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CHAPTER NOTES

Solutions Class 12 Notes — CBSE Chemistry Chapter 2

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Solutions Class 12 Notes — CBSE Chemistry Chapter 2

Chapter 2 of Class 12 Chemistry deals with **Solutions** — one of the most important chapters for both Boards and competitive exams. This chapter covers how substances dissolve, how to express concentration, and the fascinating **colligative properties** that depend only on the number of solute particles, not their identity. Expect **5-7 marks** from this chapter, mainly from numericals on colligative properties.

Key Concepts

Types of Solutions

Solute	Solvent	Type	Example
Gas	Gas	Gaseous	Air (O_2 in N_2)
Gas	Liquid	Liquid	Soda water (CO_2 in water)
Liquid	Liquid	Liquid	Ethanol in water
Solid	Liquid	Liquid	Sugar in water
Solid	Solid	Solid	Alloys (brass = Cu + Zn)

Expressing Concentration

Molarity (M) = moles of solute / volume of solution (L)

Molality (m) = moles of solute / mass of solvent (kg)

Mole Fraction (x) = moles of component / total moles

Mass % = (mass of solute / mass of solution) × 100

ppm = (mass of solute / mass of solution) × 10^6

Why Molality is preferred: Molality doesn't change with temperature (it depends on mass, not volume), while Molarity does. This is why colligative property formulas use molality!

Solubility

Henry's Law (Gas in Liquid)

p = K_H × x

Where: p = partial pressure of gas, K_H = Henry's law constant, x = mole fraction of gas in solution

Higher K_H → lower solubility

Applications:

- Soda/cold drinks: CO_2 dissolved under high pressure (low solubility at atmospheric pressure → fizz)
- Scuba diving: High pressure → more N_2 dissolves in blood → "bends" on rapid ascent
- Oxygen dissolved at higher altitude decreases → altitude sickness

Factors Affecting Solubility

- **Gas in Liquid:** Solubility increases with pressure, decreases with temperature
- **Solid in Liquid:** Solubility generally increases with temperature (exceptions: $CeSO_4$, Li_2CO_3)

Raoult's Law

For volatile solute: $p_1 = x_1 \cdot p_1^0$ and $p_2 = x_2 \cdot p_2^0$

Total pressure: $P_{total} = p_1 + p_2 = x_1p_1^0 + x_2p_2^0$

For non-volatile solute: $p = x_{solvent} \cdot p^0_{solvent}$

Or: $(p^0 - p)/p^0 = x_{solute}$ (relative lowering of vapour pressure)

Ideal vs Non-ideal Solutions

Property	Ideal Solution	Positive Deviation	Negative Deviation
Obeys Raoult's Law?	Yes	P_observed > P_Raoult	P_observed < P_Raoult
ΔH_mix	0	Positive (endothermic)	Negative (exothermic)
ΔV_mix	0	Positive	Negative
A-B interactions vs A-A, B-B	Equal	Weaker	Stronger
Examples	Benzene + Toluene, n-hexane + n-heptane	Ethanol + Water, Acetone + CS_2	$CHCl_3$ + Acetone, HNO_3 + Water

Azeotropes: Solutions that boil at constant temperature and cannot be separated by distillation. Minimum boiling azeotrope → positive deviation (e.g., ethanol-water at 95.4%). Maximum boiling azeotrope → negative deviation (e.g., HNO_3-water at 68%).

Colligative Properties

Properties that depend only on the **number** of solute particles, not their nature:

1. Relative Lowering of Vapour Pressure (RLVP)

$(p^0 - p)/p^0 = x_{solute} = n_2/(n_1 + n_2) \approx n_2/n_1$ (for dilute solutions)

Where n_2 = moles of solute, n_1 = moles of solvent

2. Elevation of Boiling Point (ΔT_b)

$\Delta T_b = K_b \times m$

K_b = molal elevation constant (ebullioscopic constant)

m = molality

$\Delta T_b = (K_b \times w_2 \times 1000) / (M_2 \times w_1)$ (for finding molar mass M_2)

3. Depression of Freezing Point (ΔT_f)

$\Delta T_f = K_f \times m$

K_f = molal depression constant (cryoscopic constant)

$$\Delta T_f = (K_f \times w_2 \times 1000) / (M_2 \times w_1)$$

Practical Application: This is why we add salt to icy roads — salt lowers the freezing point of water! Also why antifreeze (ethylene glycol) is added to car radiators in cold regions.

4. Osmotic Pressure (π)

$\pi = CRT = (n/V)RT$

C = molarity, R = 0.0821 L·atm/mol·K, T = temperature in K

For molar mass: $M_2 = (w_2RT) / (\pi V)$

Osmosis: Flow of solvent from lower concentration to higher concentration through a semipermeable membrane.

- **Isotonic solutions:** Same osmotic pressure (e.g., saline drip = 0.9% NaCl)
- **Hypertonic:** Higher osmotic pressure → cells shrink (crenation)
- **Hypotonic:** Lower osmotic pressure → cells swell/burst (haemolysis)
- **Reverse osmosis:** Apply pressure > π to force solvent from concentrated to dilute side (used in water purification)

van't Hoff Factor (i)

i = observed colligative property / calculated colligative property

i = total particles after dissociation or association / particles before

For electrolytes that dissociate: $i > 1$ (e.g., NaCl → $i \approx 2$, $CaCl_2$ → $i \approx 3$)

For solutes that associate: $i < 1$ (e.g., acetic acid in benzene → $i \approx 0.5$)

Modified formulas: $\Delta T_b = i \cdot K_b \cdot m$, $\Delta T_f = i \cdot K_f \cdot m$, $\pi = iCRT$

Important Definitions

Term	Definition
Solution	Homogeneous mixture of two or more substances
Molality	Moles of solute per kilogram of solvent
Mole Fraction	Ratio of moles of a component to total moles
Henry's Law	Partial pressure of a gas $\propto$ its mole fraction in solution
Raoult's Law	Vapour pressure of a component = mole fraction × pure vapour pressure
Ideal Solution	Solution that obeys Raoult's law at all concentrations
Azeotrope	Constant-boiling mixture that cannot be separated by distillation
Colligative Property	Property depending only on number of solute particles, not their nature
Osmotic Pressure	Minimum pressure needed to prevent osmosis
van't Hoff Factor	Ratio of observed to calculated colligative property

Solved Examples — NCERT Based

Example 1: Molality Calculation

Q: Calculate the molality of a solution containing 20 g of NaOH dissolved in 500 g of water. (M of NaOH = 40)

Solution:

Moles of NaOH = 20/40 = 0.5 mol

Mass of solvent = 500 g = 0.5 kg

Molality = 0.5/0.5 = **1 m (molal)**

Example 2: Boiling Point Elevation

Q: 18 g of glucose (M = 180) is dissolved in 500 g of water. Calculate the boiling point of the solution. (K_b for water = 0.52 K·kg/mol)

Solution:

Moles of glucose = 18/180 = 0.1 mol

Molality = 0.1/0.5 = 0.2 m

$\Delta T_b = K_b \times m = 0.52 \times 0.2 = 0.104$ K

Boiling point = 100 + 0.104 = **100.104°C**

Example 3: Osmotic Pressure

Q: 200 mL of an aqueous solution of a protein contains 1.26 g of protein. The osmotic pressure at 300 K is 2.57×10^{-3} atm. Calculate the molar mass of the protein.

Solution:

$\pi = CRT \rightarrow C = \pi/RT = (2.57 \times 10^{-3}) / (0.0821 \times 300) = 1.044 \times 10^{-4}$ mol/L

Moles = $C \times V = 1.044 \times 10^{-4} \times 0.2 = 2.088 \times 10^{-5}$ mol

M = mass/moles = $1.26 / (2.088 \times 10^{-5})$ = **60,345 g/mol ≈ 6.03×10^4 g/mol**

Example 4: van't Hoff Factor

Q: 0.1 m NaCl solution has a freezing point depression of 0.348°C. Calculate the van't Hoff factor. (K_f = 1.86 K·kg/mol)

Solution:

Expected $\Delta T_f = K_f \times m = 1.86 \times 0.1 = 0.186$°C

i = Observed ΔT_f / Expected $\Delta T_f = 0.348 / 0.186$ = **1.87**

(Close to 2, as expected for $NaCl \rightarrow Na^+ + Cl^-$, though not fully dissociated)

Important Questions for Board Exams

1 Mark Questions

1. Define molality.
2. State Henry's law.
3. What are isotonic solutions?
4. Why is osmotic pressure preferred for molar mass of polymers?
5. What is the van't Hoff factor for Na_2SO_4 if completely dissociated?

2 Mark Questions

1. Distinguish between ideal and non-ideal solutions.
2. Why do gases become less soluble at higher temperatures?
3. Explain reverse osmosis and its application.
4. Calculate mole fraction of solute and solvent if 10 g of NaCl is dissolved in 90 g of water.

3 Mark Questions

1. State Raoult's law. Explain positive and negative deviations with examples.
2. Derive the relationship between relative lowering of vapour pressure and mole fraction of solute.
3. Calculate the freezing point of a solution containing 60 g of glucose in 250 g of water.
4. What are azeotropes? Explain minimum and maximum boiling azeotropes.

5 Mark Questions

1. Describe all four colligative properties with formulas and examples. How is van't Hoff factor used to modify them?
2. What is osmosis? Explain osmotic pressure, its measurement, and applications. Define isotonic, hypertonic and hypotonic solutions.

Quick Revision Points

- Molality is temperature-independent; molarity is not
- Henry's law: $p = K_H \times x$ (higher $K_H \rightarrow$ less soluble)
- Raoult's law: $p = x \times p^0$ (ideal solutions obey this)
- Positive deviation: weaker A-B forces, $\Delta H > 0$ | Negative: stronger A-B, $\Delta H < 0$
- 4 colligative properties: RLVP, ΔT_b, ΔT_f, π — all depend on number of particles only

- $\Delta T_b = K_b \cdot m$ | $\Delta T_f = K_f \cdot m$ | $\pi = CRT$
- For electrolytes: multiply by van't Hoff factor (i)
- NaCl: $i \approx 2$ | $CaCl_2$: $i \approx 3$ | Glucose: $i = 1$
- Osmotic pressure is best for molar mass of macromolecules (measurable even at low conc.)
- Reverse osmosis: pressure > π → pure water through membrane

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