What Is The Freezing Point Of A Solution In Which 2.50 Grams Of Sodium Chloride Are Added To 230.0 Ml

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Understanding the freezing point depression of a solution involves exploring the principles of colligative properties, the nature of sodium chloride (NaCl) as a solute, and the calculations needed to determine the new freezing point when NaCl is dissolved in water. This article provides an in-depth analysis of how to calculate the freezing point of a solution formed by dissolving 2.50 grams of sodium chloride into 230.0 mL of water, elucidating the underlying concepts, necessary formulas, step-by-step calculations, and implications.

Fundamentals of Freezing Point Depression

What Are Colligative Properties?

Colligative properties are physical properties of solutions that depend solely on the number of solute particles present, regardless of their identity. These include boiling point elevation, vapor pressure lowering, osmotic pressure, and freezing point depression.

The key idea is that dissolving solutes into a solvent alters the solvent’s physical properties because of the change in particle concentration, particularly affecting the solvent's ability to form a solid lattice during freezing.

The Concept of Freezing Point Depression

Freezing point depression refers to the lowering of a solvent's freezing point caused by the addition of a solute. When a non-volatile solute such as sodium chloride dissolves in water, it disrupts the formation of the orderly crystalline structure of ice, requiring a lower temperature to freeze the solution compared to pure water.

The mathematical relationship governing freezing point depression is expressed as:

$$\Delta T_f = K_f \times m \times i$$

where:

- ΔT_f is the freezing point depression (in °C),
- K_f is the cryoscopic constant (freezing point depression constant) for the solvent,
- m is the molality of the solution (moles of solute per kilogram of solvent),
- i is the van 't Hoff factor, representing the number of particles the solute dissociates into in solution.

Calculating the Freezing Point of the Sodium Chloride Solution

To determine the new freezing point, several steps are necessary: calculating the molality of the solution, identifying the van 't Hoff factor, and applying the freezing point depression formula.

Step 1: Determine the Mass of Water (Solvent)

Given:
- Volume of water = 230.0 mL
- Density of water ≈ 1.00 g/mL (at room temperature)

Therefore:
- Mass of water = volume × density = 230.0 mL × 1.00 g/mL = 230.0 g

Convert this to kilograms:
- 230.0 g = 0.230 kg

Step 2: Calculate the Moles of Sodium Chloride

Given:
- Mass of NaCl = 2.50 g
- Molar mass of NaCl = 58.44 g/mol

Calculate:
$$ \text{Moles of NaCl} = \frac{\text{Mass}}{\text{Molar mass}} = \frac{2.50\, g}{58.44\, g/mol} \approx 0.0428\, \text{mol} $$

Step 3: Determine the Molality of the Solution

Molality (m) is moles of solute per kilogram of solvent:
$$ m = \frac{\text{Moles of NaCl}}{\text{Mass of water in kg}} = \frac{0.0428\, \text{mol}}{0.230\, \text{kg}} \approx 0.186\, \text{mol/kg} $$

Step 4: Identify the Van 't Hoff Factor (i)

NaCl dissociates into sodium (Na^+) and chloride (Cl^-) ions in water:

$$ \text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^- $$

Therefore:
- $i \approx 2$

In practice, slight deviations may occur due to ion pairing, but for dilute solutions, assuming $i=2$ is acceptable.

Step 5: Calculate the Freezing Point Depression (ΔT_f)

Using the known value for water's cryoscopic constant:
- $K_f = 1.86\, ^\circ C\, \text{kg/mol}$

Calculate:

$$ \Delta T_f = K_f \times m \times i = 1.86\, ^\circ C \times 0.186\, \text{mol/kg} \times 2 \approx 0.693\, ^\circ C $$

Determine the New Freezing Point

Pure water freezes at 0.00°C. The addition of sodium chloride causes the freezing point to decrease by approximately 0.693°C:

$$ \text{Freezing point of solution} = 0.00^\circ C - \Delta T_f \approx -0.693^\circ C $$

Thus, the solution's freezing point is approximately -0.69°C.

Implications and Real-World Applications

Practical Significance

Understanding freezing point depression is crucial in various fields, including:
- Food preservation: Adding salt to ice in ice cream making or to brine for pickling.
- De-icing roads: Salt lowers the freezing point of water, preventing ice formation.
- Chemical processes: Controlling phase transitions in reactions and manufacturing.

Limitations and Considerations

While the calculation provides a good approximation, real-world factors can influence the actual freezing point:
- Ion pairing or clustering can reduce the effective number of particles.
- Temperature and pressure conditions may vary.
- The solution's concentration might approach saturation, affecting dissociation.

Conclusion

In summary, dissolving 2.50 grams of sodium chloride into 230.0 mL (or approximately 0.230 kg) of water results in a solution with a freezing point lowered to about -0.69°C. This calculation hinges on the principles of colligative properties, molality, and dissociation behavior of NaCl. Such understanding is fundamental in scientific and industrial applications where precise control of phase changes is required. By mastering these calculations, one can predict and manipulate the physical properties of solutions in various contexts.

Frequently Asked Questions

What is the freezing point depression caused by adding 2.50 grams of sodium chloride to 230.0 mL of solution?

The freezing point depression can be calculated using the formula $\Delta T_f = i\ K_f\ m$. First, determine molality: moles of NaCl = 2.50 g / 58.44 g/mol ≈ 0.0428 mol. Convert volume to kilograms of solvent (assuming density similar to water, 230.0 mL ≈ 0.230 kg). Molality m ≈ 0.0428 mol / 0.230 kg ≈ 0.186 mol/kg. Using Kf for water (1.86°C·kg/mol) and i=2 for NaCl, $\Delta T_f \approx 2\ 1.86\ 0.186 \approx 0.692$°C. Therefore, the solution's freezing point is lowered by approximately 0.69°C.

How do you determine the new freezing point of a solution after adding sodium chloride?

To determine the new freezing point, calculate the freezing point depression using the formula $\Delta T_f = i\ K_f\ m$, then subtract this value from the pure solvent's freezing point (0°C for water). For example, with 2.50 g of NaCl in 230.0 mL of water, the depression is approximately 0.69°C, so the new freezing point is about -0.69°C.

Why does adding sodium chloride lower the freezing point of a solution?

Adding sodium chloride introduces dissolved ions into the solution, which disrupt the formation of the solid crystalline structure of the solvent (like water). This phenomenon, called freezing point depression, occurs because the presence of solutes reduces the solvent's tendency to freeze, requiring a lower temperature to solidify.

What assumptions are made when calculating the freezing point of a salt solution like this?

Assumptions include that the solvent is water with a density close to 1 g/mL, the solution behaves ideally, complete dissociation of NaCl occurs, and temperature effects are uniform throughout the solution. Additionally, it is assumed that no other solutes are present and that the solution's volume does not significantly change upon dissolving the salt.

How does the molality of sodium chloride affect the freezing point of the solution?

The molality, which is moles of solute per kilogram of solvent, directly influences the freezing point depression. Higher molality results in a greater depression, meaning the solution freezes at a lower temperature. In this case, approximately 0.186 mol/kg of NaCl causes a depression of about 0.69°C.

Additional Resources

Understanding the Freezing Point of a Sodium Chloride Solution: A Detailed Exploration

When it comes to solutions and their physical properties, the concept of freezing point depression is both fundamental and fascinating. In this comprehensive review, we will explore the question: What is the freezing point of a solution in which 2.50 grams of sodium chloride (NaCl) are added to 230.0 mL of water? To answer this precisely, we will delve into the principles of colligative properties, the calculations involved, and the factors influencing the freezing point of such a solution.

Introduction to Colligative Properties and

Freezing Point Depression

Colligative properties are properties of solutions that depend solely on the number of solute particles present, not their identity or chemical nature. These include vapor pressure lowering, boiling point elevation, osmotic pressure, and freezing point depression.

Freezing point depression (ΔTf) refers to the lowering of the solvent's freezing point caused by the addition of solutes. This phenomenon is particularly important in various fields, including chemistry, environmental science, and industry.

Key concept: When a non-volatile solute such as NaCl dissolves in water, it disrupts the formation of an orderly solid lattice, requiring a lower temperature to freeze.

Theoretical Foundations of Freezing Point Depression

1. The Formula for Freezing Point Depression

The quantitative relationship is expressed as:

$$
\Delta T_f = i \times K_f \times m
$$

Where:
- ΔT_f = freezing point depression (in °C)
- i = van 't Hoff factor (number of particles the solute dissociates into)
- K_f = cryoscopic constant (freezing point depression constant for the solvent)
- m = molality of the solution (moles of solute per kilogram of solvent)

Understanding each component:
- The van 't Hoff factor (i) accounts for the dissociation of ionic compounds like NaCl.
- The cryoscopic constant (K_f) is specific to the solvent; for water, it is approximately 1.86 °C·kg/mol.
- Molality (m) is a measure of concentration independent of volume, making it suitable for colligative property calculations.

Calculating the Number of Moles of NaCl

Step 1: Determine the molar mass of NaCl

- Sodium (Na): approximately 22.99 g/mol
- Chlorine (Cl): approximately 35.45 g/mol

$$
\text{Molar mass of NaCl} = 22.99 + 35.45 = 58.44\, \text{g/mol}
$$

Step 2: Convert the given mass to moles

Given:
- Mass of NaCl = 2.50 g

$$
\text{Moles of NaCl} = \frac{2.50\, \text{g}}{58.44\, \text{g/mol}} \approx 0.0428\, \text{mol}
$$

Determining the Mass and Moles of Water (Solvent)

Step 1: Convert the volume of water to mass

Assuming the density of water is approximately 1 g/mL:

$$
\text{Mass of water} = 230.0\, \text{mL} \times 1\, \text{g/mL} = 230.0\, \text{g}
$$

Step 2: Convert grams to kilograms

$$
\text{Mass in kg} = 230.0\, \text{g} \div 1000 = 0.230\, \text{kg}
$$

Step 3: Calculate molality

\[

$$
m = \frac{\text{moles of solute}}{\text{kg of solvent}} = \frac{0.0428\, \text{mol}}{0.230\, \text{kg}} \approx 0.186\, \text{mol/kg}
$$

Accounting for Dissociation of Sodium Chloride

NaCl is an ionic compound that dissociates completely in water:

$$
\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^-
$$

This dissociation results in two particles per formula unit.

van 't Hoff factor (i):

$$
i = 2
$$

This effectively doubles the number of particles contributing to colligative effects.

Calculating the Freezing Point Depression

Using the formula:

$$
\Delta T_f = i \times K_f \times m
$$

Plugging in the known values:

$$
\Delta T_f = 2 \times 1.86\, \text{°C·kg/mol} \times 0.186\, \text{mol/kg} \approx 0.692\, \text{°C}
$$

Interpretation: The freezing point of water is lowered by approximately 0.692°C when 2.50 grams of NaCl are dissolved in 230 mL of water.

Determining the Final Freezing Point of the Solution

Pure water freezes at 0°C. With the depression:

$$
\text{Freezing point of solution} = 0\, \text{°C} - \Delta T_f \approx -0.692\, \text{°C}
$$

Conclusion:

The freezing point of the NaCl solution is approximately -0.69°C.

Factors Affecting the Accuracy and Practical Considerations

While the theoretical calculation provides a solid estimate, several factors can influence the actual freezing point:

- Incomplete Dissolution: If the NaCl does not fully dissolve, the number of particles in solution would be less, resulting in a higher freezing point than predicted.
- Impurities in Water: Presence of other solutes or impurities can alter colligative properties.
- Temperature Measurement Precision: Accurate thermometry is essential for precise freezing point determination.
- Temperature Equilibration: Ensuring the solution reaches thermal equilibrium before measurement.

Real-World Applications and Significance

Understanding freezing point depression of NaCl solutions has numerous practical applications:

- De-icing of roads: Salt lowers the freezing point of water on surfaces, preventing ice formation.
- Food preservation: Salt solutions inhibit microbial growth by altering freezing and boiling points.
- Chemical industry: Precise control of freezing points is crucial in

formulation and storage.
- Environmental science: Studying how saltwater bodies respond to temperature changes.

Extensions and Related Concepts

1. Effect of Concentration Variations

- Increasing the amount of NaCl will proportionally increase the freezing point depression until solubility limits are reached.
- For highly concentrated solutions, deviations from ideal behavior may occur.

2. Temperature-Dependent Dissociation

- At higher concentrations, ion pairing or activity coefficients may alter the effective van 't Hoff factor.

3. Other Solutes and Their Effects

- Non-electrolyte solutes like sugars also depress freezing point but to a lesser extent per mole because they do not dissociate into multiple particles.

4. Boiling Point Elevation

- Similar calculations apply, with the boiling point elevation formula, using the same parameters.

Summary and Final Remarks

In conclusion, adding 2.50 grams of sodium chloride to 230.0 mL of water results in a solution with a freezing point approximately -0.69°C below that of pure water. This calculation integrates fundamental principles of colligative properties, molality, dissociation factors, and thermodynamic constants, illustrating how a small amount of solute can significantly alter the physical properties of a solvent.

Understanding these principles is essential not only for academic purposes but also for practical applications across industries and environmental management. The precise calculation of freezing point depression allows scientists and engineers to predict and manipulate the behavior of solutions under various conditions, ensuring safety, efficiency, and innovation.

References & Further Reading:

- Atkins, P., & de Paula, J. (2010). Physical Chemistry. Oxford University Press.
- Zumdahl, S. S., & Zumdahl, S. A. (2014). Chemistry: An Atoms First Approach. Cengage Learning.
- CRC Handbook of Chemistry and Physics, 97th Edition.
- Educational resources on colligative properties from reputable chemistry textbooks and online platforms.

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