▶ Video explanation (YouTube Short)

Solutions: Chapter Introduction

1. Two everyday mysteries

Soda fizzes the moment the bottle is opened, and sea water does not freeze at 0 °C. Both are explained by the behaviour of **solutions**.

2. Solution, solvent, solute

A **solution** is a homogeneous mixture of two or more components: its composition is the same throughout. The component present in the largest amount is the **solvent**; the others are **solutes**. In salt water, water is the solvent and salt the solute. Solutions can be gaseous (air), liquid (salt water, soda) or solid (alloys such as brass).

3. Concentration

Term	Definition	Unit
Mass %	(mass of solute / mass of solution) × 100	%
ppm	parts of solute per 10^6 parts of solution	ppm
Mole fraction x	moles of component / total moles	none
Molarity M	moles of solute / litres of solution	mol L^{-1}
Molality m	moles of solute / kg of solvent	mol kg^{-1}

Molality does not change with temperature (it uses mass); molarity does (volume changes).

Solved example. 4 g NaOH (40 g mol^{-1}) in 500 g water: n = 0.1 mol; m = 0.1 / 0.5 = **0.2 mol kg^{-1}**.

4. Henry's law: why soda fizzes

p = K_H · x: the partial pressure of a gas above a solution is proportional to its mole fraction in the solution. Higher pressure → more gas dissolves.

Soda is bottled under high CO_2 pressure. Opening the bottle lowers the pressure, the solubility of CO_2 falls, and the extra gas escapes as bubbles. (Divers breathe helium-diluted air because more nitrogen dissolves in blood at depth.)

5. Raoult's law and vapour pressure

For a solution of volatile liquids, the partial vapour pressure of each component = its mole fraction × its pure vapour pressure ($p_1 = x_1 p_1^\circ$). With a **non-volatile solute**: $(p^\circ - p) / p^\circ = x$(solute).

The solute particles occupy part of the surface, so fewer solvent molecules escape and the vapour pressure falls. Solutions that obey Raoult's law at all concentrations are **ideal**; others show positive or negative deviations (**non-ideal**).

6. Colligative properties

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

- Relative lowering of vapour pressure.
- Elevation of boiling point ($\Delta T_b = K_b m$).
- Depression of freezing point ($\Delta T_f = K_f m$): sea water freezes at about −1.9 °C.
- Osmotic pressure ($\pi = CRT$): solvent flows through a semipermeable membrane into the more concentrated solution.

Key takeaways

- Solvent = largest component; solute = dissolved.
- Molarity uses litres of solution, molality kg of solvent.
- Henry: gas solubility $\propto$ pressure. Raoult: solute lowers vapour pressure.
- Colligative properties count particles.

Practice questions (CBSE style)

1. (MCQ) Which concentration term does not change with temperature? (a) molarity (b) molality (c) normality (d) % by volume
2. (MCQ) Which is a colligative property? (a) colour (b) osmotic pressure (c) density (d) refractive index
3. (Very short) State Henry's law.
4. (Very short) Why does salt water boil above 100 °C?
5. (Short) Explain why soda fizzes when opened.
6. (Numerical) Find the molarity of 0.5 mol NaCl in 2 L of solution.
7. (Numerical) 18 g glucose (180 g mol^{-1}) in 90 g water (18 g mol^{-1}): find the mole fraction of glucose.
8. (Application) Why is salt spread on icy roads?

Answer key

1. (b) molality.
2. (b).
3. The partial pressure of a gas is proportional to its mole fraction in the solution: $p = K_H x$.
4. The non-volatile solute lowers the vapour pressure, so a higher temperature is needed to reach atmospheric pressure (elevation of boiling point).
5. CO_2 is dissolved under high pressure; opening lowers the pressure, so its solubility falls and the excess escapes as bubbles (Henry's law).
6. M = 0.5 / 2 = 0.25 M.
7. n(glucose) = 0.1, n(water) = 5; x = 0.1 / 5.1 ≈ 0.0196.
8. Dissolved salt depresses the freezing point of water, so the ice melts at temperatures below 0 °C.

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