

Colligative (collective) properties are vapor pressure lowering, boiling point elevation (ΔT_b), freezing point depression (ΔT_f), and osmotic pressure (Π). **Colligative properties depend only on the number of solute particles and not their chemical identity.** This is best, specifically, the properties of a pure solvent compare with those of a solution:

- The vapor pressure of the solution is lower.
- The boiling point of the solution is higher.
- The freezing (or melting) point of the solution is lower.
- The solution causes osmosis, the flow of solvent from a solution of lower solute concentration to one of higher solute concentration a semipermeable membrane.

Below are the key equations you need for solving problems related to colligative properties.

Key Equations for Colligative Properties

Raoult's Law: Relationship between the Vapor Pressure of a Solution (P_{solution}), the Mole Fraction of the Solvent (χ_{solvent}), and the Vapor Pressure of the Pure Solvent ($P^{\circ}_{\text{solvent}}$)

$$P_{\text{solution}} = \chi_{\text{solvent}} P^{\circ}_{\text{solvent}}$$

$$K_{\text{solvent}} = \frac{V_{\text{solvent}}}{n_{\text{solvent}} + n_{\text{solute}}}$$

The Vapor Pressure of a Solution Containing Two Volatile Components

$$P_A = \chi_A P^{\circ}_A \quad P_B = \chi_B P^{\circ}_B$$

$$P_{\text{total}} = P_A + P_B$$

Relationship between Osmotic Pressure (Π), Molarity (M), the Ideal Gas Constant (R), and Temperature (T , in K)

$$\Pi = MRT$$

$$M = \frac{n_{\text{solute}}}{V_{\text{solution}}} = 0.08206 \text{ L}\cdot\text{atm} / \text{mol}\cdot\text{K}$$

The dependence of ΔT_f and ΔT_b on the Molarity (m) and Freezing Point Depression Constant (K_f) and Boiling Point Elevation Constant (K_b) respectively:

$$\Delta T_f = m \cdot K_f$$

$$\Delta T_b = m \cdot K_b$$

$$m = \frac{\text{mol (solute)}}{\text{kg (solvent)}}$$

Boiling-Point and Freezing-Point Constants

| Solvent | Freezing point (°C) | K_f (K·kg·mol ⁻¹) | Boiling point (°C) | K_b (K·kg·mol ⁻¹) |
|----------------------|---------------------|---------------------------------|--------------------|---------------------------------|
| acetone | -93.85 | 2.00 | 56.5 | 0.71 |
| benzene | 5.5 | 5.12 | 80.1 | 2.53 |
| carbon tetrachloride | -23 | 29.8 | 76.5 | 4.93 |
| cyclohexane | 6.5 | 20.1 | 80.7 | 2.79 |
| naphthalene | 80.3 | 6.94 | 217.7 | 5.80 |
| phenol | 41 | 7.27 | 182 | 3.04 |
| water | 0 | 1.86 | 100.0 | 0.51 |

An important trait in colligative properties is, the larger the number of the nodes, the stronger the effect. For example, dissolving 2 moles of methanol (nonelectrolyte) will not affect the colligative properties of the solution as much as 2 moles of NaCl (electrolyte) would. This is because NaCl dissociates into ions, so 2 moles of NaCl will produce 4 moles of particles (ions).



Electrolytes have stronger effect on colligative properties of solutions.

Therefore, we divide our summary in two parts: colligative properties of nonelectrolytes solutions and those of electrolyte solutions.

Nonelectrolyte Solutions

Vapor Pressure Lowering Example: Calculate the vapor pressure of a solution at 25°C that is made by adding 47.6 g of glucose ($C_6H_{12}O_6$) to 340.0 g of water. The vapor pressure of pure water at 25°C is 23.8 torr.

The vapor pressure of the solution is calculated by the following formula:

$$P_{\text{solution}} = \chi_{\text{solvent}} P^{\circ}_{\text{solvent}}$$

We have the vapor pressure of pure water: $P^{\circ}(\text{H}_2\text{O}) = 23.8 \text{ Torr}$, so the only missing part is the mole fraction of the solvent (water in this case) which is the ratio of the moles of solvent over the total moles of the solvent and the solute.

Calculate the moles and plug the numbers on the equation for the mole fraction of H_2O :

$$\chi_{\text{H}_2\text{O}} = \frac{n(\text{H}_2\text{O})}{n(\text{H}_2\text{O}) + n(\text{C}_6\text{H}_{12}\text{O}_6)} = \frac{18.9 \text{ mol}}{(18.9 + 0.298) \text{ mol}} = 0.868$$

The vapor pressure of the solution would be: $P_{\text{solution}} = 0.868 \times 23.8 \text{ torr} = 20.6 \text{ torr}$

Freezing Point Depression Example: Using the appropriate data in the table, determine the freezing point depression of the solution that contains 24.1 g urea (NH_2CO) in 485 mL of water.

The freezing point depression is calculated by the following formula: $\Delta T_f = m \cdot K_f$

So, we need to first determine the moles of urea, then calculate the molarity and use it in the equation for the ΔT_f .

$$n(\text{NH}_2\text{CO}) = 24.1 \cancel{\text{ g}} \times \frac{1 \text{ mol}}{60.1 \cancel{\text{ g}}} = 0.401 \text{ mol}$$

$$m(\text{H}_2\text{O}) = 485 \cancel{\text{ mL}} \times \frac{1.00 \cancel{\text{ g}}}{1 \cancel{\text{ mL}}} \times \frac{1 \text{ kg}}{1000 \cancel{\text{ g}}} = 0.485 \text{ kg}$$

$$m = \frac{0.401 \text{ mol}}{0.485 \text{ kg}} = 0.827 \text{ m}$$

$$\Delta T_f = K_{fm} = 1.86 \text{ }^\circ\text{C} \cdot \text{m} \times 0.827 \text{ m} = 0.591 \text{ }^\circ\text{C}$$

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