6) When 5.00 Grams Of Ammonium Chloride, NH4Cl, Is Added To 100. ML Of Water The Temperature Drops By4.2C,

6) When 5.00 Grams Of Ammonium Chloride, NH4Cl, Is Added To 100. ML Of Water The Temperature Drops By 4.2°C

The scenario where 5.00 grams of ammonium chloride (NH_4Cl) is introduced into 100 mL of water resulting in a temperature decrease of 4.2°C provides a fascinating insight into the thermodynamics of dissolving processes. This phenomenon is characteristic of an endothermic dissolution, where energy is absorbed from the surrounding water, leading to a temperature drop. Understanding this process involves exploring concepts such as enthalpy of solution, solubility, and the energy exchanges involved in dissolving salts in water. This article delves into the scientific principles underlying this observation, the calculations involved in determining the enthalpy change, and the broader implications for thermochemistry and solution chemistry.

Understanding the Dissolution of Ammonium Chloride

What Is Ammonium Chloride?

- Ammonium chloride (NH_4Cl) is an inorganic salt composed of ammonium (NH_4^+) and chloride (Cl^-) ions.
- It is commonly used in medicine, as a food additive, and in industrial applications.
- It is highly soluble in water, which makes it useful for various chemical processes.

Solubility and Dissolution Process

The dissolution of NH_4Cl in water involves breaking ionic bonds in the salt and forming interactions with water molecules. This process can be represented as:

$$\text{NH}_4\text{Cl (s)} \rightarrow \text{NH}_4^+ (aq) + \text{Cl}^- (aq)$$

The process involves several energetic considerations, including lattice energy, hydration energy, and entropy changes, which collectively determine whether the dissolution is endothermic or exothermic.

Thermodynamics of Dissolving NH_4Cl in Water

Enthalpy Change of Solution (ΔH_{sol})

The enthalpy change associated with dissolving NH_4Cl is a critical parameter. It indicates whether energy is absorbed or released during dissolution.

- If ΔH_{sol} is positive, the process is endothermic, and the temperature drops upon dissolution.
- If ΔH_{sol} is negative, the process is exothermic, and the temperature rises.

Significance of Observed Temperature Drop

The observed temperature decrease of 4.2°C upon dissolving 5.00 g of NH_4Cl suggests an endothermic process. This aligns with known thermodynamic data that NH_4Cl dissolves endothermically in water.

Calculating the Enthalpy Change (ΔH_{sol})

Step 1: Determine the Moles of NH_4Cl Dissolved

Given data:

- Mass of NH_4Cl = 5.00 g
- Molar mass of NH_4Cl: N = 14.01 g/mol, H = 1.008 g/mol, Cl = 35.45 g/mol

Calculating molar mass:
\[

$$
\text{Molar mass} = 14.01 + (4 \times 1.008) + 35.45 = 53.49\, \text{g/mol}
$$

Number of moles:

$$
n = \frac{5.00\, \text{g}}{53.49\, \text{g/mol}} \approx 0.0935\, \text{mol}
$$

Step 2: Convert Temperature Change to Heat Absorbed

The heat (q) absorbed or released is related to the temperature change via the specific heat capacity (c) of water and its mass (m).

- Specific heat capacity of water, $c = 4.18\, \text{J/g}^\circ \text{C}$
- Mass of water, $m = 100\, \text{mL} \approx 100\, \text{g}$

Calculating the heat absorbed:

$$
q = m \times c \times \Delta T = 100\, \text{g} \times 4.18\, \text{J/g}^\circ \text{C} \times 4.2^\circ \text{C} \approx 1756\, \text{J}
$$

Since the temperature drops, the water absorbs heat, which is equal in magnitude to the heat absorbed during dissolution.

Step 3: Calculate ΔH_{sol} per mole

Enthalpy change per mole of NH_4Cl is:

$$
\Delta H_{sol} = \frac{q}{n} = \frac{1756\, \text{J}}{0.0935\, \text{mol}} \approx 18,778\, \text{J/mol} \approx 18.78\, \text{kJ/mol}
$$

This positive value confirms that dissolving NH_4Cl is an endothermic process with an enthalpy of solution of approximately +18.78 kJ/mol.

Implications of the Enthalpy of Solution

Understanding Endothermic Dissolution

- The absorption of heat during dissolution explains the observed cooling of water.
- This process is driven by the energy required to overcome lattice energy and hydrate the ions.

Practical Applications

1. Cooling packs: NH_4Cl solutions are used in instant cold packs, where the dissolution absorbs heat, providing a cooling effect.
2. Thermochemical studies: Understanding enthalpy changes helps in designing processes involving heat absorption or release.

Factors Affecting the Dissolution and Temperature Change

Concentration and Quantity of Salt

- Adding more NH_4Cl increases the heat absorbed, possibly leading to a more significant temperature drop.
- The maximum solubility at a given temperature limits how much salt can be dissolved.

Temperature of Water Before Dissolution

- Initial temperature influences the extent of temperature change observed.
- Pre-cooling or heating water can modify thermodynamic measurements.

Nature of the Solvent

- The polarity and specific heat capacity of water make it an excellent solvent for ionic salts like NH_4Cl.
- Different solvents would lead to different enthalpy changes and temperature effects.

Broader Context and Related Concepts

Comparison with Other Salts

- Many salts, such as sodium hydroxide ($NaOH$), dissolve exothermically, releasing heat and increasing temperature.
- Understanding the thermodynamics of these processes helps in applications ranging from industrial manufacturing to biological systems.

Thermochemical Equations and Data

Standard enthalpy of solution values are tabulated for many salts, providing reference data for calculations and predictions.

Limitations and Assumptions in Calculations

- Assumed complete dissolution and no heat loss to surroundings.
- Neglects minor heat exchanges and impurities.
- Assumes uniform temperature throughout the solution.

Conclusion

The observation that dissolving 5.00 g of ammonium chloride in 100 mL of water causes a temperature drop of 4.2°C exemplifies an endothermic process characterized by an enthalpy of solution of approximately +18.78 kJ/mol. This phenomenon illustrates fundamental principles of thermochemistry, including energy absorption during ionic dissolution, the balancing of lattice and hydration energies, and the practical uses of such processes in cooling applications. Understanding these principles enhances our grasp of solution chemistry and the energetic considerations involved in dissolving ionic compounds, contributing to various scientific and industrial fields.

Frequently Asked Questions

What is the significance of the temperature drop when ammonium chloride is dissolved in water?

The temperature drop indicates an endothermic process where heat is absorbed from the surroundings, cooling the solution as ammonium chloride dissolves.

How much heat is absorbed during the dissolution of 5.00 grams of ammonium chloride?

The heat absorbed can be calculated using the specific heat capacity of water, the mass of water, and the temperature change, indicating the enthalpy of solution per gram of NH_4Cl.

What is the molar mass of ammonium chloride (NH_4Cl)?

The molar mass of NH_4Cl is approximately 53.5 grams per mole.

How many moles of ammonium chloride were dissolved in the water?

About 0.0935 moles of NH_4Cl were dissolved, calculated as 5.00 g divided by 53.5 g/mol.

What is the enthalpy change (ΔH) per mole for the dissolution of ammonium chloride based on this data?

Using the temperature change and the heat capacity of water, ΔH per mole can be estimated, indicating an endothermic process with a specific enthalpy value.

Why does ammonium chloride dissolve endothermically in water?

The dissolution process requires energy to break ionic bonds in NH4Cl, which is absorbed from the surroundings, resulting in a temperature decrease.

How does the amount of water affect the temperature change during dissolution?

Larger volumes of water can absorb more heat without a significant temperature change, whereas smaller amounts result in a more noticeable temperature drop.

Can the temperature change be used to determine the enthalpy of solution experimentally?

Yes, by measuring the temperature change, mass of solute, and water, one can calculate the enthalpy of solution per mole of solute dissolved.

Additional Resources

When 5.00 grams of ammonium chloride (NH_4Cl) is added to 100 mL of water, resulting in a temperature drop of 4.2°C, this phenomenon exemplifies the intricate interplay between physical chemistry principles such as dissolution, thermodynamics, and calorimetry. This scenario not only highlights the energy changes during solvation but also provides insights into the properties of salts and their effects on aqueous solutions. Analyzing this process reveals critical concepts relevant to both theoretical understanding and practical applications, ranging from industrial processes to laboratory experiments.

Introduction to the Phenomenon

Adding a solid solute like ammonium chloride into water often results in temperature changes, which are directly linked to the energy exchanges involved in dissolving the compound. In this specific case, 5.00 grams of NH_4Cl causes a temperature drop of 4.2°C in 100 mL of water, indicating an endothermic process where heat is absorbed from the surroundings to facilitate dissolution. Understanding the underlying mechanisms requires a detailed exploration of the dissolution process, thermodynamic principles, and the implications of the observed temperature change.

Understanding the Dissolution Process of NH_4Cl in Water

Nature of Ammonium Chloride

Ammonium chloride (NH_4Cl) is an ionic salt composed of ammonium (NH_4^+) and chloride (Cl^-) ions. It is highly soluble in water, making it a common laboratory reagent and industrial chemical. Its solubility increases with temperature, but the dissolution process itself can be either endothermic or exothermic depending on the interactions involved.

Mechanism of Dissolution

When NH_4Cl dissolves in water, the following steps occur:

1. Breaking of Ionic Lattice: The ionic lattice of solid NH_4Cl must be broken, which requires energy called lattice energy.

2. Hydration of Ions: The free ammonium and chloride ions are surrounded by water molecules, a process known as hydration, which releases energy called hydration energy.

3. Net Energy Change: The overall enthalpy change (ΔH) of the dissolution process depends on the balance between lattice energy and hydration energy.

Mathematically, this can be expressed as:

$\Delta H\_dissolution$ = Hydration Energy - Lattice Energy

If hydration energy exceeds lattice energy, the process is exothermic; if not, it is endothermic.

Relevance to the Observed Temperature Drop

The observed temperature decrease of 4.2°C indicates that the dissolution of NH_4Cl in water is endothermic, meaning that the energy required to overcome the lattice energy is greater than the energy released during hydration. Consequently, heat is absorbed from the surrounding water and container, leading to cooling.

Quantitative Analysis of the Temperature Change

Calculating the Heat Absorbed (q)

The heat absorbed or released during the process can be calculated using the formula:

$q = mc\Delta T$

where:
- m = mass of water in grams
- c = specific heat capacity of water (approximately 4.18 J/g°C)
- ΔT = temperature change (negative for cooling)

Given data:
- Volume of water = 100 mL
- Density of water ≈ 1 g/mL → mass (m) = 100 g
- ΔT = -4.2°C
- c = 4.18 J/g°C

Calculating the heat:
$q = 100 \text{ g} \times 4.18 \text{ J/g°C} \times (-4.2\text{°C}) = -1757.6 \text{ J}$

The negative sign indicates heat absorption from the surroundings.

Implication of the Heat Calculation

This amount of heat (approximately 1758 Joules) is required to facilitate the dissolution of 5.00 grams of NH_4Cl in 100 mL of water, confirming the endothermic nature of the process.

Thermodynamic Considerations

Enthalpy of Solution (ΔH_{sol})

The enthalpy change per mole of solute provides a more standardized measure of the process. To compute this, we need to determine the number of moles of NH_4Cl dissolved:

Molar mass of NH_4Cl:
- N: 14.01 g/mol
- H_4: 4 × 1.008 g/mol = 4.032 g/mol
- Cl: 35.45 g/mol

Total molar mass:
14.01 + 4.032 + 35.45 ≈ 53.49 g/mol

Number of moles:
$n = 5.00 \text{ g} / 53.49 \text{ g/mol} \approx 0.0934 \text{ mol}$

Now, the enthalpy change per mole:
$\Delta H_{sol} = q / n \approx 1758 \text{ J} / 0.0934 \text{ mol} \approx 18{,}837 \text{ J/mol}$

Since the process is endothermic, ΔH_{sol} is positive, indicating energy absorption.

Significance of the Enthalpy of Solution

An enthalpy of roughly +18.8 kJ/mol suggests moderate endothermicity, consistent with the observed temperature drop. This value aligns with known thermodynamic data for NH_4Cl, which typically exhibits an enthalpy of solution around +14 to +20 kJ/mol.

Practical and Industrial Implications

Applications in Cooling Systems

The endothermic dissolution of NH_4Cl forms the basis for various cooling applications, such as instant cold packs used in first aid. When the salt dissolves, it absorbs heat, effectively lowering the temperature of the affected area.

Relevance in Thermochemical Studies

Studying such dissolution processes helps chemists understand heat transfer mechanisms, thermodynamic properties of solutions, and energy management in chemical processes.

Industrial Processes and Reactions

In industrial settings, controlling temperature during dissolution is crucial for processes such as crystallization, purification, and chemical synthesis. Knowledge of the enthalpy changes allows engineers to optimize conditions for efficiency and safety.

Environmental and Safety Considerations

While NH_4Cl is generally safe for laboratory use, its endothermic dissolution also implies that handling large quantities requires attention to thermal effects. Managing heat flow and ensuring proper ventilation are essential for safe operation, especially when scaling up.

Conclusion

The scenario of dissolving 5.00 grams of ammonium chloride in 100 mL of water, resulting in a temperature decrease of 4.2°C, offers a compelling illustration of thermodynamic principles in action. The process exemplifies an endothermic dissolution, characterized by the absorption of approximately 1.76 kJ of heat. Such phenomena are not only academically intriguing but also practically significant, underpinning technologies ranging from medical cold packs to industrial chemical processes. Understanding the energy exchanges involved in salt dissolution enhances our capacity to manipulate and harness these reactions for technological and scientific advancements.

In essence, this case underscores the importance of thermochemistry in everyday applications and the necessity of detailed analysis to fully comprehend the energy dynamics involved in seemingly simple chemical processes.

6 When 5 00 Grams Of Ammonium Chloride Nh4cl Is Added To 100 Ml Of Water The Temperature Drops By4 2c

6 When 5 00 Grams Of Ammonium Chloride Nh4cl Is Added To 100 Ml Of Water The Temperature Drops By4 2c

Related Articles

- Which Of The Following Statements Is True Of Supply Chain Management In Todays Marketplace?a.) Businesses
- The ______is A Financial Statement Detailing A Firm's Assets, Liabilities, And Owners' Equity. A. Balance
- There Are 8 Ball M, N, O, P, Q, R, S And T. 7 Of Them Are Identical, The 8th I Either Heavier Or Lighter.

Back to Home