

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 in 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)

$$\Pi = MRT$$

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

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	3.84
water	0	1.86	100.0	0.51

An important trait in colligative properties is, the larger the number of the moles, 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 :

$$\chi_{\text{H}_2\text{O}} = \frac{n(\text{H}_2\text{O})}{n(\text{H}_2\text{O}) + n(\text{C}_6\text{H}_{12}\text{O}_6)} = \frac{18.9 \text{ mol}}{(18.9 + 0.266) \text{ mol}} = 0.986$$

The vapor pressure of the solution would be: $P_{\text{solution}} = 0.986 \times 23.8 \text{ torr} = 23.5 \text{ torr}$

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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Table of Contents Colligative Properties Study Guide Answers

1. Understanding the eBook Colligative Properties Study Guide Answers
 - The Rise of Digital Reading Colligative Properties Study Guide Answers
 - Advantages of eBooks Over Traditional Books
2. Identifying Colligative Properties Study Guide Answers
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in an Colligative Properties Study Guide Answers
 - User-Friendly Interface

4. Exploring eBook Recommendations from Colligative Properties Study Guide Answers
 - Personalized Recommendations
 - Colligative Properties Study Guide Answers User Reviews and Ratings
 - Colligative Properties Study Guide Answers and Bestseller Lists
5. Accessing Colligative Properties Study Guide Answers Free and Paid eBooks
 - Colligative Properties Study Guide Answers Public Domain eBooks
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 - Highlighting and Note-Taking Colligative Properties Study Guide Answers
 - Interactive Elements Colligative Properties Study Guide Answers
8. Staying Engaged with Colligative Properties Study Guide Answers
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Colligative Properties Study Guide Answers
9. Balancing eBooks and Physical Books Colligative Properties Study Guide Answers
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Colligative Properties Study Guide Answers
10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
11. Cultivating a Reading Routine Colligative Properties Study Guide Answers
 - Setting Reading Goals Colligative Properties Study Guide Answers
 - Carving Out Dedicated Reading Time

12. Sourcing Reliable Information of Colligative Properties Study Guide Answers
 - Fact-Checking eBook Content of Colligative Properties Study Guide Answers
 - Distinguishing Credible Sources
13. Promoting Lifelong Learning
 - Utilizing eBooks for Skill Development
 - Exploring Educational eBooks
14. Embracing eBook Trends
 - Integration of Multimedia Elements
 - Interactive and Gamified eBooks

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