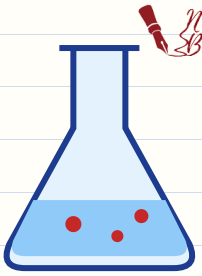
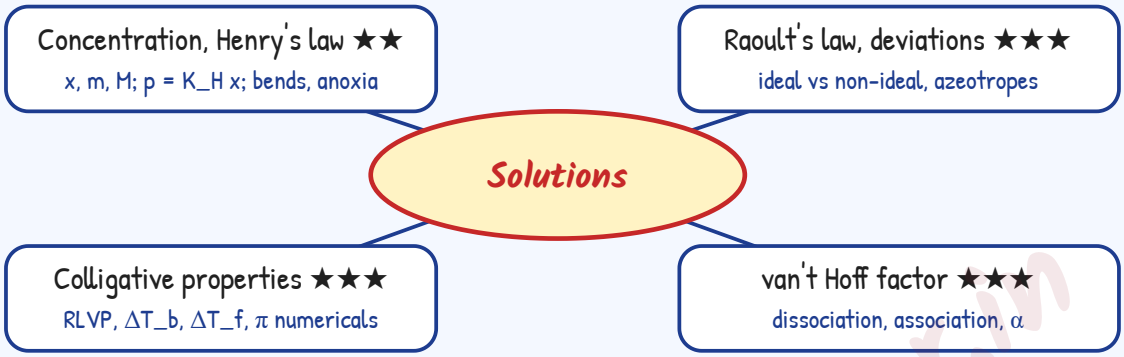
Class 12 · Chemistry · Unit 1

Chapter 1 Solutions



NCERT Chemistry Part I, Reprint 2026-27 · CBSE board 2027

Chapter map (★ = asked again and again in CBSE 2024–2026)



What is inside

1. Types of solutions; concentration terms (p. 2-3)
2. Solubility; Henry's law (p. 4)
3. Vapour pressure; Raoult's law; ideal and non-ideal solutions (p. 5-6)
4. Colligative properties (p. 7-8)
5. Abnormal molar masses; van't Hoff factor (p. 9)
6. All NCERT intext and exercise questions solved (p. 3-13)
7. CBSE PYQs 2024-2026, fully solved (p. 14-19)
8. Formula sheet; practice A-G, 20 MCQs, 6 A-R, answers, summary (p. 20-24)

Colour code used

red = key term green = definition
purple = sub-heading black = working
★ = high-frequency board topic
1 = mark the examiner gives for that step

★ **Exam tip:** Solutions carries **about 7 marks** in every CBSE paper 2024-2026: a **3-mark colligative numerical** (often with i), a **Henry's law or deviation** 2-marker, 1-2 MCQs or A-R, and often a 5-mark Section E question. All the board questions from the 16 papers are solved in full on pages 14-19.

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Based on NCERT Chemistry Part I (2026-27), Unit 1. Figures drawn by us (Ex 1.37 plotted from the book's data); all answers written by us. Board questions are quoted from the CBSE papers named in each tag; every answer checked against the CBSE marking schemes.

1.1 Types of solutions

Definition ✎ **Solution**: a homogeneous mixture of two or more components. The component present in the largest amount is the **solvent**; the others are **solutes**. A **binary** solution has two components.

Solute in solvent	Example
Gaseous solutions	
gas in gas	oxygen + nitrogen (air)
liquid in gas	chloroform mixed with nitrogen
solid in gas	camphor in nitrogen
Liquid solutions	
gas in liquid	oxygen in water
liquid in liquid	ethanol in water
solid in liquid	glucose in water
Solid solutions	
gas in solid	H_2 in palladium (NCERT Ex 1.2)
liquid in solid	amalgam of mercury with sodium
solid in solid	copper in gold

1.2 Expressing concentration

Term	Formula
mass %	$\frac{w_B}{w_{\text{solution}}} \times 100$
volume %	$\frac{V_B}{V_{\text{solution}}} \times 100$
mass by volume %	g of solute per 100 mL solution
ppm	$\frac{\text{parts of B}}{\text{total parts}} \times 10^6$
mole fraction	$x_B = \frac{n_B}{n_A + n_B}$, $\Sigma x = 1$
molarity M	mol of solute per **litre of solution**
molality m	mol of solute per **kg of solvent**

✗ **Don't do this:** Molarity changes with temperature (volume does); mass %, ppm, mole fraction and molality do not. That is why colligative-property formulas use **molality**.

NCERT Example 1.1 Mole fraction of ethylene glycol in a 20% (by mass) solution.

> **Ans.** In 100 g: 20 g glycol (20/62 = 0.322 mol), 80 g water (4.444 mol). x(glycol) = 0.322/4.766 = **0.068**; x(water) = **0.932**.

NCERT Example 1.2 Molarity of 5 g NaOH in 450 mL solution.

> **Ans.** 0.125 mol/0.450 L = **0.278 M**.

NCERT Example 1.3 Molality of 2.5 g ethanoic acid in 75 g benzene.

> **Ans.** 2.5/60 = 0.0417 mol; m = 0.0417/0.075 = **0.556 mol kg⁻¹**.

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Intext questions, solved

1.1 Mass % of benzene and CCl_4 when 22 g of benzene is dissolved in 122 g of CCl_4.

> **Ans.** Total = 144 g. Benzene = 22/144 × 100 = **15.28%**; CCl_4 = 122/144 × 100 = **84.72%**.

1.2 Mole fraction of benzene in a 30% (by mass) solution in CCl_4.

> **Ans.** In 100 g: benzene 30/78 = 0.385 mol, CCl_4 70/154 = 0.455 mol. x(benzene) = 0.385/0.840 = **0.459**; x(CCl_4) = 0.541.

1.3 Molarity of (a) 30 g of $Co(NO_3)_2{\cdot}6H_2O$ in 4.3 L of solution; (b) 30 mL of 0.5 M H_2SO_4 diluted to 500 mL.

> **Ans.** (a) M of $Co(NO_3)_2{\cdot}6H_2O$ = 291 g mol^{-1}: 30/291 = 0.103 mol; 0.103/4.3 = **0.024 M**. (b) $M_1V_1 = M_2V_2$: 0.5 × 30 = M_2 × 500, M_2 = **0.03 M**.

1.4 Mass of urea for 2.5 kg of 0.25 molal aqueous solution.

> **Ans.** 0.25 mol (15 g) of urea per 1000 g water: 1015 g of solution holds 15 g. In 2500 g: 15 × 2500/1015 = **36.95 g**.

1.5 Molality, molarity and mole fraction of KI in 20% (m/m) aqueous KI of density 1.202 g mL^{-1}.

> **Ans.** In 100 g: 20 g KI (20/166 = 0.120 mol), 80 g water (4.444 mol).
> (a) m = 0.120/0.080 = **1.51 mol kg^{-1}**.
> (b) V = 100/1.202 = 83.2 mL: M = 0.120/0.0832 = **1.45 M**.
> (c) x(KI) = 0.120/(0.120 + 4.444) = **0.0263**.

★ ***Exam tip:*** Converting between M and m: assume 1 L (or 100 g) of solution, find the masses with the density, then subtract the solute to get the solvent mass.

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1.3 Solubility; Henry's law ★★

Solubility: the maximum amount of a substance that dissolves in a given amount of solvent at a given temperature.

Solids in liquids: "like dissolves like" (polar in polar, non-polar in non-polar). At **saturation**, dissolution and crystallisation are in dynamic equilibrium. Solubility rises with T if dissolving is endothermic (Le Chatelier); pressure has no real effect.

Gases in liquids: solubility rises with pressure and falls with temperature (dissolving a gas is exothermic).

> ***Definition*** **Henry's law**: at constant temperature, the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above it. Equivalently, the partial pressure equals K_H times the mole fraction of the gas in solution:
>
> $$p = K_H\, x$$

- K_H depends on the gas and on T. A **larger K_H means lower solubility** (for the same p, x = p/K_H is smaller).
- K_H **increases** with temperature, so solubility falls as T rises.

Applications

- Soda water and soft drinks are sealed under high CO_2 pressure.
- **Bends** (scuba divers): at depth, high pressure dissolves more N_2 in blood; on rising quickly, the pressure falls and N_2 forms bubbles that block capillaries. Air tanks use He-diluted mixtures (11.7% He, 56.2% N_2, 32.1% O_2).
- **Anoxia** (high altitude): low partial pressure of O_2 means less O_2 in blood: weakness, unclear thinking.
- Aquatic life is more comfortable in **cold water**: more O_2 dissolves.

NCERT Example 1.4 Millimoles of N_2 dissolving in 1 L of water at 293 K; $p(N_2)$ = 0.987 bar, K_H = 76.48 kbar.

> **Ans.** $x = p/K_H = 0.987/76480 = 1.29 \times 10^{-5}$. 1 L of water = 55.5 mol: $n = x \times 55.5 = 7.16 \times 10^{-4}$ mol = **0.716 mmol**.

Intext questions

1.6 Solubility of H_2S in water at STP is 0.195 m. Henry's law constant?

> **Ans.** 0.195 mol in 1 kg (55.56 mol) water: $x = 0.195/(0.195 + 55.56) = 3.50 \times 10^{-3}$. At STP p = 0.987 bar: $K_H = 0.987/3.50 \times 10^{-3}$ = **282 bar**.

1.7 K_H for CO_2 in water = 1.67×10^8 Pa at 298 K. Quantity of CO_2 in 500 mL soda water packed under 2.5 atm.

> **Ans.** $p = 2.5 \times 1.01325 \times 10^5 = 2.533 \times 10^5$ Pa. $x = p/K_H = 1.517 \times 10^{-3}$. 500 mL water ≈ 27.78 mol: $n(CO_2) = x \times 27.78 = 0.0421$ mol = **1.85 g**.

Raoult's law as a special case of Henry's law: for a volatile component, $p = p^\circ x$ (Raoult) has the same form as $p = K_H x$ (Henry). Raoult's law is Henry's law with $K_H = p^\circ$ (CBSE 2025, 56/4/1).

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1.4 Vapour pressure of liquid solutions; Raoult's law ★★★

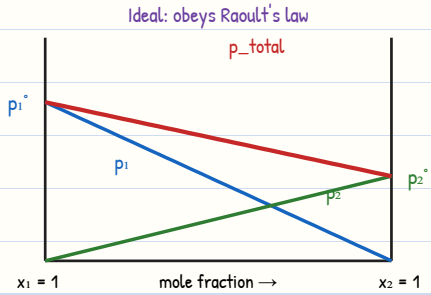
Definition ✎ **Raoult's law** (volatile components): the partial vapour pressure of each component is directly proportional to its mole fraction:

$$p_1 = p_1^\circ x_1, \quad p_2 = p_2^\circ x_2$$

Dalton's law: $p_{\text{total}} = p_1 + p_2 = p_1^\circ + (p_2^\circ - p_1^\circ)x_2$, a straight line.

Vapour composition: $y_1 = p_1/p_{\text{total}}$. The vapour is richer in the more volatile component.

Non-volatile solute: only the solvent has vapour: p = p₁°x₁ (the vapour pressure falls).



p₁ = p₁°x₁, p₂ = p₂°x₂; p_total = p₁° + (p₂° − p₁°)x₂ is a straight line

NCERT Example 1.5 p°(CHCl₃) = 200, p°(CH₂Cl₂) = 415 mm Hg at 298 K. (i) p of a mixture of 25.5 g CHCl₃ and 40 g CH₂Cl₂; (ii) vapour composition.

Ans. n(CH₂Cl₂) = 40/85 = 0.47; n(CHCl₃) = 25.5/119.5 = 0.213. x(CH₂Cl₂) = 0.688, x(CHCl₃) = 0.312.
(i) p = 200 + (415 − 200) × 0.688 = **347.9 mm Hg**.
(ii) p(CH₂Cl₂) = 0.688 × 415 = 285.5; y = 285.5/347.9 = **0.82**; y(CHCl₃) = **0.18**.

Intext 1.8 p°(A) = 450, p°(B) = 700 mm Hg at 350 K. Liquid composition if p_total = 600 mm Hg; vapour composition.

Ans. 600 = 450x_A + 700(1 − x_A) ⇒ 250x_A = 100: **x_A = 0.4, x_B = 0.6**.
p_A = 0.4 × 450 = 180; y_A = 180/600 = **0.30**, y_B = **0.70**.

Definition ✎ **Ideal solution**: obeys Raoult's law over the entire range of concentration; **Δ_mixH = 0** and **Δ_mixV = 0**; A–B interactions equal A–A and B–B. Examples: n-hexane + n-heptane, bromoethane + chloroethane, benzene + toluene.

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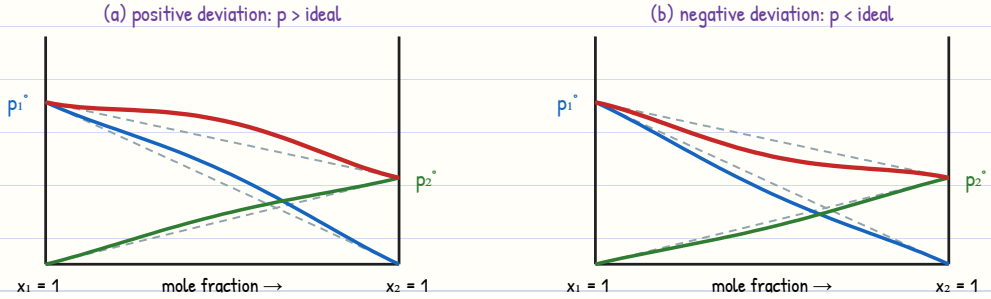
1.5 Non-ideal solutions and azeotropes ★★★



Dashed lines: Raoult's law. (a) A–B attraction weaker (ethanol + acetone): minimum-boiling azeotrope. (b) A–B attraction stronger (chloroform + acetone): maximum-boiling azeotrope

	Positive deviation	Negative deviation
vapour pressure	higher than Raoult's law predicts	lower than predicted
interactions	A–B **weaker** than A–A and B–B	A–B **stronger** than A–A and B–B
$\Delta_{mix}H$	> 0 (endothermic, cooling on mixing)	< 0 (exothermic, warming on mixing)
$\Delta_{mix}V$	> 0 (volume increases)	< 0 (volume decreases)
examples	ethanol + acetone; CS_2 + acetone; ethanol + water	chloroform + acetone; phenol + aniline; HNO_3 + water
azeotrope	**minimum-boiling** (about 95% ethanol + water)	**maximum-boiling** (68% HNO_3 + 32% water)

Why ethanol + acetone deviates positively: ethanol molecules are hydrogen-bonded to one another. Acetone breaks some of these H-bonds, and the new ethanol–acetone attractions are weaker, so molecules escape more easily: higher vapour pressure.

Why chloroform + acetone deviates negatively: the C–H of chloroform forms a hydrogen bond with the C=O oxygen of acetone, which is stronger than the original attractions: lower vapour pressure.

> ***Definition*** ✎ **Azeotropes**: binary mixtures with the **same composition in liquid and vapour**, which boil at a constant temperature; they cannot be separated by fractional distillation.

Ethanol cannot be distilled beyond about 95% purity (by volume) because that mixture is a minimum-boiling azeotrope.

NCERT Ex 1.14 What are positive and negative deviations from Raoult's law, and how is the sign of $\Delta_{mix}H$ related to them?

> ***Ans.*** Table above: positive deviation, $\Delta_{mix}H > 0$; negative deviation, $\Delta_{mix}H < 0$.

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1.6 Colligative properties ★★★

Colligative properties depend only on the **number of solute particles**, not on their nature. There are four: relative lowering of vapour pressure, elevation of boiling point, depression of freezing point and osmotic pressure.

1. Relative lowering of vapour pressure

For a non-volatile solute, $p_1 = p_1^\circ x_1 = p_1^\circ(1 - x_2)$, so

$$\frac{p_1^\circ - p_1}{p_1^\circ} = x_2 = \frac{n_2}{n_1 + n_2} \approx \frac{w_2 M_1}{M_2 w_1}$$

(the last form for dilute solutions, $n_2 \ll n_1$).

NCERT Example 1.6 p°(benzene) = 0.850 bar; 0.5 g of a solid in 39.0 g benzene gives 0.845 bar. M of the solid?

> **Ans.** (0.850 − 0.845)/0.850 = 0.5 × 78/(M_2 × 39): 0.00588 = 1/M_2: M_2 = **170 g mol⁻¹**.

2. Elevation of boiling point

The solution's vapour pressure is lower, so it must be heated more to reach 1 atm: it boils higher.

$$\Delta T_b = K_b\, m = \frac{K_b \times w_2 \times 1000}{M_2 \times w_1}$$

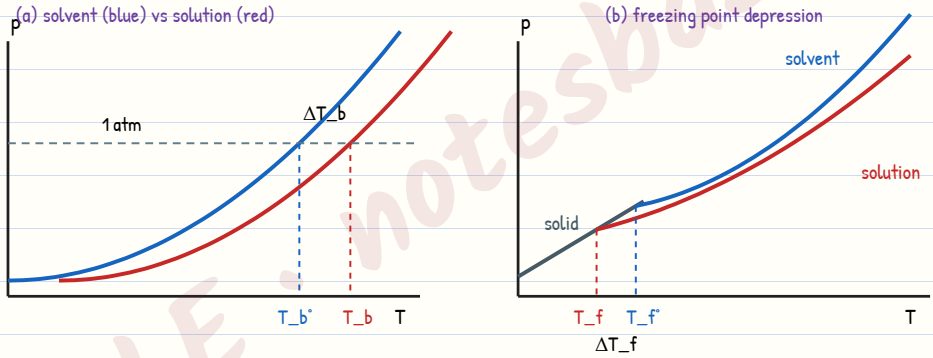
K_b = molal elevation (ebullioscopic) constant: the elevation for a 1 molal solution. Water 0.52, benzene 2.53 K kg mol⁻¹.

NCERT Example 1.7 18 g glucose in 1 kg water. Boiling point at 1.013 bar? (K_b = 0.52)

> **Ans.** m = 0.1; ΔT_b = 0.052 K: T_b = **373.202 K**.

NCERT Example 1.8 1.80 g solute in 90 g benzene raises the boiling point from 353.23 K to 354.11 K (K_b = 2.53). M?

> **Ans.** ΔT_b = 0.88 K: M_2 = 2.53 × 1.8 × 1000/(0.88 × 90) = **58 g mol⁻¹**.



A non-volatile solute lowers the vapour pressure: the solution boils higher (a) and freezes lower (b) than the pure solvent

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3. Depression of freezing point

At the freezing point, solid and liquid have equal vapour pressure. The solution's lower vapour pressure meets the solid's curve at a lower temperature.

$$\Delta T_f = K_f\, m = \frac{K_f \times w_2 \times 1000}{M_2 \times w_1}$$

Definition ✎ **Molal depression (cryoscopic) constant** K_f: the depression in freezing point of a 1 molal solution. Water 1.86, benzene 5.12 K kg mol⁻¹. It is related to the enthalpy of fusion:

$$K_f = \frac{R \times M_1 \times T_f^2}{1000 \times \Delta_{fus}H}$$

(similarly $K_b = \frac{RM_1T_b^2}{1000\,\Delta_{vap}H}$, with M₁ in g mol⁻¹).

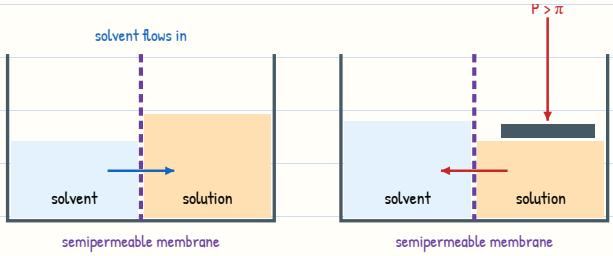
NCERT Example 1.9 45 g of ethylene glycol in 600 g water. ΔT_f and freezing point?

> **Ans.** n = 0.73 mol; m = 0.73/0.6 = 1.2 mol kg⁻¹; ΔT_f = 1.86 × 1.2 = **2.2 K**; T_f = **270.95 K**.

NCERT Example 1.10 1.00 g of solute in 50 g benzene lowers its freezing point by 0.40 K (K_f = 5.12). M?

> **Ans.** M₂ = 5.12 × 1.00 × 1000/(0.40 × 50) = **256 g mol⁻¹**.

4. Osmosis and osmotic pressure



(a) Osmosis: solvent moves through the membrane into the solution. (b) Reverse osmosis: a pressure larger than π pushes pure water out (desalination; cellulose acetate membrane)

Definition ✎ **Osmosis**: flow of solvent through a semipermeable membrane (SPM) from pure solvent (or dilute solution) into the solution (or more concentrated one). **Osmotic pressure** π: the extra pressure that must be applied to the solution to just stop osmosis.

$$\pi = CRT = \frac{n_2RT}{V}, \text{ so } M_2 = \frac{w_2RT}{\pi V}$$

Best method for **macromolecules** (proteins, polymers): it works at room temperature and π is large enough to measure even for dilute solutions.

Isotonic solutions have the **same osmotic pressure** (blood: 0.9% NaCl). Hypertonic: higher π; hypotonic: lower. Red blood cells **shrink** in hypertonic and **swell** (may burst) in hypotonic solution.

Reverse osmosis: a pressure larger than π on the solution side drives solvent out through the membrane (cellulose acetate): desalination of sea water.

NCERT Example 1.11 200 cm³ of solution with 1.26 g protein; π = 2.57 × 10⁻³ bar at 300 K. M?

> **Ans.** M₂ = 1.26 × 0.083 × 300/(2.57 × 10⁻³ × 0.200) = **61 022 g mol⁻¹**.

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Intext questions, solved

1.9 p°(water) = 23.8 mm Hg at 298 K. 50 g urea in 850 g water: vapour pressure and relative lowering?

> **Ans.** n(urea) = 50/60 = 0.833; n(water) = 850/18 = 47.22. x_2 = 0.833/48.06 = **0.0173** (relative lowering). p = 23.8 × (1 − 0.0173) = **23.4 mm Hg**.

1.10 Water boils at 99.63 °C at 750 mm Hg. Sucrose to add to 500 g water so it boils at 100 °C?

> **Ans.** ΔT_b = 0.37 K: $w_2 = \Delta T_b M_2 w_1/(K_b \times 1000)$ = 0.37 × 342 × 500/(0.52 × 1000) = **121.67 g**.

1.11 Ascorbic acid ($C_6H_8O_6$, 176) to dissolve in 75 g acetic acid to lower its melting point by 1.5 °C (K_f = 3.9).

> **Ans.** w_2 = 1.5 × 176 × 75/(3.9 × 1000) = **5.08 g**.

1.12 Osmotic pressure (Pa) of 1.0 g of a polymer (M = 185 000) in 450 mL water at 37 °C.

> **Ans.** π = (1.0/185 000) × 8.314 × 10^3 × 310/0.450 = **30.96 Pa** (R = 8.314 kPa L K^{-1} mol^{-1}).

1.7 Abnormal molar masses ★★★

If the solute **dissociates** ($NaCl \rightarrow Na^+ + Cl^-$) there are more particles: colligative effects are larger and the molar mass found is **lower** than the true one. If it **associates** ($2CH_3COOH \rightleftharpoons$ dimer in benzene) there are fewer particles: the molar mass found is **higher**.

Definition ✎ van't Hoff factor:

$$i = \frac{\text{normal molar mass}}{\text{abnormal molar mass}}$$

$$i = \frac{\text{observed colligative property}}{\text{calculated colligative property}}$$

i > 1: dissociation; i < 1: association; i = 1: neither (glucose, urea).

$$\Delta T_f = iK_f m, \quad \Delta T_b = iK_b m$$

$$\pi = iCRT, \quad \frac{\Delta p}{p^\circ} = ix_2$$

	degree
dissociation into n ions	$\alpha = \frac{i-1}{n-1}$
association into $(A)_n$	$\alpha = \frac{1-i}{1-1/n}$; dimer: α = 2(1 − i)

Complete dissociation: KCl, NaCl i = 2; $CaCl_2$, K_2SO_4, $Na_2SO_4 \cdot 10H_2O$ i = 3; $AlCl_3$ i = 4. Complete dimerisation: i = 0.5.

NCERT Example 1.12 2 g benzoic acid in 25 g benzene: ΔT_f = 1.62 K (K_f = 4.9). % association (dimer)?

> **Ans.** M(obs) = 4.9 × 2 × 1000/(25 × 1.62) = 241.98. i = 122/241.98 = 0.504 = 1 − x/2: x = **0.992 (99.2%)**.

NCERT Example 1.13 0.6 mL acetic acid (d = 1.06) in 1 L water: ΔT_f = 0.0205 °C. i and K_a?

> **Ans.** n = 0.6 × 1.06/60 = 0.0106 mol; m = 0.0106. ΔT_f(calc) = 1.86 × 0.0106 = 0.0197 K. i = 0.0205/0.0197 = **1.041** = 1 + α: α = 0.041. $K_a = c\alpha^2/(1-\alpha)$ = 0.0106 × 0.041^2/0.959 = **1.86 × 10^{-5}**.

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NCERT Exercises, solved

1.1 Define solution. How many types of solutions are formed? Describe each with an example.

> **Ans.** A homogeneous mixture of two or more components. Nine types: gas, liquid or solid solute in a gaseous, liquid or solid solvent (table on p. 2).

1.2 A solid solution in which the solute is a gas.

> **Ans.** Hydrogen in palladium.

1.3 Define mole fraction, molality, molarity and mass percentage.

> **Ans.** Mole fraction: moles of a component divided by total moles. Molality: moles of solute per kg of solvent. Molarity: moles of solute per litre of solution. Mass %: mass of the component per 100 g of solution.

1.4 68% HNO_3 by mass, density 1.504 g mL^{-1}. Molarity?

> **Ans.** 100 g has 68 g HNO_3 = 1.079 mol; V = 100/1.504 = 66.5 mL. M = 1.079/0.0665 = **16.23 M**.

1.5 10% w/w glucose: molality and mole fractions; molarity if density 1.2 g mL^{-1}.

> **Ans.** In 100 g: 10 g glucose (0.0556 mol), 90 g water (5 mol). m = 0.0556/0.090 = **0.617 m**. x(glucose) = 0.0556/5.0556 = **0.011**, x(water) = **0.989**. V = 100/1.2 = 83.3 mL: M = **0.67 M**.

1.6 mL of 0.1 M HCl to react completely with 1 g of an equimolar Na_2CO_3 + $NaHCO_3$ mixture.

> **Ans.** Let x mol of each: 106x + 84x = 1, x = 5.26×10^{-3}. Na_2CO_3 needs 2 HCl, $NaHCO_3$ 1: HCl = 3x = 0.0158 mol. Volume = 0.0158/0.1 L = **157.9 mL**.

1.7 300 g of 25% and 400 g of 40% solutions are mixed. Mass % of the result?

> **Ans.** Solute = 75 + 160 = 235 g in 700 g: **33.57%** (solvent 66.43%).

1.8 222.6 g ethylene glycol + 200 g water. Molality; molarity if density 1.072 g mL^{-1}.

> **Ans.** n = 222.6/62 = 3.59 mol; m = 3.59/0.2 = **17.95 m**. V = 422.6/1.072 = 394.2 mL: M = **9.11 M**.

1.9 Water contaminated with 15 ppm (by mass) $CHCl_3$: (i) mass %; (ii) molality.

> **Ans.** (i) $15/10^6 \times 100 = \mathbf{1.5 \times 10^{-3}}$ **%**. (ii) 15 g $CHCl_3$ (0.1255 mol) per 10^6 g ≈ 10^3 kg water: m = $\mathbf{1.26 \times 10^{-4}}$ **m**.

1.10 Role of molecular interaction in alcohol + water.

> **Ans.** Both are H-bonded. On mixing, the new alcohol–water H-bonds are weaker than the original ones, so the solution shows **positive deviation** (higher vapour pressure, lower boiling point than expected).

1.11 Why are gases less soluble in liquids as temperature rises?

> **Ans.** Dissolving a gas is exothermic; by Le Chatelier's principle, raising T shifts the equilibrium back towards gas. (K_H also increases with T.)

1.12 State Henry's law and some applications.

> **Ans.** At constant T, the solubility of a gas is proportional to its partial pressure ($p = K_H x$). Applications: sealing soda bottles under pressure; bends in scuba divers (He-diluted air tanks); anoxia at high altitude (p. 4).

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1.13 p(ethane) = 1 bar over a solution with 6.56×10^{-3} g of ethane. Partial pressure if the solution contains 5.00×10^{-2} g?

> **Ans.** $p \propto$ amount dissolved: $p = 1 \times 5.00 \times 10^{-2} / 6.56 \times 10^{-3} =$ **7.62 bar**.

1.15 A 2% aqueous solution of a non-volatile solute has p = 1.004 bar at the solvent's normal boiling point. M?

> **Ans.** $p^\circ = 1.013$ bar. $(1.013 - 1.004)/1.013 = 2 \times 18/(M_2 \times 98)$: $0.00888 = 0.367/M_2$, $M_2 =$ **41.35 g mol^{-1}**.

1.16 Heptane and octane (ideal). $p^\circ = 105.2$ and 46.8 kPa at 373 K. p of 26.0 g heptane + 35 g octane?

> **Ans.** n = 0.26 and 0.307; x(heptane) = 0.459. $p = 0.459 \times 105.2 + 0.541 \times 46.8 =$ **73.6 kPa**.

1.17 p°(water) = 12.3 kPa at 300 K. p of a 1 molal solution?

> **Ans.** $x_2 = 1/(1 + 55.5) = 0.0177$; $p = 12.3 \times 0.9823 =$ **12.08 kPa**.

1.18 Mass of a non-volatile solute (M = 40) in 114 g octane to reduce its vapour pressure to 80%.

> **Ans.** $x_2 = 0.20$. n(octane) = 1 mol: $n_2/(1 + n_2) = 0.2 \Rightarrow n_2 = 0.25$ mol = **10 g**.

1.19 30 g solute in 90 g water: p = 2.8 kPa at 298 K. After adding 18 g water, p = 2.9 kPa. (i) M; (ii) p°(water).

> **Ans.** Let $n_2 = 30/M$; water is 5 mol, then 6 mol. Raoult: $2.8 = p^\circ \times 5/(5 + n_2)$ and $2.9 = p^\circ \times 6/(6 + n_2)$. Dividing: $2.8 \times 6(5 + n_2) = 2.9 \times 5(6 + n_2) \Rightarrow 84 + 16.8n_2 = 87 + 14.5n_2 \Rightarrow n_2 = 1.304$ mol. (i) $M = 30/1.304 =$ **23 g mol^{-1}**. (ii) $p^\circ = 2.8 \times 6.304/5 =$ **3.53 kPa**.

1.20 5% cane sugar freezes at 271 K. Freezing point of 5% glucose? ($T_f^\circ = 273.15$ K)

> **Ans.** $\Delta T_f \propto 1/M$ (same mass %): ΔT_f(glucose) $= 2.15 \times 342/180 = 4.09$ K: $T_f =$ **269.07 K**.

1.21 AB_2 and AB_4 in 20 g benzene: 1 g AB_2 gives $\Delta T_f = 2.3$ K, 1.0 g AB_4 gives 1.3 K ($K_f = 5.1$). Atomic masses of A and B?

> **Ans.** $M(AB_2) = 5.1 \times 1 \times 1000/(2.3 \times 20) = 110.9$; $M(AB_4) = 5.1 \times 1000/(1.3 \times 20) = 196.2$. $a + 2b = 110.9$, $a + 4b = 196.2 \Rightarrow b =$ **42.6 u**, $a =$ **25.6 u**.

1.22 36 g glucose per litre has $\pi = 4.98$ bar at 300 K. Concentration if $\pi = 1.52$ bar?

> **Ans.** $\pi \propto C$: $C_1 = 36/180 = 0.2$ M; $C_2 = 0.2 \times 1.52/4.98 =$ **0.061 M**.

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1.23 Main intermolecular interaction in: (i) n-hexane + n-octane; (ii) I_2 + CCl_4; (iii) $NaClO_4$ + water; (iv) methanol + acetone; (v) acetonitrile + acetone.

> **Ans.** (i) London dispersion forces. (ii) London dispersion. (iii) Ion–dipole. (iv) Dipole–dipole (with H-bonding between the O–H of methanol and the C=O of acetone). (v) Dipole–dipole.

1.24 Increasing solubility in n-octane: cyclohexane, KCl, CH_3OH, CH_3CN.

> **Ans.** Octane is non-polar, so the least polar dissolves best: **KCl < CH_3OH < CH_3CN < cyclohexane**.

1.25 Insoluble, partially soluble or highly soluble in water: phenol, toluene, formic acid, ethylene glycol, chloroform, pentanol.

> **Ans.** Insoluble: **toluene, chloroform**. Partially soluble: **phenol, pentanol** (large non-polar part). Highly soluble: **formic acid, ethylene glycol** (H-bonding).

1.26 Lake water: density 1.25 g mL^{-1}, 92 g Na^+ per kg of water. Molarity of Na^+?

> **Ans.** 92/23 = 4 mol. Taking the solution as about 1 kg: V = 1000/1.25 = 800 mL: **5 M** (including the 92 g in the mass, V = 873.6 mL and 4.6 M).

1.27 $K_{sp}(CuS) = 6 \times 10^{-16}$. Maximum molarity of CuS in water?

> **Ans.** $K_{sp} = S^2$: $S = \sqrt{(6 \times 10^{-16})} = \mathbf{2.45 \times 10^{-8}\ M}$.

1.28 Mass % of aspirin when 6.5 g is dissolved in 450 g CH_3CN.

> **Ans.** 6.5/456.5 × 100 = **1.42%**.

1.29 Mass of 1.5×10^{-3} m aqueous nalorphene ($C_{19}H_{21}NO_3$, 311) solution for a 1.5 mg dose.

> **Ans.** $n = 1.5 \times 10^{-3}/311 = 4.82 \times 10^{-6}$ mol; water needed = $n/m = 3.21 \times 10^{-3}$ kg = 3.21 g. Solution ≈ **3.22 g**.

1.30 Benzoic acid for 250 mL of 0.15 M solution in methanol.

> **Ans.** 0.15 × 0.250 = 0.0375 mol × 122 = **4.575 g**.

1.31 ΔT_f for the same amount of acetic, trichloroacetic and trifluoroacetic acids increases in that order. Explain.

> **Ans.** F is more electronegative than Cl (–I effect), so acid strength is $CF_3COOH > CCl_3COOH > CH_3COOH$. A stronger acid dissociates more, giving more particles (larger i) and a larger ΔT_f.

1.32 ΔT_f when 10 g $CH_3CH_2CHClCOOH$ (M = 122.5) is added to 250 g water; $K_a = 1.4 \times 10^{-3}$, $K_f = 1.86$.

> **Ans.** $m = (10/122.5)/0.25 = 0.327$. $\alpha = \sqrt{(K_a/c)} = \sqrt{(1.4 \times 10^{-3}/0.327)} = 0.0655$; $i = 1 + \alpha = 1.065$. $\Delta T_f = 1.065 \times 1.86 \times 0.327 = \mathbf{0.65\ K}$.

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1.33 19.5 g CH_2FCOOH (78) in 500 g water: ΔT_f = 1.00 °C. van't Hoff factor and K_a.

> **Ans.** m = 0.25/0.5 = 0.5; $\Delta T_f(calc)$ = 1.86 × 0.5 = 0.93 K. i = 1.00/0.93 = **1.075**; α = 0.075. $K_a = c\alpha^2/(1-\alpha)$ = 0.5 × 0.075^2/0.925 = **3.07×10^{-3}**.

1.34 p°(water) = 17.535 mm Hg at 293 K. p when 25 g glucose is dissolved in 450 g water.

> **Ans.** x_2 = 0.139/(0.139 + 25) = 0.00552: p = 17.535 × 0.99448 = **17.44 mm Hg**.

1.35 K_H for methane in benzene = 4.27×10^5 mm Hg at 298 K. Solubility under 760 mm Hg?

> **Ans.** $x = 760/4.27 \times 10^5 = \mathbf{1.78 \times 10^{-3}}$.

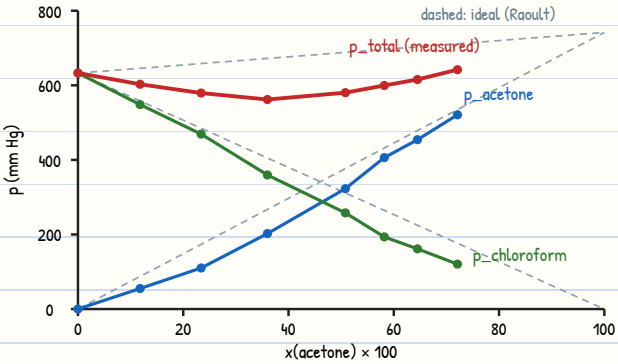
1.36 100 g A (140) in 1000 g B (180). p°(B) = 500 torr; p_{total} = 475 torr. p°(A) and p(A) in solution?

> **Ans.** n_A = 0.714, n_B = 5.556; x_A = 0.114. p_B = 500 × 0.886 = 443 torr; p_A = 475 − 443 = **32 torr**; p°_A = 32/0.114 = **280.7 torr**.

1.38 Benzene (50.71) and toluene (32.06 mm Hg) at 300 K. Mole fraction of benzene in vapour for 80 g benzene + 100 g toluene.

> **Ans.** n = 1.026 and 1.087; x_b = 0.486. p_b = 24.6, p_t = 16.5 mm Hg: y_b = 24.6/41.1 = **0.60**.

1.37 Plot p_{total}, $p_{chloroform}$, $p_{acetone}$ against x(acetone) for ideal behaviour and for the given data. Positive or negative deviation?



NCERT Ex 1.37 plotted from the book's data: every measured curve lies below the ideal line, so acetone + chloroform shows negative deviation

> **Ans.** The measured curves lie below the ideal lines: **negative deviation** (acetone–chloroform H-bonding).

1.39 Air (20% O_2, 79% N_2) at 10 atm over water at 298 K; K_H = 3.30×10^7 mm (O_2), 6.51×10^7 mm (N_2). Composition in water?

> **Ans.** P = 7600 mm Hg: $p(O_2)$ = 1520, $p(N_2)$ = 6004 mm. $x(O_2) = 1520/3.30 \times 10^7 = \mathbf{4.6 \times 10^{-5}}$; $x(N_2) = \mathbf{9.22 \times 10^{-5}}$.

1.40 $CaCl_2$ (i = 2.47) in 2.5 L water for π = 0.75 atm at 27 °C.

> **Ans.** $n = \pi V/(iRT)$ = 0.75 × 2.5/(2.47 × 0.0821 × 300) = 0.0308 mol × 111 = **3.42 g**.

1.41 π of 25 mg K_2SO_4 in 2 L water at 25 °C (complete dissociation).

> **Ans.** i = 3; n = 0.025/174 = 1.44×10^{-4} mol. $\pi = 3 \times 1.44 \times 10^{-4} \times 0.0821 \times 298/2 = \mathbf{5.27 \times 10^{-3}}$ atm.

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CBSE board questions 2024–2026, fully solved

Every Solutions question from the 16 papers (sets 56/1/1 to 56/7/1, 2024–2026), solved step by step. Mark steps follow the CBSE marking schemes.

A. Concentration and Henry's law

CBSE 2026 · 56/3/1 The mole fraction of a solute in a 2.0 molal aqueous solution is: (A) 1.87 (B) 0.347 (C) 0.0347 (D) 0.00347 [1]

Ans. 2 mol solute in 1 kg water (1000/18 = 55.56 mol). x = 2/(2 + 55.56) = **(C) 0.0347**.

CBSE 2025 · 56/4/1 The value of Henry's constant K_H is: (A) greater for gases with higher solubility (B) greater for gases with lower solubility (C) constant for all gases (D) not related to solubility [1]

Ans. $x = p/K_H$: at the same p, a larger K_H gives a smaller x (less solubility): **(B)**.

CBSE 2025 · 56/5/1 A–R. (A) K_H decreases with increase in temperature. (R) As temperature increases, solubility of gases in liquids decreases. [1]

Ans. Solubility falls with T, so K_H (= p/x) **increases**: A is false, R is true: **(D)**.

CBSE 2025 · 56/6/1 State Henry's law. Why are aquatic species more comfortable in cold water than in warm water? [2]

Ans. At constant temperature, the solubility of a gas in a liquid is directly proportional to its partial pressure above the liquid ($p = K_H x$) 1. K_H rises with temperature, so less O_2 dissolves in warm water; cold water holds more O_2 1.

CBSE 2024 · 56/4/1 (a)(ii) Why are aquatic species more comfortable in cold water? [1]

Ans. Solubility of gases decreases with rise in temperature, so more oxygen is available in cold water.

CBSE 2024 · 56/5/1 (b)(ii) State Henry's law and write any one application. [2]

Ans. Law as above 1. Application: soft-drink and soda bottles are sealed under high CO_2 pressure to increase its solubility 1 (or: bends, anoxia).

CBSE 2026 · 56/4/1 (a) Explain with Henry's law: (i) bends, (ii) anoxia. OR (b) How does sprinkling salt help clear snow-covered roads? Name the colligative property. [2]

Ans. (a)(i) Deep under water the pressure is high, so more N_2 dissolves in a diver's blood. When the diver rises, the pressure falls, N_2 comes out as bubbles and blocks capillaries: a painful condition called bends 1.
(ii) At high altitude the partial pressure of O_2 is low, so blood and tissues hold less O_2: people feel weak and cannot think clearly (anoxia) 1.
OR (b) Salt dissolves in the thin layer of water on the snow and **lowers its freezing point**, so the snow melts at the road temperature and can be cleared 1. Property: **depression of freezing point** 1.

CBSE 2026 · 56/5/1 (b) State Henry's law. Mole fraction of CO_2 in water at 298 K under 760 mm Hg ($K_H = 1.25 \times 10^6$ mm Hg). [2]

Ans. Law 1. $p = K_H x$: $x = 760/1.25 \times 10^6 = \mathbf{6.08 \times 10^{-4}}$ 1.

CBSE 2026 · 56/4/1 A liquid and its vapour are at equilibrium and the pressure is suddenly decreased: (A) vapour pressure of the solution increases (B) heating occurs (C) cooling occurs (D) equilibrium unaffected [1]

Ans. Lower pressure makes more liquid evaporate. Evaporation absorbs latent heat from the liquid: **(C) cooling occurs**.

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B. Raoult's law, ideal and non-ideal solutions, azeotropes

CBSE 2025 · 56/4/1, Q29 (case, 4 marks): ideal solutions as limiting cases; Raoult's law deduced thermodynamically.

(a) One miscible pair showing negative deviation, and the reason. [2]

Ans. Chloroform + acetone (1). A–B interactions (H-bond between C–H of $CHCl_3$ and C=O of acetone) are **stronger** than A–A and B–B interactions, so fewer molecules escape (1).

(b)(i) State Raoult's law for volatile components. [1]

Ans. The partial vapour pressure of each volatile component is directly proportional to its mole fraction in the solution: $p_1 = p_1^\circ x_1$.

OR (b)(ii) Raoult's law is a special case of Henry's law. Comment.

Ans. Henry: $p = K_H x$; Raoult: $p = p^\circ x$. Both say $p \propto x$; Raoult's law is Henry's law with $K_H = p^\circ$.

(c) Two characteristics of an ideal solution. [1]

Ans. $\Delta_{mix}H = 0$ and $\Delta_{mix}V = 0$ (also: obeys Raoult's law at all concentrations).

CBSE 2024 · 56/2/1 The relative lowering of vapour pressure of an aqueous solution of a non-volatile solute is 0.0225. Mole fraction of the solute: (A) 0.80 (B) 0.725 (C) 0.15 (D) 0.0225 [1]

> **Ans.** Relative lowering = $(p^\circ - p)/p^\circ = x_2$: **(D) 0.0225.**

CBSE 2025 · 56/5/1 (b) X and Y form an ideal solution; $p^\circ(X) = 120$, $p^\circ(Y) = 160$ mm Hg. p of a solution with equal moles of X and Y. [2]

> **Ans.** $x_X = x_Y = 0.5$ (½). $p = p^\circ_X x_X + p^\circ_Y x_Y = 120 \times 0.5 + 160 \times 0.5$ (1) $= 60 + 80 =$ **140 mm Hg** (½).

CBSE 2025 · 56/6/1 A solution of acetone in ethanol: (A) obeys Raoult's law (B) is ideal (C) shows positive deviation (D) shows negative deviation [1]

> **Ans.** Acetone breaks the H-bonds between ethanol molecules and the new attractions are weaker: **(C).**

CBSE 2024 · 56/3/1 OR (b) What deviation is shown by an ethanol + acetone mixture? Reason. Which azeotrope does it form? [2]

> **Ans.** **Positive deviation** (½), because ethanol–acetone interactions are weaker than ethanol–ethanol (H-bonded) and acetone–acetone interactions (½). It forms a **minimum-boiling azeotrope** (1).

CBSE 2024 · 56/5/1 (a)(ii) Deviation shown by ethanol + acetone? Reason. [2]

> **Ans.** Positive deviation (1): ethanol–ethanol attraction (H-bonding) is stronger than ethanol–acetone attraction (1).

CBSE 2025 · 56/7/1 What is positive deviation from Raoult's law? Example. Which azeotrope does it form? [2]

> **Ans.** The vapour pressure of the solution is higher than expected from Raoult's law (1). Example: ethanol + acetone (or CS_2 + acetone) (½). Minimum-boiling azeotrope (½).

CBSE 2026 · 56/1/1 Deviation shown by phenol + aniline? Reason. What happens to the boiling point on mixing? [2]

> **Ans.** **Negative deviation** (½). The H-bond between the phenolic –OH and the lone pair on aniline's N is stronger than the H-bonds between like molecules (1). Vapour pressure falls, so the **boiling point increases** (½).

CBSE 2026 · 56/3/1 An azeotrope of water and ethanol boils below both pure liquids. It shows: (A) positive deviation (B) negative deviation (C) no deviation (D) that it is ideal [1]

> **Ans.** A minimum-boiling azeotrope comes from higher vapour pressure: **(A).**

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CBSE 2025 · S6/1/1 (A)(a) Why is cooking faster in a pressure cooker? (b) Mixing liquids X and Y decreases the volume. Type of deviation? Temperature change on mixing? OR (B) Define azeotrope. Which azeotrope forms by negative deviation? Example. [2]

Ans. (a) The pressure inside is high, so water boils at a higher temperature and food cooks faster 1. (b) $\Delta_{mix}V < 0$: **negative deviation**; $\Delta_{mix}H < 0$, so the **temperature rises** ½ + ½.
OR (B) A binary mixture with the same composition in liquid and vapour that boils at a constant temperature 1. **Maximum-boiling azeotrope** ½, e.g. 68% HNO_3 + 32% water ½.

CBSE 2025 · S6/7/1 A-R. (A) Cooking time is reduced in a pressure cooker. (R) The boiling point of water inside is elevated. [1]

Ans. Both true; the higher boiling point is the reason: (A).

CBSE 2026 · S6/5/1 (a)(ii) Two differences between ideal and non-ideal solutions. [2]

Ans. Ideal: obeys Raoult's law at all concentrations; $\Delta_{mix}H = 0$, $\Delta_{mix}V = 0$. Non-ideal: does not obey Raoult's law; $\Delta_{mix}H \neq 0$, $\Delta_{mix}V \neq 0$ 1 + 1.

CBSE 2026 · S6/5/1 (b)(ii) Which azeotrope does acetone + chloroform form? Reason. [2]

Ans. **Maximum-boiling azeotrope** 1, because the mixture shows negative deviation: acetone–chloroform interactions (H-bonding) are stronger 1.

CBSE 2024 · S6/4/1 (b)(ii) Define ideal solution. [1]

Ans. A solution that obeys Raoult's law over the entire range of concentration ($\Delta_{mix}H = 0$, $\Delta_{mix}V = 0$).

C. Relative lowering of vapour pressure

CBSE 2024 · S6/3/1 5 g of a non-volatile solute in 200 g water: p = 31.84 mm Hg at 300 K (p° = 32 mm Hg). Molar mass of the solute? [3]

Ans. $(p^\circ - p)/p^\circ = w_2M_1/(M_2w_1)$ 1
$(32 - 31.84)/32 = 5 \times 18/(M_2 \times 200)$ 1
$0.005 = 0.45/M_2 \Rightarrow$ **M_2 = 90 g mol^{-1}** 1

CBSE 2024 · S6/4/1 (a)(iii) 2 g glucose (180) in 100 g water at 303 K; p°(water) = 32.8 mm Hg. Vapour pressure of the solution? [3]

Ans. $(32.8 - p)/32.8 = 2 \times 18/(180 \times 100) = 0.002$ 1 + 1
$32.8 - p = 0.0656 \Rightarrow$ **p = 32.73 mm Hg** 1

CBSE 2025 · S6/2/1 At 25 °C p°(water) = 24 mm Hg. Vapour pressure of a 5% aqueous urea solution (urea = 60)? [3]

Ans. In 100 g: 5 g urea (0.0833 mol), 95 g water (5.278 mol) 1.
$x_2 = 0.0833/(0.0833 + 5.278) = 0.01554$ 1.
$p = p^\circ(1 - x_2) = 24 \times 0.98446 =$ **23.63 mm Hg** 1.

CBSE 2025 · S6/6/1 p°(water) = 24.8 mm Hg at 298 K. Lowering of vapour pressure of an aqueous solution that freezes at −0.3 °C (K_f = 1.86). [3]

Ans. $m = \Delta T_f/K_f = 0.3/1.86 = 0.161$ mol kg^{-1} 1.
$x_2 \approx m \times M_1/1000 = 0.161 \times 18/1000 = 2.90 \times 10^{-3}$ 1.
$\Delta p = x_2\, p^\circ = 2.90 \times 10^{-3} \times 24.8 =$ **0.072 mm Hg** 1.

CBSE 2026 · S6/4/1 Vapour pressure of 61 g benzoic acid (122) in 500 g benzene (78) if p° = 66 torr; benzoic acid dimerises completely. [3]

Ans. Complete dimerisation: i = 0.5 1.
$(p^\circ - p)/p^\circ = i \times w_2M_1/(M_2w_1) = 0.5 \times (61 \times 78)/(122 \times 500) = 0.039$ 1.
$66 - p = 66 \times 0.039 = 2.57 \Rightarrow$ **p = 63.4 torr** 1.

★ ***Exam tip:*** For RLVP with i, the dilute form $(p^\circ - p)/p^\circ = i\, x_2 \approx i\, w_2M_1/(M_2w_1)$ is what the schemes use.

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D. Elevation of boiling point and depression of freezing point

CBSE 2024 · S6/3/I (a) Define molal depression constant. How is it related to the enthalpy of fusion? [2]

Ans. The depression in freezing point of a 1 molal solution (1 mol of non-volatile solute in 1 kg of solvent) (1). $K_f = \frac{R \times M_1 \times T_f^2}{1000 \times \Delta_{fus}H}$ (1).

CBSE 2024 · S6/4/I 18 g of a non-volatile solute in 200 g water freezes at 272.07 K. Molecular mass (K_f = 1.86)? [2]

Ans. $\Delta T_f = 273.15 - 272.07 = 1.08$ K (½). $M_2 = K_f w_2 \times 1000/(w_1 \Delta T_f) = 1.86 \times 18 \times 1000/(200 \times 1.08)$ (1) = **155 g mol⁻¹** (½).

CBSE 2026 · S6/2/I 1.5 g of a non-volatile solute in 90 g benzene raises its boiling point from 353.23 K to 353.93 K (K_b = 2.52). Molar mass? [3]

Ans. $\Delta T_b = 353.93 - 353.23 = 0.70$ K (½). $\Delta T_b = K_b w_2 \times 1000/(M_2 w_1)$ (½): $0.70 = 2.52 \times 1.5 \times 1000/(M_2 \times 90)$ (½). $M_2 = 3780/63 =$ **60 g mol⁻¹** (1½).

CBSE 2026 · S6/3/I Antifreeze: 31 g ethylene glycol (62) in 600 g water. Freezing point (K_f = 1.86)? [3]

Ans. $\Delta T_f = K_f w_2 \times 1000/(M_2 w_1)$ (½) $= 1.86 \times 31 \times 1000/(62 \times 600)$ (1) = 1.55 K (½). $T_f = 0 - 1.55 =$ **−1.55 °C (271.6 K)** (1).

CBSE 2025 · S6/1/I Glucose solution boils at 100.20 °C. Its freezing point? (K_f = 1.86, K_b = 0.512) [3]

Ans. $\Delta T_b = 0.20$ K (½); $m = \Delta T_b/K_b = 0.20/0.512 = 0.391$ mol kg⁻¹ (1). $\Delta T_f = K_f m = 1.86 \times 0.391 = 0.727$ K (1); $T_f =$ **−0.73 °C** (½).

CBSE 2024 · S6/5/I (a)(i) Ishan's radiator holds 1.0 kg water. Grams of ethylene glycol (62) to lower the freezing point to −2.8 °C (K_f = 1.86)? [3]

Ans. $\Delta T_f = 2.8$ K. $\Delta T_f = K_f w_2/(M_2 \times W_1$ in kg) (1): $2.8 = 1.86 \times w_2/(62 \times 1)$ (1). $w_2 = 2.8 \times 62/1.86 =$ **93.33 g** (1).

CBSE 2024 · S6/5/I OR (b)(i) Water boils at 99.68 °C at 750 mm Hg. Sucrose (342) to add to 500 g water so it boils at 100 °C (K_b = 0.52)? [3]

Ans. $\Delta T_b = 100 - 99.68 = 0.32$ K (1). $0.32 = 0.52 \times w_2 \times 1000/(342 \times 500)$ (1) $\Rightarrow w_2 = 0.32 \times 342 \times 500/(0.52 \times 1000) =$ **105.2 g** (1).

CBSE 2026 · S6/4/I Boiling point of 1 molal NaCl (complete dissociation, K_b = 0.52): (A) 100.52 °C (B) 101.04 °C (C) 100.04 °C (D) 101.52 °C [1]

Ans. i = 2: $\Delta T_b = 2 \times 0.52 \times 1 = 1.04$ K: **(B) 101.04 °C.**

CBSE 2025 · S6/6/I Freezing point of 1 molal KCl (complete dissociation, K_f = 1.86): (A) −3.72 °C (B) +3.72 °C (C) −1.86 °C (D) +2.72 °C [1]

Ans. i = 2: $\Delta T_f = 2 \times 1.86 \times 1 = 3.72$ K: **(A) −3.72 °C.**

CBSE 2025 · S6/5/I Highest freezing point: (A) 1.0 M KCl (B) 1.0 M Na_2SO_4 (C) 1.0 M glucose (D) 1.0 M $AlCl_3$ [1]

Ans. Fewest particles gives the smallest depression. i = 2, 3, 1, 4: **(C) glucose.**

CBSE 2026 · S6/1/I Lowest freezing point: (A) 0.1 M glucose (B) 0.1 M $CaCl_2$ (C) 0.1 M KCl (D) 0.1 M urea [1]

Ans. Most particles: $CaCl_2$ (i = 3): **(B).**

CBSE 2025 · S6/7/I Highest boiling point: (A) 1% KCl (B) 1% glucose (C) 1% urea (D) 1% $CaCl_2$ [1]

Ans. Moles of particles per gram: KCl 2/74.5 = 0.0268; glucose 1/180 = 0.0056; urea 1/60 = 0.0167; $CaCl_2$ 3/111 = 0.0270. Largest: **(D) $CaCl_2$.**

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CBSE 2026 · S6/3/I 1 mol of KCl in 1 L water raises the boiling point, but 1 mol of methyl alcohol lowers it. Explain. [2]

Ans. KCl is a non-volatile electrolyte: it dissociates into 2 ions and lowers the vapour pressure, so the boiling point rises (1). Methyl alcohol is **volatile** (b.p. 65 °C): it adds its own vapour and disturbs water–water H-bonding, so the vapour pressure rises and the boiling point falls (1).

CBSE 2024 · S6/1/I A–R. (A) Adding NaCl to water depresses its freezing point. (R) NaCl dissociates in water. [1]

Ans. Both true. The depression happens with any non-volatile solute; dissociation only makes it larger. R is not the explanation: **(B)**.

CBSE 2024 · S6/3/I A–R. (A) Ethylene glycol lowers the freezing point of water. (R) Ethylene glycol is insoluble in water as it cannot H-bond with water. [1]

Ans. A is true (antifreeze). R is false: glycol is highly soluble because it H-bonds with water: **(C)**.

CBSE 2025 · S6/5/I (a) Elevation of boiling point when 3 g $CaCl_2$ (111) is dissolved in 260 g water, with complete dissociation (K_b = 0.52). [2]

Ans. $\Delta T_b = iK_b m$ (½); i = 3 (½).
$\Delta T_b = 3 \times 0.52 \times 3 \times 1000/(111 \times 260)$ (½) = **0.162 K** (½).

CBSE 2026 · S6/5/I (a)(i) Freezing point of 10.5 g $MgBr_2$ (184) in 250 g water, complete dissociation (K_f = 1.86). [3]

Ans. $MgBr_2 \rightarrow Mg^{2+} + 2Br^-$: i = 3 (½); $\Delta T_f = iK_f m$ (½).
$\Delta T_f = 3 \times 1.86 \times 10.5 \times 1000/(184 \times 250)$ = **1.27 K** (1).
$T_f = 273.15 - 1.27$ = **271.88 K (−1.27 °C)** (1).

CBSE 2024 · S6/4/I OR (b)(i) Is i less or greater than 1 for ethanoic acid in benzene? (ii) Define ideal solution. (iii) Mass of $CaCl_2$ (111) in 500 g water to lower the freezing point by 2 K (complete dissociation, K_f = 1.86). [5]

Ans. (i) **Less than 1**: ethanoic acid dimerises in benzene (H-bonding), so there are fewer particles (1).
(ii) A solution obeying Raoult's law at all concentrations (1).
(iii) i = 3. $\Delta T_f = iK_f w_2 \times 1000/(M_2 w_1)$: $2 = 3 \times 1.86 \times w_2 \times 1000/(111 \times 500)$ (1 + 1).
$w_2 = 2 \times 111 \times 500/(3 \times 1.86 \times 1000)$ = **19.89 g** (1).

E. Osmosis

CBSE 2024 · S6/2/I 6% glucose (180) is isotonic with a 2.5% solution of an unknown organic substance. Its molecular weight? [2]

Ans. Isotonic ⇒ same molar concentration: C(glucose) = C(unknown) (½).
6/180 = 2.5/M (1) ⇒ M = 2.5 × 180/6 = **75 g mol^{-1}** (½).

CBSE 2024 · S6/5/I Isotonic solutions have the same: (A) density (B) refractive index (C) osmotic pressure (D) volume [1]

Ans. **(C) osmotic pressure.**

CBSE 2025 · S6/1/I An unripe mango in concentrated salt solution shrivels because: (A) it gains water by osmosis (B) loses water by reverse osmosis (C) gains water by reverse osmosis (D) loses water by osmosis [1]

Ans. The brine is hypertonic, so water flows out of the mango's cells: **(D)**.

CBSE 2024 · S6/4/I (a)(i) Define reverse osmosis. [1]

Ans. If a pressure larger than the osmotic pressure is applied to the solution side, solvent flows from the solution to the pure solvent through the SPM.

CBSE 2026 · S6/5/I OR (b)(i) Osmotic pressure of 0.025 g K_2SO_4 (174) in 2 L water at 27 °C, complete dissociation (R = 0.082). [3]

Ans. $K_2SO_4 \rightarrow 2K^+ + SO_4^{2-}$: i = 3 (½ + ½).
$\pi = i(n/V)RT = 3 \times (0.025/174)/2 \times 0.082 \times 300$ (1) = **5.3×10^{-3} atm** (1).

CBSE 2025 · 56/2/1, Q29 (case, 4 marks): two eggs with shells removed; one in distilled water swells, one in saturated salt solution shrinks.
(a) Define reverse osmosis; name one SPM used. **Ans.** Solvent is forced from solution to solvent side by a pressure larger than π (1); cellulose acetate (1).
(b)(i) RBCs in 0.5% NaCl? **Ans.** 0.5% is hypotonic to blood (isotonic = 0.9%), so water enters and the cells **swell** and may burst (1). OR (ii) Higher π: 1 M KCl or 1 M urea? **Ans. 1 M KCl**: it gives 2 ions (i = 2), urea does not dissociate (i = 1).
(c) Why is osmotic pressure colligative? **Ans.** It depends only on the number of solute particles (π = iCRT), not on their nature (1).

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F. van't Hoff factor, dissociation and association

CBSE 2025 · S6/1/1 In case of association, the abnormal molar mass of the solute will: (A) increase (B) decrease (C) remain same (D) first increase, then decrease [1]

> **Ans.** Fewer particles ⇒ smaller colligative effect ⇒ larger calculated M: **(A) increase.**

CBSE 2024 · S6/S/1 van't Hoff factor for $Na_2SO_4·10H_2O$ (complete ionisation): (A) 1 (B) 3 (C) 13 (D) 2 [1]

> **Ans.** Water of crystallisation just joins the solvent; $Na_2SO_4 \longrightarrow 2Na^+ + SO_4^{2-}$: **(B) 3.**

CBSE 2025 · S6/4/1 The molar mass of NaOH from osmotic pressure at 27 °C is 25 g mol^{-1}. % dissociation of NaOH? [3]

> **Ans.** i = normal M/abnormal M ½ = 40/25 = 1.6 1.
> $\alpha = (i - 1)/(n - 1)$ with n = 2 ½: α = 0.6/1 = 0.6 = **60%** 1.

CBSE 2025 · S6/7/1 0.3 g acetic acid (60) in 30 g benzene: ΔT_f = 0.45 °C (K_f = 5.12). % association (dimer)? [3]

> **Ans.** $\Delta T_f = iK_f m$ ½: 0.45 = i × 5.12 × 0.3 × 1000/(60 × 30) 1.
> 0.45 = 0.8533 i ⇒ i = 0.527 ½.
> $\alpha = (1 - i)/(1 - 1/2)$ = 0.473/0.5 ½ = 0.946 = **94.6%** ½.

CBSE 2026 · S6/1/1 Boiling point of 0.61 g benzoic acid (122) in 5 g CS_2, dimerised to 88%. $T_b(CS_2)$ = 46.2 °C, K_b = 2.3. [3]

> **Ans.** For a dimer: $\alpha = (1 - i)/(1 - ½) \Rightarrow 0.88 = 2(1 - i) \Rightarrow i$ = **0.56** ½ + ½.
> $\Delta T_b = iK_b w_2 \times 1000/(M_2 w_1)$ = 0.56 × 2.3 × 0.61 × 1000/(122 × 5) = **1.288 K** 1.
> T_b = 46.2 + 1.288 = **47.49 °C** 1.

CBSE 2026 · S6/S/1 (a) A 1.00 molal aqueous solution of trichloroacetic acid boils at 100.18 °C. van't Hoff factor? (K_b = 0.512) [2]

> **Ans.** Calculated $\Delta T_b = K_b m$ = 0.512 × 1 = 0.512 K ½ + ½. Observed ΔT_b = 0.18 K.
> i = observed/calculated = 0.18/0.512 = **0.352** 1.
> (The data give i < 1; the scheme accepts 0.3516.)

★ **Exam tip:** i-questions in 3 lines: (1) write the dissociation or association; (2) get i (from n, from α, or from observed/calculated); (3) put i into the colligative formula, or use $\alpha = (i - 1)/(n - 1)$ or $\alpha = (1 - i)/(1 - 1/n)$.

What the board asked (16 papers)

Topic	Questions	Marks
ΔT_b / ΔT_f numericals (with or without i)	14	1–3
Raoult's law, deviations, azeotropes	13	1–4
Henry's law and applications	8	1–2
RLVP numericals	6	1–3
Osmosis, isotonic, reverse osmosis	6	1–4
van't Hoff factor, α	6	1–3

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Formula sheet

Quantity	Formula	Units / note
Henry's law	$p = K_H x$	larger K_H, lower solubility
Raoult's law	$p_i = p_i^\circ x_i$; $p = p_1^\circ + (p_2^\circ - p_1^\circ)x_2$	$y_i = p_i/p$
RLVP	$(p^\circ - p)/p^\circ = i\,x_2 \approx i\,w_2M_1/(M_2w_1)$	no units
Elevation of b.p.	$\Delta T_b = iK_b m$	water $K_b = 0.52$ K kg mol^{-1}
Depression of f.p.	$\Delta T_f = iK_f m$	water $K_f = 1.86$ K kg mol^{-1}
Osmotic pressure	$\pi = iCRT$	$R = 0.0821$ L atm or 0.083 L bar K^{-1} mol^{-1}
Molar mass (any)	$M_2 = K\,w_2 \times 1000/(\Delta T\,w_1)$; $M_2 = w_2RT/(\pi V)$	w in g
K_f, K_b	$K = RM_1T^2/(1000\Delta H)$	ΔH_{fus} or ΔH_{vap}
van't Hoff factor	i = M(normal)/M(observed)	dissociation $\alpha = (i-1)/(n-1)$; association $\alpha = (1-i)/(1-1/n)$

Don't-confuse table

Pair	Difference
Molarity vs molality	per litre of solution (changes with T) vs per kg of solvent (does not)
Positive vs negative deviation	A–B weaker, p higher, $\Delta H > 0$, min-boiling vs A–B stronger, p lower, $\Delta H < 0$, max-boiling
Osmosis vs reverse osmosis	solvent into the solution, natural vs solvent out of the solution under $P > \pi$
Isotonic, hypertonic, hypotonic	same π vs higher π (cells shrink) vs lower π (cells swell)

✗ **Don't do this:** Forgetting to convert the freezing point: ΔT_f is a **difference**; $T_f(\text{solution}) = T_f^\circ - \Delta T_f$. The same ΔT in °C and K.

✗ **Don't do this:** Using the mass of **solution** instead of solvent in molality, or forgetting i for electrolytes (NaCl 2, $CaCl_2$ 3, K_2SO_4 3, $AlCl_3$ 4).

Case study (practice): the antifreeze test. A car radiator in Leh must not freeze at −10 °C. A mechanic has ethylene glycol (62 g mol^{-1}) and common salt.
(i) Mass of glycol per kg of water needed ($K_f = 1.86$)? (ii) Why is salt not used in radiators though it lowers the freezing point more per gram? (iii) Does glycol also raise the boiling point? By how much ($K_b = 0.52$)? (iv) Name the colligative property used.

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Practice set: test yourself *Exam-style*

Try every question before looking at the answers on page 24. 64 or more out of 80 means you are ready for the board.

A. Fill in the blanks (1 mark each)

1. Molality is moles of solute per _____ of solvent.
2. $p = K_H x$ is _____'s law.
3. For an ideal solution $\Delta_{mix}H$ = _____.
4. Chloroform + acetone shows _____ deviation.
5. K_f of water is _____ K kg mol^{-1}.
6. Solutions with the same π are _____.
7. i for K_2SO_4 (complete dissociation) is _____.
8. The SPM used in reverse osmosis is _____.

B. True or false? (1 mark each)

1. Molarity is independent of temperature.
2. Gases are more soluble at higher temperature.
3. Ethanol + acetone forms a maximum-boiling azeotrope.
4. Colligative properties depend on the nature of the solute.
5. Association gives an abnormally high molar mass.
6. RBCs swell in 0.5% NaCl solution.
7. Osmotic pressure suits proteins best.
8. An azeotrope can be separated by fractional distillation.

C. Very short answer (1 mark each)

1. Define mole fraction.
2. Give one example of a solid solution with a gas solute.
3. What is a colligative property?
4. Define ebullioscopic constant.
5. What are isotonic solutions?
6. Name the cause of bends.
7. Why is K_H larger for a less soluble gas?
8. Give an example of positive deviation.
9. Write i for a solute that dimerises completely.
10. Why is molality preferred to molarity?

D. Short answer I (2 marks each)

1. State Raoult's law and show it is a special case of Henry's law.
2. Calculate the molality of 9 g glucose in 250 g water.
3. Distinguish between positive and negative deviations (two points).
4. Why does a solution of a non-volatile solute boil at a higher temperature?
5. Find the osmotic pressure of 0.1 M glucose at 27 °C.
6. Define azeotrope; give one example of each type.
7. Why does the molar mass of acetic acid in benzene come out about 120?
8. Find ΔT_f for 0.2 m NaCl (complete dissociation).

My score (fill in after checking the answers)

Section	A /8	B /8	C /10	D /16	E /18	F /20	G (bonus)	Total /80
My marks								

Rough work

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E. Short answer II (3 marks each)

1. 3.0 g of a non-volatile solute in 100 g water lowers the vapour pressure from 23.76 to 23.52 mm Hg. Find its molar mass.
2. 2.0 g of a solute in 50 g benzene lowers the freezing point by 0.82 K (K_f = 5.12). Find M.
3. Draw vapour pressure–composition curves for positive and negative deviations and explain.
4. Calculate π of a solution of 1.8 g glucose in 500 mL at 300 K (R = 0.0821).
5. A 0.1 m solution of a weak acid HA freezes at –0.205 °C. Find i and α (K_f = 1.86).
6. Derive the relation between relative lowering of vapour pressure and molar mass.

F. Long answer (5 marks each)

1. (a) State Henry's law and give two applications. (b) Explain bends and anoxia. (c) Find the mass of CO_2 dissolved in 1 L of water at 298 K under 2.5 atm (K_H = 1.67×10^8 Pa).
2. (a) Define van't Hoff factor. (b) What is i for solutes that associate and dissociate? (c) 1.0 g of KCl in 100 g water freezes at –0.48 °C. Find i and the degree of dissociation (K_f = 1.86).

G. HOTS (higher-order thinking)

1. Why does a bottle of soda fizz when opened?
2. Why is common salt spread on roads in snowy hills, while glycol goes in car radiators?
3. Two beakers, one of water and one of sugar solution, are kept under one bell jar. What happens after a long time?
4. Why should intravenous injections be isotonic with blood?

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20 MCQs

1. Unit of molality
(a) mol L^{-1} (b) g L^{-1} (c) mol kg^{-1} (d) mol

2. Temperature independent
(a) molality (b) molarity (c) normality (d) formality

3. Henry's law: p equals
(a) K_H/x (b) K_H + x (c) x/K_H (d) K_H x

4. Gas solubility on heating
(a) increases (b) decreases (c) unchanged (d) doubles

5. Ideal solution
(a) $\Delta H > 0$ (b) $\Delta H = 0$ (c) $\Delta H < 0$ (d) $\Delta V > 0$

6. Negative deviation pair
(a) ethanol + acetone (b) hexane + heptane (c) benzene + toluene (d) chloroform + acetone

7. Minimum-boiling azeotrope from
(a) positive deviation (b) negative deviation (c) ideal (d) none

8. Not colligative
(a) osmotic pressure (b) ΔT_f (c) vapour pressure (d) ΔT_b

9. K_f of water
(a) 0.52 (b) 2.53 (c) 5.12 (d) 1.86

10. i for $AlCl_3$ (complete)
(a) 2 (b) 4 (c) 3 (d) 1

11. Best for molar mass of proteins
(a) ΔT_b (b) ΔT_f (c) osmotic pressure (d) RLVP

12. RBC in hypertonic solution
(a) shrinks (b) swells (c) bursts (d) no change

13. RLVP equals
(a) x_1 (b) x_2 (c) p° (d) p

14. Complete dimerisation: i
(a) 2 (b) 1 (c) 0.5 (d) 0

15. Bends is related to
(a) Raoult's law (b) azeotrope (c) osmosis (d) Henry's law

16. Lowest freezing point (0.1 m)
(a) K_2SO_4 (b) NaCl (c) urea (d) glucose

17. Reverse osmosis is used for
(a) desalination (b) antifreeze (c) soda (d) welding

18. π equals
(a) CR/T (b) RT/C (c) C/RT (d) CRT

19. Association makes molar mass
(a) lower (b) higher (c) same (d) zero

20. H_2 in palladium is
(a) gas in liquid (b) solid in gas (c) gas in solid (d) liquid in solid

1 (c) · 2 (a) · 3 (d) · 4 (b) · 5 (b) · 6 (d) · 7 (a) · 8 (c) · 9 (d) · 10 (b) · 11 (c) · 12 (a) · 13 (b) · 14 (c) · 15 (d) · 16 (a) · 17 (a) · 18 (d) · 19 (b) · 20 (c)

Assertion–Reason (6)

(a) Both true, R explains A. (b) Both true, R does not explain A. (c) A true, R false. (d) A false, R true.

1. A: 0.1 M NaCl has a lower freezing point than 0.1 M glucose. R: NaCl gives more particles.
2. A: Ethanol + water cannot be purified beyond about 95% by distillation. R: They form a minimum-boiling azeotrope.
3. A: Molarity of a solution changes with temperature. R: Volume of a solution changes with temperature.
4. A: Gases are more soluble in hot water. R: Dissolving a gas is exothermic.
5. A: Chloroform and acetone show positive deviation. R: They form hydrogen bonds with each other.
6. A: Acetic acid in benzene shows a higher molar mass. R: It dimerises through hydrogen bonds.

Key: 1 (a) · 2 (a) · 3 (a) · 4 (d) · 5 (d) · 6 (a)

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Answers to the practice set

A. 1. kg 2. Henry 3. zero 4. negative 5. 1.86 6. isotonic 7. 3 8. cellulose acetate

B. 1. F 2. F 3. F 4. F 5. T 6. T 7. T 8. F

C. 1. $x_B = n_B/(n_A + n_B)$ 2. H_2 in Pd 3. depends only on the number of solute particles 4. ΔT_b of a 1 molal solution 5. same osmotic pressure 6. N_2 bubbles in blood when pressure drops 7. $x = p/K_H$ 8. ethanol + acetone 9. 0.5 10. it does not change with temperature

D. 1. p. 4–5 2. 0.05/0.25 = 0.2 m 3. p. 6 4. its lower vapour pressure must be raised to 1 atm 5. $0.1 \times 0.0821 \times 300 = 2.46$ atm 6. p. 6 7. dimerisation ($i = 0.5$) 8. $2 \times 1.86 \times 0.2 = 0.744$ K

E. 1. $0.24/23.76 = 3 \times 18/(M \times 100)$: $M = 53.5$ g mol^{-1} 2. $5.12 \times 2 \times 1000/(0.82 \times 50) = 250$ g mol^{-1} 3. p. 6 4. 0.02 M: 0.49 atm 5. $\Delta T_f(\text{calc}) = 0.186$; $i = 1.10$; $\alpha = 0.10$ 6. p. 7

F. 1. p. 4; $x = 1.52 \times 10^{-3}$, $n = 0.0844$ mol per L = 3.7 g 2. p. 9; $m = 0.134$; $\Delta T_f(\text{calc}) = 0.250$ K; $i = 0.48/0.250 = 1.92$; $\alpha = 0.92$

G. 1. the pressure falls, so less CO_2 stays dissolved (Henry) 2. salt is cheap and works on open roads but corrodes metal; glycol does not corrode and also raises the boiling point 3. water distils into the solution (higher p over water), so the water level falls and the solution level rises 4. a hypertonic or hypotonic drip would make blood cells shrink or swell

Case. (i) $m = 10/1.86 = 5.38$; $5.38 \times 62 = 333$ g (ii) salt corrodes the radiator (iii) yes: $0.52 \times 5.38 = 2.8$ K (iv) depression of freezing point

Last-night summary

- Concentration: mass %, ppm, x, M (changes with T), m (does not). Henry: $p = K_H x$; solubility falls as T rises; bends, anoxia, soda.
- Raoult: $p_i = p_i^\circ x_i$; ideal: $\Delta H = \Delta V = 0$. Positive deviation (weaker A–B; ethanol + acetone) ⟶ min-boiling azeotrope; negative (stronger A–B; $CHCl_3$ + acetone, phenol + aniline) ⟶ max-boiling.
- Colligative: RLVP = ix_2; $\Delta T_b = iK_b m$; $\Delta T_f = iK_f m$; $\pi = iCRT$; $K_f = RM_1T_f^2/1000\Delta H_{fus}$. Water: K_b 0.52, K_f 1.86.
- Osmosis: isotonic = same π; RBC shrinks in hypertonic, swells in hypotonic (0.9% NaCl is isotonic); reverse osmosis desalinates (cellulose acetate).
- i = M(normal)/M(observed): dissociation $i > 1$, $\alpha = (i - 1)/(n - 1)$; association $i < 1$, dimer $\alpha = 2(1 - i)$.

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