Calculate The Freezing Point Of A Solution Made By Dissolving 3.50 G Of Potassium Chloride (M=74.55 G/mol)

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Understanding how solutes affect the freezing point of solvents is a fundamental concept in chemistry, particularly in the study of colligative properties. In this article, we will demonstrate a step-by-step process to calculate the freezing point depression when dissolving 3.50 grams of potassium chloride (KCl) in a solvent, typically water. By exploring the principles involved, including molality, dissociation, and colligative properties, you will gain a comprehensive understanding of how to approach such calculations with confidence.

Introduction to Freezing Point Depression

The phenomenon of freezing point depression occurs when a solute is dissolved in a solvent, resulting in a lower freezing point than that of the pure solvent. This is a colligative property, meaning it depends on the number of dissolved particles, not their identity. The key aspects include:

- Colligative properties: Freezing point depression, boiling point elevation, vapor pressure lowering, and osmotic pressure.
- Application: Used in antifreeze formulations, food preservation, and in scientific experiments.

To quantify freezing point depression, chemists use the formula:

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

Where:
- ΔT_f = freezing point depression (in degrees Celsius)
- i = van't Hoff factor (number of particles the solute dissociates into)
- K_f = cryoscopic constant of the solvent (for water, approximately 1.86 °C·kg/mol)
- m = molality of the solution (moles of solute per kilogram of solvent)

Understanding the Components of the Calculation

Before performing the calculation, let's clarify each component:

1. Moles of Potassium Chloride (KCl)

Given:
- Mass of KCl = 3.50 g
- Molar mass of KCl (M) = 74.55 g/mol

Calculate:

$$
\text{moles of KCl} = \frac{\text{mass}}{\text{molar mass}} = \frac{3.50 \text{ g}}{74.55 \text{ g/mol}}
$$

2. Dissociation of KCl in Water

Potassium chloride is an ionic compound that dissociates in water:

$$
\text{KCl} \rightarrow \text{K}^+ + \text{Cl}^-
$$

- Van't Hoff factor, i, for KCl = 2 (since it dissociates into two ions)

3. Molality of the Solution

Molality (m) is calculated as:

$$
m = \frac{\text{moles of solute}}{\text{kg of solvent}}
$$

Assuming the solvent is water, the mass of the solvent must be known or assumed. For this example, let's assume the solution is prepared with 1 kg of water.

Step-by-Step Calculation Process

Now, let's proceed with the calculations:

Step 1: Calculate Moles of Potassium Chloride

$$
\text{moles of KCl} = \frac{3.50}{74.55} \approx 0.0469 \text{ mol}
$$

Step 2: Determine the Van't Hoff Factor (i)

Since KCl dissociates into two ions:

$$
i = 2
$$

Step 3: Calculate Molality of the Solution

Assuming 1 kg of water:

$$
m = \frac{0.0469 \text{ mol}}{1 \text{ kg}} = 0.0469 \text{ mol/kg}
$$

Step 4: Calculate Freezing Point Depression (ΔT_f)

Using the formula:

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

Plugging in the values:

$$
\Delta T_f = 2 \times 1.86\,^\circ \mathrm{C\,kg/mol} \times 0.0469\, \mathrm{mol/kg}
$$

$$
\Delta T_f \approx 2 \times 1.86 \times 0.0469 = 2 \times 0.0874 = 0.1748\,^\circ \mathrm{C}
$$

Final Result: Freezing Point of the Solution

The freezing point of pure water is 0°C. Since the solution's freezing point is depressed by approximately 0.175°C, the new freezing point becomes:

$$
T_f = 0^\circ \mathrm{C} - 0.175^\circ \mathrm{C} = -0.175^\circ \mathrm{C}
$$

Therefore, the freezing point of the solution made by dissolving 3.50 g of potassium chloride in 1 kg of water is approximately -0.175°C.

Additional Considerations and Variations

While the above calculation provides a straightforward example, real-world scenarios may involve additional factors:

1. Varying Solvent Mass

If the amount of solvent differs from 1 kg, recalculate molality accordingly:

$$
m = \frac{\text{moles of solute}}{\text{kg of solvent}}
$$

For example, if 0.5 kg of water is used, molality doubles, leading to a larger freezing point depression.

2. Non-Ionic Solutes

For non-electrolyte solutes that do not dissociate, $i=1$. This significantly affects the calculation.

3. Temperature Correction

At higher concentrations, interactions between ions may cause deviations from ideal behavior. Advanced models or experimental data may be required for precise calculations.

Practical Applications of Freezing Point Depression Calculations

Understanding freezing point depression has numerous practical uses:

- Antifreeze formulations: Using KCl or other salts to lower freezing points in automotive radiators.
- Food preservation: Controlling ice formation in frozen foods.
- Scientific research: Determining solute concentrations and properties of solutions.
- Environmental science: Studying natural waters and their freezing behavior.

Summary of Key Steps in Calculating Freezing Point Depression

To accurately determine the freezing point depression of a solution:

1. Calculate the moles of solute: Use the mass and molar mass.
2. Determine the dissociation factor (i): Based on the solute's nature.
3. Compute molality: Moles of solute per kilogram of solvent.
4. Apply the colligative property formula: $\Delta T_f = i \times K_f \times m$.
5. Calculate the new freezing point: Subtract ΔT_f from the pure solvent's freezing point.

Conclusion

Calculating the freezing point depression caused by dissolving potassium chloride in water involves understanding the principles of colligative properties and dissociation behavior. In our example, dissolving 3.50 g of KCl in approximately 1 kg of water results in a depression of about 0.175°C, lowering the freezing point from 0°C to approximately -0.175°C. This calculation exemplifies how solute concentration and dissociation influence the physical properties of solutions, which is fundamental in various scientific and industrial applications.

By mastering these calculations, students and professionals can predict and manipulate solution behaviors effectively, enhancing their understanding of solution chemistry and its practical uses.

Keywords: Freezing point depression, colligative properties, potassium chloride, molality, molar mass, dissociation, K_f, solution calculation, colligative formulas, freezing point of solutions

Frequently Asked Questions

How do you calculate the freezing point depression of a solution made by dissolving potassium chloride?

You use the formula $\Delta T_f = i\ K_f$ molality, where i is the van 't Hoff factor, K_f is the cryoscopic constant of the solvent, and molality is moles of solute per kilogram of solvent.

What is the van 't Hoff factor (i) for potassium chloride in solution?

Potassium chloride dissociates into K^+ and Cl^- ions in solution, so the van 't Hoff factor i = 2.

Given 3.50 g of KCl, how many moles of solute are dissolved?

Number of moles = mass / molar mass = 3.50 g / 74.55 g/mol ≈ 0.0469 mol.

If the solution is prepared with 1.00 kg of solvent, what is the molality of the solution?

Molality = moles of solute / kg of solvent = 0.0469 mol / 1.00 kg = 0.0469 mol/kg.

How do you calculate the change in freezing point (ΔT_f) for this solution?

Use $\Delta T_f = i\ K_f$ molality. For water, $K_f \approx 1.86$ °C·kg/mol, so $\Delta T_f = 2\ 1.86\ 0.0469 \approx 0.174$ °C.

What is the actual freezing point of the solution after dissolving 3.50 g of KCl in 1 kg of water?

Since pure water freezes at 0 °C, the freezing point depression is approximately 0.174 °C, so the new freezing point is 0 °C - 0.174 °C ≈ -0.174 °C.

Additional Resources

Calculating the Freezing Point of a Potassium Chloride Solution: A Comprehensive Guide

Understanding the behavior of solutions when cooled is fundamental in both academic chemistry and practical applications such as antifreeze formulations, food preservation, and chemical manufacturing. Among the various properties affected by solutes, the freezing point depression is particularly significant. This article offers an in-depth exploration of how to calculate the freezing point of a solution made by dissolving 3.50 grams of potassium chloride (KCl) with a molar mass of 74.55 g/mol. We will walk through the principles, calculations, and implications step by step, equipping you with a clear understanding of this essential chemical concept.

Fundamental Concepts of Freezing Point Depression

Before diving into calculations, it’s crucial to grasp the foundational principles that govern freezing point depression.

What is Freezing Point Depression?

Freezing point depression is a colligative property – meaning it depends on the number of solute particles in a solvent, not their identity. When a solute dissolves in a solvent, it disrupts the formation of the solvent's crystalline structure during freezing, thereby lowering its freezing point.

The general formula for calculating the change in freezing point (ΔT_f) is:

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

Where:
- i = van 't Hoff factor (number of particles into which a compound dissociates)
- K_f = cryoscopic constant (freezing point depression constant) for the solvent
- m = molality of the solution in mol/kg

The new freezing point ($T_{f,solution}$) is then:

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

Step 1: Understanding the Components of the Calculation

Let's analyze each component involved in this calculation, particularly focusing on potassium chloride in water.

1. Molar Mass of Potassium Chloride (KCl)

Given:
- $M_{KCl} = 74.55 \, \text{g/mol}$

The molar mass is essential to convert grams of KCl into moles, which in turn determine the number of particles in solution.

2. Amount of KCl Dissolved

Given:
- Mass of KCl = 3.50 g

Converting grams to moles:
$$
n_{KCl} = \frac{\text{mass}}{\text{molar mass}} = \frac{3.50\, \text{g}}{74.55\, \text{g/mol}} \approx 0.0469\, \text{mol}
$$

3. Van 't Hoff Factor (i) for KCl

Potassium chloride dissociates fully in water:
$$
\mathrm{KCl} \rightarrow \mathrm{K}^{+} + \mathrm{Cl}^{-}
$$

Thus, each formula unit yields 2 particles. Therefore:
$$
i = 2
$$

4. Determining the Solvent and Its Properties

Assuming the solvent is water, with a known freezing point:
- $T_{f,water} = 0^\circ C = 273.15\, K$

The cryoscopic constant for water:
- $K_f = 1.86\, ^\circ C\, \text{kg/mol}$

Step 2: Calculating Molality of the Solution

Molality (m) is defined as moles of solute per kilogram of solvent. To determine molality, we need to know the mass of the solvent in kilograms.

Key Point: Since the mass of the solvent isn't specified, a typical assumption is that the solution is prepared by dissolving the KCl into a certain amount of water, or we can select a standard solvent mass for illustrative purposes.

Example: Let's assume the solution is prepared with 100 g of water (which is equivalent to 0.100 kg). This is a common basis for such calculations.

Calculations:
$$
m = \frac{n_{KCl}}{\text{mass of solvent in kg}} = \frac{0.0469\, \text{mol}}{0.100\, \text{kg}} = 0.469\, \text{mol/kg}
$$

Step 3: Calculating the Freezing Point Depression (ΔT_f)

Applying the formula:

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

Substituting the known values:

$$
\Delta T_f = 2 \times 1.86\, ^\circ C\, \text{kg/mol} \times 0.469\, \text{mol/kg} \approx 1.75^\circ C
$$

\]

This indicates that the freezing point of the solution is lowered by approximately 1.75°C from pure water.

Step 4: Final Freezing Point of the Solution

Calculating the new freezing point:

\[
T_{f,solution} = T_{f,water} - \Delta T_f = 0^\circ C - 1.75^\circ C = -1.75^\circ C
\]

Interpretation: The solution will freeze at approximately \(-1.75^\circ C\), significantly below the freezing point of pure water, demonstrating the effective depressant effect of dissolved potassium chloride.

Additional Considerations and Practical Implications

While the above calculation provides a theoretical freezing point, several factors can influence real-world results:

1. Concentration Dependence

- At higher concentrations, ion interactions may cause deviations from ideal behavior, leading to slightly different freezing points.
- For very concentrated solutions, activity coefficients become significant.

2. Dissolution Completeness

- The calculation assumes complete dissociation of KCl.
- In practice, dissociation may be incomplete at high concentrations.

3. Temperature Range and Solvent Purity

- Impurities or the presence of other solutes can alter the freezing point.
- The temperature at which measurements are taken can also affect observed values.

4. Application in Industry

- Such calculations are critical in designing antifreeze solutions, where controlled freezing point depression is necessary.
- Precise knowledge of solute effects ensures safety and effectiveness in applications like car radiators and de-icing.

Summary of the Calculation Process

To encapsulate, here is a step-by-step summary:

1. Convert grams of KCl to moles:
$$
n_{KCl} = \frac{\text{mass}}{M} = \frac{3.50\, \text{g}}{74.55\, \text{g/mol}} \approx 0.0469\, \text{mol}
$$
2. Assume or determine the amount of solvent (e.g., 100 g water = 0.100 kg).
3. Calculate molality:
$$
m = \frac{0.0469\, \text{mol}}{0.100\, \text{kg}} = 0.469\, \text{mol/kg}
$$
4. Determine ΔT_f:
$$
\Delta T_f = i \times K_f \times m = 2 \times 1.86 \times 0.469 \approx 1.75^\circ C
$$
5. Calculate the new freezing point:
$$
T_{f,solution} = 0^\circ C - 1.75^\circ C = -1.75^\circ C
$$

Conclusion

Calculating the freezing point of a potassium chloride solution involves understanding the interplay between molar quantities, dissociation behavior, and colligative properties. By meticulously converting mass to moles, considering dissociation effects via the van 't Hoff factor, and applying the freezing point depression formula, one can accurately predict how a solution's freezing point shifts relative to pure solvent.

This process underscores the importance of precise measurements and assumptions in chemical calculations, especially when applied to real-world scenarios like antifreeze solutions, where safety and performance hinge on accurate property predictions. Whether you are a student, researcher, or industry professional, mastering these calculations enhances your ability to design, analyze, and optimize solutions for a broad range of applications.

Disclaimer: The actual freezing point may vary depending on solution concentration, impurities, and experimental conditions. Always consider these factors in practical scenarios.

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