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Solutions and Concentration

Definition and Constituents

A solution is a homogeneous mixture of two or more pure substances. It consists of two main constituents:

Solute: The substance present in lesser amount that dissolves in the solvent (e.g., sugar in water).

Solvent: The substance present in larger amount in which the solute dissolves (e.g., water).

Types of Solutions

Solutions can be classified based on the physical states of solute and solvent. There are nine types, such as solid in liquid (salt in water), liquid in liquid (alcohol in water), gas in liquid

(oxygen in water), solid in solid (brass), and gas in gas (oxygen in nitrogen).

Aqueous solutions contain water as solvent, while non-aqueous solutions have solvents other than water.

Saturated solutions cannot dissolve more solute at a given temperature, whereas unsaturated solutions can.

Methods of Expressing Concentration

Concentration indicates the amount of solute in a given quantity of solution or solvent. Common expressions include:

Mass percentage (w/w): (Mass of solute / Total mass of solution) × 100

Volume percentage (v/v): (Volume of solute / Total volume of solution) × 100

Mass by volume percentage (w/v): (Mass of solute / Volume of solution) × 100

Parts per million (ppm): (Mass or volume of component / Total mass or volume of solution) × 10^6

Mole fraction: Ratio of moles of a component to total moles in solution. Sum of mole fractions equals one.

Molarity (M): Moles of solute per litre of solution (mol L^{-1}).

Molality (m): Moles of solute per kilogram of solvent (mol kg^{-1}).

Normality (N): Gram equivalents of solute per litre of solution.

Relationships between these concentration units depend on solution density and molar masses.

Solubility and Henry’s Law

Solubility is the maximum amount of solute dissolved in 100 g of solvent at a given temperature.

Factors affecting solubility include the nature of solute and solvent ("like dissolves like"), temperature (generally increases solubility for endothermic dissolution), and pressure (significant for gases).

Henry’s Law states that the mass of a gas dissolved in a liquid at constant temperature is proportional to the pressure of the gas above the liquid. Mathematically, $p = K_H \times x$, where p is partial pressure, x is mole fraction of gas in solution, and K_H is Henry’s constant.

Applications include carbonation of soft drinks, scuba diving gas mixtures, and effects of altitude on oxygen availability.

Solved Examples

Example: Calculate the mole fraction and amount of nitrogen gas dissolved in 1 litre of water at 293 K with a partial pressure of 0.987 bar. Given Henry’s constant $K_H = 76,480$ bar.

Solution:

Using Henry's law: $x = p / K_H = 0.987 / 76{,}480 = 1.29 \times 10^{-5}$

Moles of nitrogen in 1 litre = $x \times 1000$ mmol = $1.29 \times 10^{-5} \times 1000 = 0.0129$ mmol

Therefore, 0.0129 millimoles of nitrogen dissolve in 1 litre of water under these conditions.

Practice Set

Conceptual Questions

- Explain the difference between molarity and molality.
- Why does Henry's law apply only under certain conditions?

Application Question

- Calculate the molarity of a solution prepared by dissolving 58.5 g of NaCl (molar mass = 58.5 g/mol) in 500 mL of solution.

Answer Key

Conceptual Questions

- Molarity is moles of solute per litre of solution and depends on volume, which changes with temperature. Molality is moles of solute per kilogram of solvent and is independent of temperature.
- Henry's law applies when gas pressure is not too high, temperature is moderate, and the gas does not chemically react or dissociate in solution.

Application Question

- Moles of NaCl = 58.5 / 58.5 = 1 mole
- Volume of solution = 0.5 L
- Molarity = 1 mole / 0.5 L = 2 M

Raoult’s Law

Vapour Pressure and Raoult’s Law

Vapour pressure is the pressure exerted by vapour in equilibrium with its liquid at a constant temperature. It depends on the nature of the liquid and temperature.

Raoult’s law states that the partial vapour pressure of each volatile component in a solution is proportional to its mole fraction.

For component A: $p_A = p°_A \times x_A$, where $p°_A$ is vapour pressure of pure A and x_A is mole fraction in solution.

For non-volatile solutes, the relative lowering of vapour pressure equals the mole fraction of solute.

Ideal solutions obey Raoult’s law over all concentrations, with similar intermolecular forces between solute and solvent.

Non-ideal solutions deviate positively or negatively due to differences in intermolecular forces.

Positive deviation: weaker A–B interactions, higher vapour pressure.

Negative deviation: stronger A–B interactions, lower vapour pressure.

Azeotropes are mixtures that boil without change in composition, formed by strong deviations.

Solved Examples

Example: Calculate the total vapour pressure of a solution containing two volatile liquids A and B with mole fractions $x_A = 0.6$ and $x_B = 0.4$. Vapour pressures of pure A and B are 400 mmHg and 300 mmHg respectively.

Solution:

$p_A = 400 \times 0.6 = 240$ mmHg

$p_B = 300 \times 0.4 = 120$ mmHg

Total vapour pressure = 240 + 120 = 360 mmHg

Practice Set

Conceptual Questions

- What conditions define an ideal solution?
- Explain the difference between positive and negative deviation from Raoult’s law.

Application Question

- A solution contains 0.3 mole fraction of solute B and 0.7 mole fraction of solvent A. If vapour pressure of pure A is 500 mmHg and pure B is 200 mmHg, calculate the total vapour pressure assuming ideal behavior.

Answer Key

Conceptual Questions

- Ideal solutions have similar intermolecular forces between solute-solute, solvent-solvent, and solute-solvent, and obey Raoult’s law at all concentrations.
- Positive deviation occurs when A–B interactions are weaker than A–A or B–B, increasing vapour pressure; negative deviation occurs when A–B interactions are stronger, decreasing vapour pressure.

Application Question

- $p_A = 500 \times 0.7 = 350$ mmHg
- $p_B = 200 \times 0.3 = 60$ mmHg
- Total vapour pressure = 350 + 60 = 410 mmHg

Colligative Properties

Definition and Types

Colligative properties depend on the number of solute particles, not their nature. They include:

Relative lowering of vapour pressure

Elevation of boiling point

Depression of freezing point

Osmotic pressure

Relative lowering of vapour pressure is equal to the mole fraction of solute.

Elevation of boiling point (ΔT_b) is proportional to molality: $\Delta T_b = K_b \times m$, where K_b is ebullioscopic constant.

Depression of freezing point (ΔT_f) is proportional to molality: $\Delta T_f = K_f \times m$, where K_f is cryoscopic constant.

Osmosis is the flow of solvent through a semipermeable membrane to a solution; osmotic pressure (π) is the pressure needed to stop this flow.

Osmotic pressure for dilute solutions: $\pi = C R T$, where C is molar concentration, R is gas constant, and T is temperature.

Reverse osmosis uses pressure greater than osmotic pressure to purify water.

Abnormal molecular mass occurs when calculated molecular mass differs from theoretical due to association or dissociation.

Van’t Hoff factor (i) is the ratio of observed to calculated colligative property, accounting for ionization or association.

Solutions can be hypertonic, hypotonic, or isotonic based on relative solute concentrations separated by semipermeable membranes.

Solved Examples

Example: Calculate the molecular mass of a compound if 0.721 g lowers the freezing point of 75 mL benzene from 5.51 °C to 5.03 °C. Given density of benzene = 0.879 g/mL, K_f = 5.12 K kg mol^{-1}.

Solution:

Mass of benzene = 0.879 × 75 = 65.925 g = 0.06593 kg

ΔT_f = 5.51 - 5.03 = 0.48 °C

Molality m = ΔT_f / K_f = 0.48 / 5.12 = 0.09375 mol/kg

Moles of solute = m × mass solvent (kg) = 0.09375 × 0.06593 = 0.00618 mol

Molecular mass = mass solute / moles = 0.721 / 0.00618 = 116.65 g/mol

Practice Set

Conceptual Questions

- Why do colligative properties depend only on the number of solute particles?
- Explain the significance of Van’t Hoff factor.

Application Question

- Calculate the osmotic pressure of a 0.1 M glucose solution at 27 °C. (R = 0.0821 L atm mol^{-1} K^{-1})

Answer Key

Conceptual Questions

- Colligative properties depend on the number of solute particles because they affect the physical properties of the solvent, regardless of solute identity.
- Van’t Hoff factor accounts for dissociation or association of solutes, adjusting colligative property calculations to observed values.

Application Question

- Temperature in Kelvin = 27 + 273 = 300 K
- $\pi = C\ R\ T = 0.1 \times 0.0821 \times 300 = 2.463$ atm

Quick Reference Table

Common Mistakes and Misconceptions

Glossary