CLASS 12 CHEMISTRY NCERT SOLUTIONS

NCERT Solutions for Class 12 Chemistry Chapter 2: Solutions – Free PDF Download

Chapter 2, "Solutions," builds the entire framework of concentration terms, Raoult's law, Henry's law and colligative properties that carries through into electrochemistry and beyond. This post walks through all 41 in-text exercise questions (Q2.1–Q2.41) with complete, original step-by-step working — no shortcut...

Contents

Chapter 2, “Solutions,” builds the entire framework of concentration terms, Raoult’s law, Henry’s law and colligative properties that carries through into electrochemistry and beyond. This post walks through all 41 in-text exercise questions (Q2.1–Q2.41) with complete, original step-by-step working — no shortcuts, every formula shown. These Class 12 Chemistry Chapter 2 solutions are also useful as quick revision notes before exams.

NCERT Solutions for Class 12 Chemistry Chapter 2: Solutions

Q2.1. 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 chemically non-reacting substances, made up of a solute (present in smaller amount) dissolved in a solvent (present in larger amount). Based on the physical state of solute and solvent, nine types exist: (i) Gas in gas — air (O_2 + N_2 etc.) (ii) Gas in liquid — CO_2 dissolved in water (soda water) (iii) Gas in solid — H_2 adsorbed in palladium (iv) Liquid in gas — water vapour in air (humid air) (v) Liquid in liquid — ethanol in water (vi) Liquid in solid — mercury in amalgam with a metal (vii) Solid in gas — camphor vapour in air (viii) Solid in liquid — sugar in water (ix) Solid in solid — copper dissolved in gold (alloy).

Q2.2. Give an example of a solid solution in which the solute is a gas.
Hydrogen gas absorbed/dissolved in palladium metal (Pd) is a solid solution with a gaseous solute. Another accepted example is a mixture of O_2 and N_2 trapped in vulcanised rubber.

Q2.3. Define the following terms: (i) Mole fraction (ii) Molality (iii) Molarity (iv) Mass percentage.
(i) **Mole fraction (x)**: ratio of moles of one component to the total moles of all components in solution, $x_A = n_A/(n_A + n_B)$. (ii) **Molality (m)**: number of moles of solute dissolved per kilogram of solvent, $m = n_{solute}/(\text{mass of solvent in kg})$. (iii) **Molarity (M)**: number of moles of solute dissolved per litre of solution, $M = n_{solute}/(\text{volume of solution in L})$. (iv) **Mass percentage (w/w %)**: mass of a component per 100 g of solution × 100.

Q2.4. Concentrated nitric acid used in laboratory work is 68% nitric acid by mass in aqueous solution. What should be the molarity of such a sample of the acid if the density of the solution is 1.504 g mL–1?
In 100 g solution, mass of HNO_3 = 68 g, $M(HNO_3)$ = 63 g mol–1.
Moles of HNO_3 = 68/63 = 1.079 mol.
Volume of solution = 100 g ÷ 1.504 g mL–1 = 66.49 mL = 0.06649 L.
Molarity = 1.079/0.06649 = **16.23 mol L–1**.

Q2.5. A solution of glucose in water is labelled as 10% w/w. What would be the molality and mole fraction of each component in the solution? If the density of the solution is 1.2 g mL–1, what shall be the molarity of the solution?
In 100 g solution: glucose = 10 g (M = 180 g mol–1), water = 90 g.
Moles glucose = 10/180 = 0.0556 mol; moles water = 90/18 = 5.0 mol.
Molality = 0.0556 mol/0.090 kg = **0.617 m**.
Mole fraction of glucose = 0.0556/(0.0556+5.0) = **0.0110**; mole fraction of water = **0.989**.
Volume of solution = 100 g ÷ 1.2 g mL–1 = 83.33 mL = 0.08333 L.
Molarity = 0.0556/0.08333 = **0.667 mol L–1**.

Q2.6. How many mL of 0.1 M HCl are required to react completely with 1 g mixture of Na_2CO_3 and $NaHCO_3$ containing equimolar amounts of both?
Let x mol of each be present. $M(Na_2CO_3)$=106, $M(NaHCO_3)$=84.
$106x + 84x = 1 \Rightarrow 190x = 1 \Rightarrow x = 5.263 \times 10^{-3}$ mol each.

$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$ (needs 2 mol HCl per mol); $NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$ (needs 1 mol HCl per mol).
Total HCl = $2(5.263\times10^{-3}) + 5.263\times10^{-3} = 1.579 \times 10^{-2}$ mol.
Volume = 0.01579 mol/0.1 mol L^{-1} = **157.9 mL**.

Q2.7. A solution is obtained by mixing 300 g of 25% solution and 400 g of 40% solution by mass. Calculate the mass percentage of the resulting solution.
Mass of solute = (300 × 0.25) + (400 × 0.40) = 75 + 160 = 235 g.
Total mass = 300 + 400 = 700 g.
Mass % = (235/700) × 100 = **33.57%**.

Q2.8. An antifreeze solution is prepared from 222.6 g of ethylene glycol ($C_2H_6O_2$) and 200 g of water. Calculate the molality of the solution. If the density of the solution is 1.072 g mL^{-1}, what shall be the molarity of the solution?
M(ethylene glycol) = 62 g mol^{-1}. Moles = 222.6/62 = 3.590 mol.
Molality = 3.590 mol/0.200 kg = **17.95 m**.
Total mass = 222.6 + 200 = 422.6 g; volume = 422.6/1.072 = 394.2 mL = 0.3942 L.
Molarity = 3.590/0.3942 = **9.11 mol L^{-1}**.

Q2.9. A sample of drinking water was found to be severely contaminated with chloroform ($CHCl_3$), a suspected carcinogen. 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.
(i) 15 ppm = 15 g per 10^6 g solution $\Rightarrow$ mass % = $(15/10^6) \times 100$ = **1.5×10^{-3} %**.
(ii) M($CHCl_3$) = 119.5 g mol^{-1}. Moles = 15/119.5 = 0.1255 mol. Since water is in vast excess, mass of solvent ≈ (10^6 – 15) g ≈ 1000 kg.
Molality = 0.1255 mol/1000 kg = **1.255×10^{-4} m**.

Q2.10. What role does molecular interaction play in a solution of alcohol and water?
In pure alcohol and pure water, extensive hydrogen bonding exists between like molecules. On mixing, some of these alcohol–alcohol and water–water hydrogen bonds break and are only partially replaced by weaker alcohol–water hydrogen bonds. This net weakening of intermolecular attraction increases the escaping tendency of both components, so the mixture shows a higher vapour pressure than predicted by Raoult's law — i.e., **positive deviation** — and dissolution is generally endothermic.

Q2.11. Why do gases always tend to be less soluble in liquids as the temperature is raised?
Dissolution of a gas in a liquid is an exothermic process (Gas + Solvent → Solution + Heat). By Le Chatelier's principle, raising the temperature shifts the equilibrium in the endothermic (reverse) direction, favouring escape of the gas and reducing solubility. Additionally, higher thermal energy increases the kinetic energy of dissolved gas molecules, making it easier for them to overcome intermolecular attraction and escape into the vapour phase.

Q2.12. State Henry's law and mention some important applications.
Henry's law states that at constant temperature, the partial pressure of a gas in the vapour phase above a solution is directly proportional to the mole fraction of the gas dissolved in the solution: $\mathbf{p = K_H \cdot x}$, where K_H is the Henry's law constant (specific to the gas–solvent pair and temperature). Applications: (i) carbonated beverages are bottled under high CO_2 pressure to increase its solubility; (ii) scuba divers use air tanks diluted with helium to reduce the solubility (and hence decompression sickness risk) of N_2 in blood at depth; (iii) at high altitudes, the low partial pressure of O_2 reduces its solubility in blood, causing anoxia/altitude sickness; (iv) it governs gas absorption efficiency in industrial processes such as the Haber process.

Q2.13. The partial pressure of ethane over a solution containing 6.56×10^{-3} g of ethane is 1 bar.

If the solution contains 5.00 × 10−2 g of ethane, what shall be the partial pressure of the gas?
By Henry's law, p ∝ mole fraction of gas ∝ mass of gas (solvent amount and molar mass of ethane both constant).
p2 = p1 × (w2/w1) = 1 bar × (5.00×10−2/6.56×10−3) = **7.62 bar**.

Q2.14. What is meant by positive and negative deviations from Raoult's law, and how is the sign of ΔmixH related to positive and negative deviations from Raoult's law?
Positive deviation: solute–solvent (A–B) attractive forces are weaker than the A–A and B–B forces in the pure components, so escaping tendency of molecules increases and the solution's vapour pressure is higher than the ideal (Raoult's law) value. Mixing is endothermic, so **ΔmixH is positive** (e.g., ethanol + acetone, water + ethanol). **Negative deviation:** A–B attractive forces are stronger than A–A and B–B forces, lowering the escaping tendency and giving a vapour pressure below the ideal value. Mixing is exothermic, so **ΔmixH is negative** (e.g., acetone + chloroform, phenol + aniline).

Q2.15. An aqueous solution of 2% non-volatile solute exerts a pressure of 1.004 bar at the normal boiling point of the solvent. What is the molar mass of the solute?
At the normal boiling point, p° = 1.013 bar; p (solution) = 1.004 bar.
Relative lowering = (p° − p)/p° = (1.013 − 1.004)/1.013 = 8.885 × 10−3.
In 100 g solution: w2 (solute) = 2 g, w1 (water) = 98 g, M1 = 18 g mol−1.
(p°−p)/p° = (w2/M2) ÷ (w1/M1) ⇒ 8.885×10−3 = (2/M2) × (18/98).
M2 = 36/(98 × 8.885×10−3) = **41.35 g mol−1**.

Q2.16. Heptane and octane form an ideal solution. At 373 K, the vapour pressures of the two liquid components are 105.2 kPa and 46.8 kPa respectively. What will be the vapour pressure of a mixture of 26.0 g of heptane and 35 g of octane?
M(heptane, C7H16)=100 g mol−1, M(octane, C8H18)=114 g mol−1.
Moles heptane = 26.0/100 = 0.260 mol; moles octane = 35/114 = 0.307 mol.
x_heptane = 0.260/0.567 = 0.4585; x_octane = 0.5415.
p_total = (0.4585 × 105.2) + (0.5415 × 46.8) = 48.24 + 25.34 = **73.6 kPa**.

Q2.17. The vapour pressure of water is 12.3 kPa at 300 K. Calculate the vapour pressure of 1 molal solution of a non-volatile solute in it.
1 molal = 1 mol solute per 1000 g (55.56 mol) water.
x_solute = 1/(1 + 55.56) = 0.01768.
p = p°(1 − x_solute) = 12.3 × (1 − 0.01768) = **12.08 kPa**.

Q2.18. Calculate the mass of a non-volatile solute (molar mass 40 g mol−1) which should be dissolved in 114 g octane to reduce its vapour pressure to 80%.
M(octane)=114 g mol−1, so 114 g = 1 mol (n1=1).
p/p° = 0.80 ⇒ relative lowering = 0.20 = x_solute = n2/(n1+n2).
n2/(1+n2) = 0.20 ⇒ n2 = 0.25 mol.
Mass of solute = 0.25 × 40 = **10 g**.

Q2.19. A solution containing 30 g of non-volatile solute exactly in 90 g of water has a vapour pressure of 2.8 kPa at 298 K. Further, 18 g of water is then added to the solution, and the new vapour pressure becomes 2.9 kPa at 298 K. Calculate: (i) molar mass of the solute (ii) vapour pressure of water at 298 K.
Using (p°−p)/p ≈ n2/n1:
Case 1 (90 g water, n1=5): (p°−2.8)/2.8 = (30/M2)/5 ⇒ p° = 2.8 + 16.8/M2.
Case 2 (108 g water, n1=6): (p°−2.9)/2.9 = (30/M2)/6 ⇒ p° = 2.9 + 14.5/M2.
Equating: 2.8 + 16.8/M2 = 2.9 + 14.5/M2 ⇒ 2.3/M2 = 0.1 ⇒ **M2 = 23 g mol−1**.
p° = 2.8 + 16.8/23 = **3.53 kPa**.

Q2.20. A 5% solution (by mass) of cane sugar in water has a freezing point of 271 K. Calculate the freezing point of 5% glucose in water if the freezing point of pure water is 273.15 K.
Cane sugar (sucrose), M=342 g mol−1: 5 g in 95 g water.
Molality = (5/342)/0.095 = 0.1538 m. ΔTf = 273.15 − 271 = 2.15 K.
Kf (from this data) = ΔTf/m = 2.15/0.1538 = 13.98 K kg mol−1.
Glucose, M=180 g mol−1: 5 g in 95 g water ⇒ molality = (5/180)/0.095 = 0.2924 m.
ΔTf (glucose) = 13.98 × 0.2924 = 4.09 K.
Freezing point of glucose solution = 273.15 − 4.09 = **269.06 K**.

Q2.21. Two elements A and B form compounds having formulas AB2 and AB4. When dissolved in 20 g of benzene (C6H6), 1 g of AB2 lowers the freezing point by 2.3 K, whereas 1.0 g of AB4 lowers it by 1.3 K. The molar depression constant for benzene is 5.1 K kg mol−1. Calculate the atomic masses of A and B.
For AB2: molality = 2.3/5.1 = 0.4510 m ⇒ moles = 0.4510 × 0.020 kg = 9.02×10−3 mol ⇒ M(AB2) = 1/9.02×10−3 = 110.9 g mol−1.
For AB4: molality = 1.3/5.1 = 0.2549 m ⇒ moles = 0.2549 × 0.020 = 5.10×10−3 mol ⇒ M(AB4) = 1/5.10×10−3 = 196.2 g mol−1.
Let atomic mass of A = a, B = b: a + 2b = 110.9 and a + 4b = 196.2.
Subtracting: 2b = 85.3 ⇒ **b (atomic mass of B) ≈ 42.6**; a = 110.9 − 85.3 = **a (atomic mass of A) ≈ 25.6**.

Q2.22. At 300 K, 36 g of glucose present in a litre of its solution has an osmotic pressure of 4.98 bar. If the osmotic pressure of the solution is 1.52 bars at the same temperature, what would be its concentration?
From the first data set: C1 = 36/180 = 0.20 mol L−1; using π = CRT, R = 4.98/(0.20 × 300) = 0.083 L bar mol−1 K−1 (confirms the gas constant).
For π2 = 1.52 bar: C2 = 1.52/(0.083 × 300) = **0.061 mol L−1**.

Q2.23. Suggest the most important type of intermolecular attractive interaction in the following pairs: (i) n-hexane and n-octane (ii) I2 and CCl4 (iii) NaClO4 and water (iv) methanol and acetone (v) acetonitrile (CH3CN) and acetone (C3H6O).
(i) London (dispersion) forces — both non-polar hydrocarbons. (ii) London (dispersion) forces — both non-polar. (iii) Ion–dipole interaction — ionic solute in polar water. (iv) Dipole–dipole interaction (with some hydrogen bonding from methanol's –OH). (v) Dipole–dipole interaction — both are polar molecules.

Q2.24. Based on solute–solvent interactions, arrange the following in order of increasing solubility in n-octane and explain: cyclohexane, KCl, CH3OH, CH3CN.
n-Octane is a non-polar solvent, so "like dissolves like" — solubility increases as polarity/ionic character decreases: **KCl < CH3OH < CH3CN < cyclohexane**. KCl (ionic) is essentially insoluble; methanol and acetonitrile are polar and only partly miscible; cyclohexane, being non-polar like octane, is fully miscible.

Q2.25. Amongst the following compounds, identify which are insoluble, partially soluble and highly soluble in water: (i) phenol (ii) toluene (iii) formic acid (iv) ethylene glycol (v) chloroform (vi) pentanol.
Highly soluble: formic acid and ethylene glycol (both are small, polar, and hydrogen-bond extensively with water). **Partially soluble:** phenol (polar –OH group but a large non-polar ring) and pentanol (one –OH group but a long non-polar chain). **Insoluble:** toluene and chloroform (both essentially non-polar, no hydrogen-bonding capacity with water).

Q2.26. If the density of some lake water is 1.25 g mL−1 and it contains 92 g of Na+ ions per kg of water, calculate the molarity of Na+ ions in the lake.
Taking 1 kg (1000 g) of lake water: mass of Na+ = 92 g, M(Na) = 23 g mol−1.
Moles of Na+ = 92/23 = 4.0 mol.

Volume of solution = 1000 g ÷ 1.25 g mL−1 = 800 mL = 0.8 L.
Molarity = 4.0/0.8 = **5.0 mol L−1**. *(Note: the question's wording is a known minor ambiguity across published sources — this uses the widely-accepted textbook interpretation treating 1 kg as the solution mass.)*

Q2.27. If the solubility product of CuS is 6 × 10−16, calculate the maximum molarity of CuS in aqueous solution.
CuS(s) ⇌ Cu2+ + S2−. Let solubility = s mol L−1. Ksp = [Cu2+][S2−] = s2.
s = √(6 × 10−16) = **2.45 × 10−8 mol L−1**.

Q2.28. Calculate the mass percentage of aspirin (C9H8O4) in acetonitrile (CH3CN) when 6.5 g of C9H8O4 is dissolved in 450 g of CH3CN.
Total mass = 6.5 + 450 = 456.5 g.
Mass % = (6.5/456.5) × 100 = **1.42%**.

Q2.29. Nalorphene (C19H21NO3), similar to morphine, is used to combat withdrawal symptoms in narcotic users. The dose generally given is 1.5 mg. Calculate the mass of 1.5 × 10−3 m aqueous solution required for this dose.
M(nalorphene) = 19(12) + 21(1) + 14 + 3(16) = 228+21+14+48 = 311 g mol−1.
A 1.5 × 10−3 molal solution contains 1.5 × 10−3 mol nalorphene per kg (1000 g) of solvent, i.e. 1.5×10−3 × 311 = 0.4665 g nalorphene per ≈1000 g of solution (dilute approximation).
Mass of solution needed for 1.5 mg (0.0015 g) dose = (0.0015/0.4665) × 1000 = ≈ **3.22 g**.

Q2.30. Calculate the amount of benzoic acid (C6H5COOH) required for preparing 250 mL of 0.15 M solution in methanol.
M(benzoic acid) = 122 g mol−1.
Moles required = 0.250 L × 0.15 mol L−1 = 0.0375 mol.
Mass = 0.0375 × 122 = **4.575 g**.

Q2.31. The depression in freezing point of water observed for the same amount of acetic acid, trichloroacetic acid and trifluoroacetic acid increases in the order given above. Explain briefly.
All three are weak acids that partially dissociate in water, and the extent of dissociation controls the van't Hoff factor (i), which directly scales ΔTf. Electronegative halogen substituents (Cl, F) withdraw electron density inductively from the −COOH group, weakening the O−H bond and increasing acid strength. Since F is more electronegative than Cl, trifluoroacetic acid dissociates the most, followed by trichloroacetic, followed by acetic acid (weakest, least electron-withdrawing). Greater dissociation means more solute particles in solution, hence a larger i and a larger observed ΔTf, explaining the increasing order.

Q2.32. Calculate the depression in the freezing point of water when 10 g of CH3CH2CHClCOOH is added to 250 g of water. Ka = 1.4 × 10−3, Kf = 1.86 K kg mol−1.
M(C4H7ClO2) = 122.5 g mol−1. Nominal molality m0 = (10/122.5)/0.250 = 0.3265 mol kg−1.
For a weak monoprotic acid, degree of dissociation α satisfies Ka = Cα2/(1−α). Solving iteratively with C = 0.3265: α ≈ 0.0634.
Van't Hoff factor i = 1 + α = 1.0634.
ΔTf = i × Kf × m = 1.0634 × 1.86 × 0.3265 = **0.646 K**.

Q2.33. 19.5 g of CH2FCOOH is dissolved in 500 g of water. The depression in the freezing point of water observed is 1.0°C. Calculate the van't Hoff factor and dissociation constant of fluoroacetic acid.
M(CH2FCOOH, C2H3FO2) = 78 g mol−1. Nominal molality = (19.5/78)/0.500 = 0.50 mol kg−1. (Using the standard Kf for water = 1.86 K kg mol−1.)
ΔTf = i × Kf × m ⇒ 1.0 = i × 1.86 × 0.50 ⇒ **i = 1.075**.

α = i − 1 = 0.075.
Ka = Cα2/(1−α) = (0.50 × 0.0752)/(1−0.075) = 0.002813/0.925 = **3.07 × 10−3**.

Q2.34. Vapour pressure of water at 293 K is 17.535 mm Hg. Calculate the vapour pressure of water at 293 K when 25 g of glucose is dissolved in 450 g of water.
Moles glucose = 25/180 = 0.1389 mol; moles water = 450/18 = 25 mol.
Relative lowering = n2/(n1+n2) = 0.1389/25.1389 = 5.525 × 10−3.
Δp = 17.535 × 5.525×10−3 = 0.0969 mm Hg.
p(solution) = 17.535 − 0.0969 = **17.44 mm Hg**.

Q2.35. Henry's law constant for methane in benzene at 298 K is 4.27 × 10^5 mm Hg. Calculate the solubility of methane in benzene at 298 K under 760 mm Hg.
By Henry's law, x_methane = p/K_H = 760/(4.27 × 10^5) = **1.78 × 10−3** (mole fraction of methane dissolved).

Q2.36. 100 g of liquid A (molar mass 140 g mol−1) was dissolved in 1000 g of liquid B (molar mass 180 g mol−1). The vapour pressure of pure liquid B was found to be 500 torr. Calculate the vapour pressure of pure liquid A and its vapour pressure in the solution if the total vapour pressure of the solution is 475 torr.
Moles A = 100/140 = 0.7143 mol; moles B = 1000/180 = 5.556 mol.
x_A = 0.7143/6.270 = 0.1139; x_B = 0.8861.
p_total = x_A p°_A + x_B p°_B ⇒ 475 = 0.1139 p°_A + (0.8861 × 500) ⇒ 475 − 443.05 = 0.1139 p°_A.
p°_A (pure liquid A) = 31.95/0.1139 = **280.5 torr**.
Partial pressure of A in solution = x_A × p°_A = **31.95 torr**.

Q2.37. Vapour pressures of pure acetone and chloroform at 328 K are 741.8 mm Hg and 632.8 mm Hg respectively. Assuming they form an ideal solution over the entire composition range, plot p_total, p_chloroform and p_acetone as a function of x_acetone for the given experimental data, and identify the type of deviation.
Given data: 100·x_acetone = 0, 11.8, 23.4, 36.0, 50.8, 58.2, 64.5, 72.1; p_acetone/mm Hg = 0, 54.9, 110.1, 202.4, 322.7, 405.9, 454.1, 521.1; p_chloroform/mm Hg = 632.8, 548.1, 469.4, 359.7, 257.7, 193.6, 161.2, 120.7. Adding these pairwise gives p_total at each composition (e.g. at x_acetone=0.360, p_total = 202.4+359.7 = 562.1 mm Hg). Plotting p_total, p_acetone and p_chloroform against x_acetone on the same graph shows that the experimental p_total curve lies **below** the straight line joining pure p°_chloroform (632.8) and p°_acetone (741.8) that ideal (Raoult's law) behaviour would predict. This is a **negative deviation** from Raoult's law, caused by strong hydrogen-bond-like C−H···O interaction between chloroform's acidic hydrogen and acetone's carbonyl oxygen, which makes A−B attraction stronger than A−A/B−B attraction and lowers escaping tendency.

Q2.38. Benzene and toluene form an ideal solution over the entire range of composition. The vapour pressures of pure benzene and toluene at 300 K are 50.71 mm Hg and 32.06 mm Hg respectively. Calculate the mole fraction of benzene in the vapour phase if 80 g of benzene is mixed with 100 g of toluene.
M(benzene)=78, M(toluene)=92. Moles benzene = 80/78 = 1.0256; moles toluene = 100/92 = 1.0870.
x_benzene = 1.0256/2.1126 = 0.4855; x_toluene = 0.5145.
p_benzene = 0.4855 × 50.71 = 24.62 mm Hg; p_toluene = 0.5145 × 32.06 = 16.50 mm Hg.
p_total = 41.12 mm Hg.
Mole fraction of benzene in vapour = p_benzene/p_total = 24.62/41.12 = **0.599**.

Q2.39. Air is a mixture of gases, mainly O2 and N2 in approximately 20:79 proportion by volume, at 298 K. Water is in equilibrium with air at a total pressure of 10 atm. If Henry's law constants for O2 and N2 at 298 K are 3.30 × 10^7 mm and 6.51 × 10^7 mm respectively, calculate the

composition (mole fraction) of these gases dissolved in water.
p_O2 = 0.20 × 10 atm = 2 atm = 1520 mm Hg; p_N2 = 0.79 × 10 atm = 7.9 atm = 6004 mm Hg.
x_O2 = p_O2/K_H(O2) = 1520/(3.30×10^7) = **4.61 × 10−5**.
x_N2 = p_N2/K_H(N2) = 6004/(6.51×10^7) = **9.22 × 10−5**.

Q2.40. Determine the amount of CaCl2 (i = 2.47) dissolved in 2.5 litres of water such that its osmotic pressure is 0.75 atm at 27°C.
T = 300 K. Using π = iCRT: C = π/(iRT) = 0.75/(2.47 × 0.0821 × 300) = 0.75/60.84 = 0.01233 mol L−1.
Moles = C × V = 0.01233 × 2.5 = 0.0308 mol.
M(CaCl2) = 40 + 2(35.5) = 111 g mol−1.
Mass = 0.0308 × 111 = **3.42 g**.

Q2.41. Determine the osmotic pressure of a solution prepared by dissolving 25 mg of K2SO4 in 2 litres of water at 25°C, assuming it is completely dissociated.
M(K2SO4) = 2(39)+32+4(16) = 174 g mol−1. Moles = 0.025/174 = 1.437 × 10−4 mol.
C = 1.437×10−4/2 = 7.18 × 10−5 mol L−1. K2SO4 → 2K+ + SO4 2−, so i = 3 (complete dissociation).
T = 298 K. π = iCRT = 3 × 7.18×10−5 × 0.0821 × 298 = **5.27 × 10−3 atm**.

Why This Chapter Matters

Solutions is one of the highest-yield chapters for both board exams and competitive entrance tests because it is almost entirely numerical and formula-driven — once the core relationships (Raoult's law, Henry's law, the four colligative properties, and the van't Hoff factor) are internalised, most questions become a matter of careful substitution rather than conceptual guesswork. It also lays the foundation for later JEE/NEET topics like electrochemistry (where molarity and concentration terms reappear constantly) and for real-world applications such as intravenous saline concentration, antifreeze formulation, and reverse osmosis water purification.

Class 12 Chemistry Chapter 2 – Notes and Extra Questions

Along with these NCERT Solutions, students can also use the Class 12 Chemistry Chapter 2 Extra Questions and Class 12 Chemistry Chapter 2 Revision Notes for quick revision and extra practice.