1) = 0.2 \times (1.52/4.98) = 0.061 \text{ mol/L}$$

✓ Concentration = 0.061 mol/L

Q 1.23 — Intermolecular Interactions

Question: Identify the most important intermolecular interaction in: (i) n-hexane + n-octane, (ii) I_2 + CCl_4 , (iii) NaClO_4 + water, (iv) methanol + acetone, (v) acetonitrile + acetone.

(i) n-hexane + n-octane: London dispersion (van der Waals) forces — both are non-polar hydrocarbons.

(ii) I_2 + CCl_4 : London dispersion forces — both are non-polar; I_2 is large and polarisable.

(iii) NaClO_4 + water: Ion-dipole interactions — Na^+ and ClO_4^- ions interact strongly with polar water molecules.

(iv) Methanol + acetone: Hydrogen bonding — O–H of methanol forms H-bond with C=O of acetone.

(v) Acetonitrile + acetone: Dipole-dipole interactions — both are polar molecules with similar dipole moments.

Q 1.24 — Solubility in n-Octane

Question: Arrange in increasing solubility in n-octane: Cyclohexane, KCl, CH_3OH , CH_3CN .

 Note: n-Octane is a non-polar hydrocarbon. Like dissolves like: non-polar solutes dissolve better.

KCl is ionic — very insoluble in non-polar octane.

CH_3OH has strong hydrogen bonding with itself — very poorly soluble in octane.

CH_3CN is polar but no H-bonding — slightly soluble.

Cyclohexane is non-polar — most soluble (like dissolves like).

✓ Increasing solubility: $\text{KCl} < \text{CH}_3\text{OH} < \text{CH}_3\text{CN} < \text{Cyclohexane}$

Q 1.25 — Solubility in Water

Question: Classify as insoluble, partially soluble, or highly soluble in water: (i) phenol, (ii) toluene, (iii) formic acid, (iv) ethylene glycol, (v) chloroform, (vi) pentanol.

- (i) Phenol — Partially soluble (has -OH for H-bonding but also non-polar benzene ring).
- (ii) Toluene — Insoluble (non-polar hydrocarbon; no H-bonding with water).
- (iii) Formic acid (HCOOH) — Highly soluble (forms strong H-bonds with water; completely miscible).
- (iv) Ethylene glycol (HO-CH₂-CH₂-OH) — Highly soluble (two -OH groups form extensive H-bonds).
- (v) Chloroform (CHCl₃) — Partially soluble (slightly polar but no H-bond donor ability to water).
- (vi) Pentanol (C₅H₁₁OH) — Partially soluble (one -OH but long non-polar chain dominates).

Q 1.26 — Molarity of Na⁺ in Lake Water

Question: Density of lake water = 1.25 g/mL; contains 92 g Na⁺ per kg water. Find molarity of Na⁺.

In 1 kg (1000 g) water, Na⁺ = 92 g; total solution = 1000 + 92 = 1092 g

Volume of solution = 1092 / 1.25 = 873.6 mL = 0.8736 L

Moles of Na⁺ = 92 / 23 = 4 mol

Molarity = 4 / 0.8736 = 4.58 mol/L

✓ Molarity of Na⁺ = 4.58 M

Q 1.27 — Maximum Molarity from K_{sp} of CuS

Question: K_{sp}(CuS) = 6 × 10⁻¹⁶. Find maximum molarity of CuS in aqueous solution.

$\text{CuS} \rightleftharpoons \text{Cu}^{2+} + \text{S}^{2-}$

Let solubility = s mol/L. Then [Cu²⁺] = s, [S²⁻] = s

$K_{sp} = s \times s = s^2 = 6 \times 10^{-16}$

$s = \sqrt{6 \times 10^{-16}} = 2.45 \times 10^{-8} \text{ mol/L}$

✓ Maximum molarity of CuS = 2.45 × 10⁻⁸ mol/L

Q 1.28 — Mass % of Aspirin in Acetonitrile

Question: 6.5 g C₉H₈O₄ (aspirin) dissolved in 450 g acetonitrile. Find mass %.

Total mass = 6.5 + 450 = 456.5 g

Mass % = (6.5 / 456.5) × 100 = 1.424%

✓ Mass percentage of aspirin = 1.42%

Q 1.29 — Mass of Nalorphene Solution

Question: Dose of nalorphene = 1.5 mg. Concentration = 1.5 × 10⁻³ m. Find mass of solution required.

Molar mass of nalorphene (C₁₉H₂₁NO₃) = 19×12 + 21 + 14 + 3×16 = 228 + 21 + 14 + 48 = 311 g/mol

1.5 × 10⁻³ m means 1.5 × 10⁻³ mol nalorphene per 1 kg water

Mass of nalorphene in 1 kg water = 1.5 × 10⁻³ × 311 = 0.4665 g = 466.5 mg

We need 1.5 mg dose. Ratio = 1.5 mg / 466.5 mg = 3.214 × 10⁻³

Mass of water needed = 3.214 × 10⁻³ × 1000 g = 3.214 g