The Freezing Point Of An Aqueous Solution That Contains A Non Elextrolyte Is -4.0 C What Is The Molal

The Freezing Point Of An Aqueous Solution That Contains A Non Electrolyte Is -4.0°C What Is The Molal

Understanding the relationship between the freezing point depression of a solution and its molality is fundamental in chemistry. When dealing with aqueous solutions, especially those containing non-electrolytes, it becomes essential to determine the molality based on the observed change in freezing point. This article provides a comprehensive guide to calculating the molal concentration of a non-electrolyte solution with a freezing point of -4.0°C, exploring key concepts, formulas, and step-by-step procedures to facilitate a clear understanding.

Fundamentals of Freezing Point Depression

What is Freezing Point Depression?

Freezing point depression is a colligative property, meaning it depends solely on the number of solute particles in a solution, not their identity. When a solute is dissolved in a solvent like water, the freezing point of the solution decreases relative to the pure solvent. This phenomenon occurs because the solute particles interfere with the formation of the solid phase, requiring a lower temperature to achieve freezing.

The magnitude of freezing point depression (ΔT_f) is directly proportional to the molality (m) of the solution, governed by the formula:

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

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

Understanding the Components of the Formula

The Van't Hoff Factor (i)

Since the solution contains a non-electrolyte, it does not dissociate into multiple ions when dissolved. Therefore:
- $i = 1$

This simplifies the calculations because the number of particles produced per solute molecule remains one.

Crucoscopic Constant (K_f) for Water

For water, which is the typical solvent in aqueous solutions, the cryoscopic constant is:
- $K_f = 1.86\,°C \cdot kg/mol$

This value is well-established and can be used directly in calculations involving water as the solvent.

Calculating the Molality of the Solution

Given:
- The freezing point of the solution, $T_f = 0°C - 4.0°C = -4.0°C$
- K_f for water = 1.86°C·kg/mol
- For a non-electrolyte, $i = 1$

The change in freezing point (ΔT_f) is:

$$
\Delta T_f = |T_{f,\, pure\, solvent} - T_{f,\, solution}| = |0°C - (-4.0°C)| = 4.0°C
$$

Using the formula:

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

Substituting the known values:

$$
4.0°C = 1 \times 1.86°C \cdot kg/mol \times m
$$

Rearranged to solve for molality (m):

$$
m = \frac{\Delta T_f}{i \times K_f} = \frac{4.0°C}{1 \times 1.86°C \cdot kg/mol}
$$

Calculating:

$$
m \approx \frac{4.0}{1.86} \approx 2.15\, mol/kg
$$

Therefore, the molality of the solution is approximately 2.15 mol/kg.

Implications and Practical Applications

Why is Knowing the Molality Important?

Determining molality is crucial in various fields, including:

- Pharmacology: Precise concentration calculations for drug formulations.
- Food Chemistry: Understanding concentration effects on freezing points in food preservation.
- Environmental Science: Studying natural water bodies and pollutant distribution.
- Industrial Processes: Designing solutions with specific freezing or boiling points.

How to Use This Information?

Knowing the molality allows chemists and engineers to:

- Predict the freezing point of new solutions.
- Control and manipulate solution properties for desired outcomes.
- Verify experimental data against theoretical predictions.

Additional Considerations in Freezing Point Calculations

Effect of Electrolytes vs. Non-Electrolytes

- Electrolytes: Dissociate into multiple ions, increasing the van't Hoff factor ($i > 1$), leading to a greater depression of freezing point.
- Non-electrolytes: Do not dissociate ($i=1$), so their effect on freezing point depression is directly proportional to molality.

Temperature Measurement Accuracy

- Accurate determination of the freezing point depression relies on precise temperature measurements.
- Small errors can significantly impact the calculated molality.

Limitations of the Model

- Assumes ideal behavior and no interactions between solute particles.
- Real solutions may deviate slightly due to intermolecular forces.

Summary and Key Takeaways

- The freezing point depression of an aqueous solution containing a non-electrolyte with a depression of -4.0°C corresponds to a molality of approximately 2.15 mol/kg.
- The calculation utilizes the colligative property formula $\Delta T_f = i \times K_f \times m$, with $i=1$ for non-electrolytes.
- Understanding the parameters and their implications allows for accurate solution characterization and process optimization.

FAQs About Freezing Point Depression and Molality

1. What is the significance of the van't Hoff factor?
It accounts for the number of particles a solute produces in solution. For non-electrolytes, it's 1; for electrolytes, it's higher.

2. Can this method be used for solutions other than water?
Yes, but you need the appropriate K_f value for the solvent in question.

3. How does molality differ from molarity?
Molality is moles of solute per kilogram of solvent, whereas molarity is moles of solute per liter of solution.

4. What are common non-electrolytes?
Sugars (sucrose, glucose), alcohols (ethanol), and urea are typical non-electrolytes.

5. What happens if the solution contains electrolytes?
The van't Hoff factor increases, leading to a larger freezing point depression for the same molality.

In conclusion, understanding how to determine the molality of an aqueous solution based on freezing point depression is a vital skill in chemistry. By applying the colligative property formula and considering the nature of the solute, chemists can accurately analyze and control solution properties for scientific and industrial purposes.

Frequently Asked Questions

What is the freezing point depression formula used to determine molality in this case?

The freezing point depression formula is $\Delta T_f = i K_f m$, where ΔT_f is the change in freezing point, i is the van 't Hoff factor, K_f is the cryoscopic constant, and m is the molality. For a non-electrolyte, i = 1.

Given the freezing point of -4.0°C for an aqueous solution, what is the freezing point of pure water?

Pure water freezes at 0.0°C; the solution's freezing point depression of 4.0°C indicates a decrease from 0.0°C to -4.0°C.

How do you calculate the molal concentration of a non-electrolyte solution with a freezing point depression of -4.0°C?

Using $\Delta T_f = i K_f m$ and knowing $i=1$ for a non-electrolyte, rearranged as $m = \Delta T_f / K_f$. For water, $K_f = 1.86°C·kg/mol$, so $m = 4.0 / 1.86 \approx 2.15$ mol/kg.

What assumptions are made about the solution when calculating molality from freezing point depression?

It assumes the solution behaves ideally, the solute is a non-electrolyte ($i=1$), and that the freezing point depression is solely due to the solute's presence, without other effects.

What is the significance of the non-electrolyte in determining the freezing point depression?

Since it is a non-electrolyte, it does not dissociate into ions, so the van 't Hoff factor $i=1$, simplifying calculations of molality from freezing point depression.

How does the freezing point depression change if the substance were an electrolyte that dissociates into multiple ions?

The van 't Hoff factor i would be greater than 1, increasing the freezing point depression for the same molality, resulting in a higher ΔT_f.

What is the molality of the solution if the solution contains a non-electrolyte causing a -4.0°C freezing point depression?

The molality is approximately 2.15 mol/kg, calculated using $m = \Delta T_f / K_f = 4.0 / 1.86 \approx 2.15$ mol/kg.

Why is understanding the freezing point depression important in chemistry?

It helps in determining molar concentrations, analyzing solution properties, and understanding colligative properties critical for various applications like antifreeze formulations and solution characterization.

Additional Resources

The Freezing Point Of An Aqueous Solution That Contains A Non-Electrolyte Is -4.0°C. What Is The Molality?

Understanding how solutes affect the freezing point of a solvent is fundamental in chemistry, especially in solutions chemistry. In this guide, we will explore the principles behind freezing point depression, analyze the given data—namely, an aqueous solution with a freezing point of -4.0°C containing a non-electrolyte—and walk through the step-by-step calculation to determine the molality of the solution. Whether you're a student preparing for an exam or a professional brushing up on colligative properties, this comprehensive approach will clarify the concepts and demonstrate practical problem-solving techniques.

Introduction to Colligative Properties and Freezing Point Depression

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

Freezing point depression is particularly useful because it provides insight into the concentration of solute particles in a solution. When a solute is dissolved in a solvent—here, water—the freezing point of the solution decreases compared to pure water. The extent of this decrease is proportional to the molal concentration of the solute particles.

The general formula for freezing point depression is:

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

Where:
- ΔT_f = change in freezing point temperature (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)

For water, the cryoscopic constant K_f is approximately 1.86°C·kg/mol.

Step 1: Recognizing the Nature of the Solute

The problem states the solution contains a non-electrolyte. This detail is critical because non-electrolytes do not dissociate into multiple particles in solution. Consequently, the van 't Hoff factor i for a non-electrolyte is 1.

Step 2: Analyzing the Data

Given:
- Freezing point of the solution, $T_{solution} = -4.0°C$
- Freezing point of pure water, $T_{pure\,water} = 0°C$

Thus, the freezing point depression:

$$\Delta T_f = T_{pure\,water} - T_{solution} = 0°C - (-4.0°C) = 4.0°C$$

Step 3: Applying the Freezing Point Depression Formula

Since the solute is a non-electrolyte, the van 't Hoff factor:

$$i = 1$$

Using the known K_f for water:

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

Substituting the known values:

$$4.0°C = 1 \times 1.86°C \cdot \mathrm{kg/mol} \times m$$

Solving for molality m:

$$m = \frac{4.0°C}{1.86°C \cdot \mathrm{kg/mol}}$$

$$m \approx 2.15\, \mathrm{mol/kg}$$

This is the molality of the solution, indicating approximately 2.15 moles of non-electrolyte solute per kilogram of water.

Understanding the Significance of the Result

The calculated molality reflects the concentration of the non-electrolyte in the solution. A molality of roughly 2.15 mol/kg suggests a relatively concentrated solution, given that typical colligative property changes are quite sensitive to solute quantity.

Additional Considerations and Common Pitfalls

1. Assumption of Ideal Behavior

The calculation presumes ideal solution behavior, meaning the solute does not significantly interact with the solvent beyond dissolving. In real-world situations, deviations might occur, but for most dilute solutions, this approximation holds well.

2. Correct Identification of the Van 't Hoff Factor

Since the problem specifies a non-electrolyte, the assumption $i=1$ is valid. For electrolytes, i could be greater than 1 depending on dissociation, which would affect the molality calculation.

3. Precision of Constants

Using the precise K_f value for water (1.86°C·kg/mol) is important for accuracy. Variations in experimental conditions or impurities can slightly influence the freezing point depression.

Practical Applications of Freezing Point Depression Calculations

Understanding the relationship between solute concentration and freezing point depression has numerous practical applications:

- Antifreeze formulation: To prevent freezing in engines, antifreeze solutions are designed with specific concentrations to lower the freezing point sufficiently.
- Determining molar mass: By measuring freezing point depression and knowing the molality, chemists can back-calculate the molar mass of unknown solutes.
- Quality control: In manufacturing, colligative properties are used to verify solution concentrations.

Summary of the Calculation Steps

1. Identify the change in freezing point: $\Delta T_f = 4.0°C$
2. Use the colligative property formula: $\Delta T_f = i \times K_f \times m$
3. Plug in known values: $i=1$, $K_f=1.86°C \cdot \mathrm{kg/mol}$
4. Solve for molality:

$$ m = \frac{\Delta T_f}{i \times K_f} = \frac{4.0}{1 \times 1.86} \approx 2.15\, \mathrm{mol/kg} $$

Final Remarks

The molality of the aqueous solution containing a non-electrolyte with a freezing point of -4.0°C is approximately 2.15 mol/kg. This calculation underscores how colligative properties serve as a powerful tool to quantify solute concentration based solely on measurable physical changes. Mastery of these concepts enables chemists to analyze solutions, design industrial processes, and understand the fundamental behaviors of mixtures at the molecular level.

References

- Atkins, P., & de Paula, J. (2010). Physical Chemistry. Oxford University Press.
- Zumdahl, S. S., & Zumdahl, S. A. (2017). Chemistry: An Atoms First Approach. Cengage Learning.
- Lide, D. R. (Ed.). (2004). CRC Handbook of Chemistry and Physics (85th ed.). CRC Press.

The Freezing Point Of An Aqueous Solution That Contains A Non Elextrolyte Is 4 0 C What Is The Molal

The Freezing Point Of An Aqueous Solution That Contains A Non Elextrolyte Is 4 0 C What Is The Molal

Related Articles

- Suppose Michael Is Willing To Drive Across Town To Save 40% On A Sweatshirt With A List Price Of $80.
- The Marginal Revenue Curve Of A Firm With Market Power Will Always Lie Below Its Demand Curve Because
- The Europeans Began To Colonize Areas Rich In Resources Such As Rubber And Petroleum After (1.)_____.

Back to Home