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Understanding how solutes affect the freezing point of solvents is a fundamental concept in chemistry, particularly in colligative properties. When a solute dissolves in a solvent like water, it alters the physical properties of the solution, including its freezing point. In this article, we will explore how to determine the freezing point of a solution containing 5.72 grams of magnesium chloride ($MgCl_2$) dissolved in 100 grams of water, considering the cryoscopic constant (K^z) for water, which is 1.86°C/m. This step-by-step guide will help you understand the underlying principles and calculations involved in freezing point depression.

Understanding Colligative Properties and Freezing Point Depression

What Are Colligative Properties?

Colligative properties are physical properties of solutions that depend on the number of dissolved particles, regardless of their identity. These include:

- Vapor pressure lowering
- Boiling point elevation
- Freezing point depression
- Osmotic pressure

The property of interest here is freezing point depression, which occurs because the addition of solute particles interferes with the formation of a solid crystalline structure in the solvent, effectively lowering the temperature at which the solvent freezes.

How Does Freezing Point Depression Work?

The relationship between the freezing point depression (ΔT_f) and the molality of the solution is given by the formula:

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

Where:
- i = van't Hoff factor (number of particles formed when the solute dissolves)
- K_f = cryoscopic constant (freezing point depression constant) for the solvent
- m = molality of the solution (moles of solute per kilogram of solvent)

Knowing these variables allows us to calculate the change in freezing point caused by the dissolved $MgCl_2$.

Step-by-Step Calculation of Freezing Point Depression for $MgCl_2$ Solution

1. Determine the molar mass of $MgCl_2$

Calculating the molar mass is the first step:

- Magnesium (Mg): 24.305 g/mol
- Chlorine (Cl): 35.453 g/mol

Thus:

$$
\text{Molar mass of MgCl}_2 = 24.305 + 2 \times 35.453 = 24.305 + 70.906 = 95.211\, \text{g/mol}
$$

2. Calculate the number of moles of $MgCl_2$ dissolved

Given:
- Mass of $MgCl_2$ = 5.72 g
- Molar mass = 95.211 g/mol

Number of moles:

$$
n = \frac{\text{mass}}{\text{molar mass}} = \frac{5.72}{95.211} \approx 0.06007\, \text{mol}

\]

3. Determine the molality of the solution

Molality (m) is defined as moles of solute per kilogram of solvent:
- Mass of water = 100 g = 0.100 kg

Therefore:

$$
m = \frac{0.06007\, \text{mol}}{0.100\, \text{kg}} = 0.6007\, \text{mol/kg}
$$

4. Find the van't Hoff factor (i) for $MgCl_2$

$MgCl_2$ dissociates in water into ions:

$$
\text{MgCl}_2 \rightarrow \text{Mg}^{2+} + 2\, \text{Cl}^-
$$

Thus, the van't Hoff factor:

$$
i = 3
$$

5. Calculate the freezing point depression (ΔT_f)

Using the formula:

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

Plugging in the values:

$$
\Delta T_f = 3 \times 1.86\, ^\circ\text{C/m} \times 0.6007\, \text{m}
\approx 3 \times 1.86 \times 0.6007
$$

$$
\Delta T_f \approx 3.36\, ^\circ\text{C}
$$

Determine the New Freezing Point of the Solution

1. Understand the freezing point of pure water

Pure water freezes at 0°C under standard atmospheric pressure.

2. Calculate the freezing point of the solution

Since the freezing point is lowered by ΔT_f:

$$
T_{f,\, \text{solution}} = T_{f,\, \text{water}} - \Delta T_f
$$

$$
T_{f,\, \text{solution}} = 0\,^\circ\text{C} - 3.36\,^\circ\text{C} = -3.36\,^\circ\text{C}
$$

Therefore, the freezing point of the solution is approximately -3.36°C.

Additional Considerations and Real-World Factors

Effect of Ion Dissociation and Non-Ideal Behavior

While calculations assume complete dissociation and ideal behavior, real solutions may deviate slightly due to ion pairing, interactions, and activity coefficients. However, for dilute solutions like this, the approximation is generally acceptable.

Implications for Practical Applications

Understanding freezing point depression is crucial in various fields:

- De-icing solutions for roads and aircraft
- Food preservation techniques
- Designing cryogenic systems

Summary of Key Points

- A solution of 5.72 g $MgCl_2$ in 100 g water has a molality of approximately 0.6007 mol/kg.

- $MgCl_2$ dissociates into three particles (i = 3), contributing to freezing point depression.
- The calculated freezing point depression is about 3.36°C.
- The resulting freezing point of the solution is approximately -3.36°C.

In conclusion, dissolving 5.72 grams of magnesium chloride in 100 grams of water lowers the freezing point of water to about -3.36°C. This calculation exemplifies the practical application of colligative properties and demonstrates how solute concentration influences phase transitions in solvents. Whether in industrial processes, scientific research, or everyday scenarios, understanding these principles helps in designing and controlling systems that depend on precise thermal behaviors.

Frequently Asked Questions

How do you calculate the freezing point depression for a solution of $MgCl_2$ in water?

The freezing point depression is calculated using $\Delta T_f = i \times K_f \times m$, where i is the van't Hoff factor, Kf is the freezing point depression constant, and m is the molality of the solution.

What is the molar mass of $MgCl_2$ used in calculating the molality?

The molar mass of $MgCl_2$ is approximately 95.21 g/mol.

What is the van't Hoff factor (i) for $MgCl_2$ in solution?

$MgCl_2$ dissociates into one Mg^{2+} ion and two Cl^- ions, so i = 3.

How do you determine the molality of the $MgCl_2$ solution in this problem?

Molality is calculated as m = mol of solute / mass of solvent in kg. First, convert 5.72 g $MgCl_2$ to mols, then divide by 0.1 kg (100 g of water).

What is the final freezing point of the solution

after calculating the depression?

The freezing point is lowered by $\Delta T_f = i \times K_f \times m$; subtract this value from 0°C to find the new freezing point.

Why is the water's cryoscopic constant (Kf) important in this calculation?

Kf represents the freezing point depression per molal concentration; it is essential for calculating how much the freezing point decreases when a solute is dissolved.

Additional Resources

Understanding the Freezing Point Depression of a Magnesium Chloride Solution in Water

When exploring the physical properties of solutions, one of the most intriguing phenomena is the change in freezing point caused by the presence of dissolved solutes. This phenomenon, known as freezing point depression, is a colligative property—meaning it depends on the number of dissolved particles in a solution rather than their identity. In this discussion, we will analyze the specific case of a solution formed by dissolving 5.72 grams of magnesium chloride ($MgCl_2$) in 100 grams of water, and determine its new freezing point. We will delve into the principles, calculations, and implications of this process with a comprehensive approach.

Fundamentals of Freezing Point Depression

What is Freezing Point Depression?

Freezing point depression refers to the lowering of a solvent's freezing point when a solute is dissolved in it. Pure water freezes at 0°C under standard atmospheric pressure. However, when a non-volatile solute such as $MgCl_2$ dissolves in water, the solution freezes at a temperature below 0°C. This phenomenon results from the disruption of the formation of an ordered solid lattice as solute particles interfere with water molecules' arrangement into ice.

Key Principles Underlying Freezing Point Depression

- Colligative Property: The extent of freezing point depression depends solely on the number of solute particles in the solution, not their chemical nature.
- Number of Particles (Ions or Molecules): The more particles in solution, the greater the depression.
- Van't Hoff Factor (i): Represents the number of particles into which a compound dissociates in solution.

Understanding the Components of the Problem

Given Data

- Mass of $MgCl_2$ dissolved, $m_{solute} = 5.72\, g$
- Mass of water, $m_{solvent} = 100\, g$
- Water's freezing point depression constant, $K_f = 1.86\, ^\circ C/m$

Objective

Calculate the new freezing point of the solution after dissolving 5.72 grams of $MgCl_2$ in 100 grams of water.

Step-by-Step Calculation Process

1. Convert Mass of Solute to Moles

- Molecular weight of $MgCl_2$:

$$
MW_{MgCl_2} = 24.305\, (Mg) + 2 \times 35.453\, (Cl) = 95.211\, g/mol
$$

- Moles of $MgCl_2$:

$$
n_{MgCl_2} = \frac{5.72\, g}{95.211\, g/mol} \approx 0.0601\, mol
$$

2. Determine the Number of Dissolved Particles (Ions)

Magnesium chloride dissociates in water as follows:

$$
MgCl_2 \rightarrow Mg^{2+} + 2Cl^-
$$

- Van't Hoff factor, i:

$$
i = 3
$$

(since 1 $MgCl_2$ produces 3 ions: 1 Mg^{2+} and 2 Cl^-)

- Total particles in solution:

$$
n_{particles} = n_{MgCl_2} \times i = 0.0601 \times 3 \approx 0.1803\, mol
$$

3. Calculate Molality of the Solution

- Molality (m) is defined as moles of solute per kilogram of solvent:

$$
m = \frac{n_{MgCl_2}}{\text{mass of solvent in kg}} = \frac{0.0601\, mol}{0.100\, kg} = 0.601\, mol/kg
$$

- Note: The molality of particles (number of particles per kg of solvent):

$$
m_{particles} = 3 \times 0.601\, mol/kg = 1.803\, mol/kg
$$

But for freezing point depression, it's the molality in terms of number of particles, which is 0.601 mol/kg, multiplied by the Van't Hoff factor.

4. Apply the Freezing Point Depression Formula

The general formula for freezing point depression is:

$$

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

where:

- ΔT_f = change in freezing point
- i = Van't Hoff factor
- K_f = freezing point depression constant
- m = molality of solute in mol/kg

Plugging in the values:

$$
\Delta T_f = 3 \times 1.86\, ^\circ C/m \times 0.601\, mol/kg
$$

$$
\Delta T_f \approx 3 \times 1.86 \times 0.601 \approx 3.36\, ^\circ C
$$

Final Calculation of the New Freezing Point

- The pure water freezing point: 0°C
- The depression: approximately 3.36°C

Thus:

$$
\boxed{
\text{Freezing point of the solution} = 0^\circ C - 3.36^\circ C \approx -3.36^\circ C
}
$$

Therefore, the solution of 5.72 grams $MgCl_2$ in 100 grams of water has a freezing point of approximately -3.36°C.

Deeper Insights and Considerations

Role of Dissociation and the Van't Hoff Factor

The Van't Hoff factor is crucial because it accounts for dissociation in solution:

- For $MgCl_2$, $i = 3$, as it dissociates into three ions.
- If the compound did not dissociate (or dissociation was negligible), i would be 1, and the depression would be less.

This dissociation significantly amplifies the colligative effect, as more particles are present in solution than just the initial molecules.

Effect of Temperature and Concentration

- The calculated freezing point depression assumes ideal behavior, meaning no ion pairing or interactions that would reduce the number of effective particles.
- At higher concentrations, deviations from ideality may occur, slightly altering the depression magnitude.
- The solution's concentration (approximately 0.6 mol/kg) is moderate, so the ideal assumption is reasonably valid here.

Implications for Practical Applications

Understanding freezing point depression has multiple real-world applications:

- Road de-icing: Salt solutions lower the freezing point of water, preventing ice formation.
- Cryopreservation: Controlling freezing points of solutions to preserve biological material.
- Chemical analysis: Determining molar masses and dissociation constants.

Additional Factors Influencing Freezing Point Depression

Ion Pairing and Non-Ideal Behavior

- At higher concentrations, ions may form pairs, reducing the effective number of particles.
- This phenomenon leads to deviations from the ideal colligative property

calculations.

Temperature-Dependent Dissociation

- The degree of dissociation can vary with temperature, affecting the Van't Hoff factor.
- For dilute solutions at room temperature, the value of $i = 3$ is a good approximation.

Solvent Purity and Experimental Conditions

- Impurities in water can influence the actual freezing point.
- Accurate measurements require high-purity water and controlled conditions.

Summary and Final Remarks

In conclusion, dissolving 5.72 grams of $MgCl_2$ in 100 grams of water results in a solution with a significantly lowered freezing point, approximately -3.36°C. This calculation underscores the importance of colligative properties and the dissociation behavior of ionic compounds like $MgCl_2$. Understanding these principles is vital in fields ranging from materials science to environmental engineering.

This detailed analysis demonstrates the power of fundamental chemistry principles to predict real-world behaviors of solutions. Whether for academic purposes, industrial applications, or environmental considerations, grasping the nuances of freezing point depression enables scientists and engineers to manipulate and anticipate solution properties effectively.

Key Takeaways:

- Colligative properties depend on the number of particles, not their identity.
- Dissociation increases the number of particles, amplifying effects like freezing point depression.
- The Van't Hoff factor is essential in accurate calculations.
- Practical applications leverage these principles for safety, preservation, and analysis.

By mastering these concepts, one can approach complex solution chemistry problems with confidence and precision.

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