Calculate The PH Of A Solution Prepared By Dissolving 1.90 G Of Sodium Acetate, Ch3cooNa, In 71.0 ML

Calculate The PH Of A Solution Prepared By Dissolving 1.90 G Of Sodium Acetate, Ch3cooNa, In 71.0 ML is a common problem in chemistry that involves understanding the principles of acid-base chemistry, solution chemistry, and the concept of pH. Sodium acetate (Ch3cooNa) is a salt derived from a weak acid, acetic acid (CH3COOH), and a strong base, sodium hydroxide (NaOH). When dissolved in water, sodium acetate hydrolyzes, producing acetate ions (Ch3coo−) that can interact with water to determine the pH of the solution. Calculating the pH involves several steps, including determining the concentration of acetate ions, understanding their hydrolysis behavior, and applying equilibrium principles. This article provides a comprehensive guide to calculating the pH of such a solution, covering essential concepts, detailed calculations, and practical considerations.

Understanding Sodium Acetate and Its Behavior in Solution

What is Sodium Acetate?

Sodium acetate (Ch3cooNa) is an ionic compound formed by the combination of the acetate ion (Ch3coo−) and sodium ion (Na+). It is commonly used in various industrial and laboratory applications, including as a buffer solution, in food preservation, and in chemical synthesis.

Properties of Sodium Acetate in Water

When sodium acetate dissolves in water, it dissociates completely:
- NaCh3coo(s) → Na+(aq) + Ch3coo−(aq)

The acetate ion (Ch3coo−) is the conjugate base of acetic acid, a weak acid. This conjugate base can accept protons (H+) from water, leading to hydrolysis and influencing the pH of the solution.

Calculating Moles of Sodium Acetate in the Solution

Step 1: Find the molar mass of sodium acetate

Calculating the molar mass is essential for converting grams to moles:
- Sodium (Na): 22.99 g/mol
- Carbon (C): 12.01 g/mol (x2 for two carbons)
- Hydrogen (H): 1.008 g/mol (x3 for three hydrogens)
- Oxygen (O): 16.00 g/mol (x2 for two oxygens)

Total molar mass:
- Na: 22.99 g/mol
- C: $2 \times 12.01 = 24.02$ g/mol
- H: $3 \times 1.008 = 3.024$ g/mol
- O: $2 \times 16.00 = 32.00$ g/mol

Sum:
$22.99 + 24.02 + 3.024 + 32.00 \approx 82.04$ g/mol

Step 2: Convert grams to moles of sodium acetate

Using the given mass:
Number of moles = mass / molar mass = 1.90 g / 82.04 g/mol ≈ 0.02317 mol

Therefore, approximately 0.02317 moles of sodium acetate are dissolved in the solution.

Step 3: Calculate the concentration of acetate ions

The concentration (molarity) of acetate ions is:
- Volume of solution = 71.0 mL = 0.071 L

Concentration:
\[\text{[Ch}_3\text{coo}^-] = \frac{0.02317\, \text{mol}}{0.071\, \text{L}} \approx 0.326\, \text{M} \]

Thus, the acetate ion concentration is approximately 0.326 M.

Understanding Acetate Hydrolysis and Its Effect on pH

What is Hydrolysis?

Hydrolysis is the reaction where the conjugate base (acetate ion) reacts with water:
$Ch3coo^- + H2O \rightleftharpoons CH3COOH + OH^-$

This reaction produces hydroxide ions (OH−), which increase the pH of the solution, making it basic.

Equilibrium Constant: The Base Dissociation Constant (Kb)

Since acetate is the conjugate base of acetic acid, its hydrolysis is governed by the base dissociation constant (Kb), which relates to the acid dissociation constant (Ka) of acetic acid:

Ka for acetic acid ≈ 1.8 × 10−5

Kb for acetate:
$$ K_b = \frac{K_w}{K_a} $$
where K_w = 1.0 × 10−14 (at 25°C)

Calculating:
$$ K_b = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} \approx 5.56 \times 10^{-10} $$

This small value indicates that hydrolysis is limited but sufficient to influence pH at higher concentrations.

Calculating the pH of the Sodium Acetate Solution

Step 1: Write the hydrolysis equilibrium expression

For acetate:
Ch3coo− + H2O ⇌ CH3COOH + OH−

The equilibrium expression:
$$ K_b = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{Ch}_3\text{coo}^-]} $$

Assuming initial concentration of acetate is C = 0.326 M, and that x is the concentration of OH− produced at equilibrium:

$$ K_b = \frac{x^2}{C - x} $$

Since K_b is small, x is much less than C, so:
$$ C - x \approx C $$

Simplifying:
$$ K_b \approx \frac{x^2}{C} $$

Solving for x:

$$ x = \sqrt{K_b \times C} $$

Plugging in the values:
$$ x = \sqrt{(5.56 \times 10^{-10}) \times 0.326} $$
$$ x = \sqrt{1.81 \times 10^{-10}} $$
$$ x \approx 1.35 \times 10^{-5}\, \text{M} $$

This is the hydroxide ion concentration.

Step 2: Calculate pOH and pH

- pOH:
$$ \text{pOH} = -\log [OH^-] = -\log(1.35 \times 10^{-5}) \approx 4.87 $$

- pH:
$$ \text{pH} = 14 - \text{pOH} = 14 - 4.87 \approx 9.13 $$

Therefore, the pH of the solution is approximately 9.13.

Summary of Key Calculations

- Mass of sodium acetate: 1.90 g
- Molar mass of sodium acetate: 82.04 g/mol
- Moles of sodium acetate: 0.02317 mol
- Volume of solution: 71.0 mL (0.071 L)
- Concentration of acetate: 0.326 M
- Base dissociation constant (Kb): 5.56×10^{-10}
- Hydroxide ion concentration: approximately 1.35×10^{-5} M
- Resulting pH: approximately 9.13

Practical Considerations and Tips

- Always ensure accurate measurements of mass and volume for precise calculations.
- Remember that strong electrolytes like NaCl or NaOH dissociate completely, simplifying calculations.
- Weak acids and bases require equilibrium considerations, often involving Ka or Kb values.
- Use logarithmic functions carefully, and consider significant figures based on your data.
- Temperature affects equilibrium constants; calculations here assume standard room temperature (25°C).

Conclusion

Calculating the pH of a sodium acetate solution involves understanding the dissolution process, determining the concentration of acetate ions, applying hydrolysis equilibrium, and calculating the resulting hydroxide ion concentration. In this specific case, dissolving 1.90 g of sodium acetate in 71.0 mL of water results in a solution with a pH of approximately 9.13, indicating a mildly basic environment. This process exemplifies fundamental principles of acid-base chemistry and highlights how salts derived from weak acids and strong bases influence solution pH. Mastery of these calculations is essential for laboratory work, industrial applications, and understanding biological systems where buffer solutions play a critical role.

Frequently Asked Questions

How do I calculate the pH of a solution prepared by dissolving 1.90 g of sodium acetate in 71.0 mL of water?

To calculate the pH, first determine the molarity of sodium acetate, then use the hydrolysis equilibrium of its conjugate base (CH3COO-) with water, and finally apply the pH formula based on the Kb or pKa values.

What is the molarity of sodium acetate in the solution if 1.90 g is dissolved in 71.0 mL?

Calculate molarity by dividing the moles of sodium acetate (1.90 g / 82.03 g/mol) by the volume in liters (71.0 mL = 0.071 L).

What is the pKa of acetic acid, and how does it relate to calculating the pH of sodium acetate solution?

The pKa of acetic acid is approximately 4.76. Since sodium acetate is the conjugate base, you can use the Henderson-Hasselbalch equation with this pKa to find the pH.

How does the hydrolysis of sodium acetate affect the pH of the solution?

Sodium acetate hydrolyzes in water to produce acetate ions, which react with water to generate hydroxide ions, making the solution slightly basic and increasing the pH.

What is the approximate pH of the solution after dissolving 1.90 g of sodium acetate in 71.0 mL of water?

After calculations, the pH is approximately 8.9, indicating a basic solution due to the hydrolysis of acetate ions.

Why is sodium acetate considered a basic salt, and how does that influence the pH calculation?

Sodium acetate is a basic salt because its conjugate base (acetate ion) hydrolyzes in water to produce OH^- ions, resulting in a pH greater than 7.

What steps are involved in calculating the pH of this sodium acetate solution?

Steps include calculating the molarity of sodium acetate, determining the hydrolysis equilibrium, calculating the concentration of OH^- ions, and then using pH = 14 - pOH to find the pH.

Additional Resources

Calculate The pH Of A Solution Prepared By Dissolving 1.90 G Of Sodium Acetate, CH_3COONa, In 71.0 ML

Introduction

Calculating the pH of a solution prepared by dissolving sodium acetate (CH_3COONa) in water is a

fundamental exercise in understanding acid-base chemistry, buffer systems, and the principles of solution chemistry. Sodium acetate is a salt derived from acetic acid, a weak acid, and sodium hydroxide, a strong base. When dissolved, sodium acetate dissociates completely in water, producing acetate ions (CH_3COO^-) and sodium ions (Na^+). The acetate ion can undergo hydrolysis, reacting with water to produce acetic acid (CH_3COOH) and hydroxide ions (OH^-), which influences the pH of the solution. This article provides a detailed, step-by-step analysis of how to determine the pH of such a solution, offering insights into relevant concepts from equilibrium chemistry, molarity calculations, and buffer systems.

Understanding the Chemistry of Sodium Acetate

The Nature of Sodium Acetate

Sodium acetate (CH_3COONa) is classified as a salt resulting from the neutralization of a weak acid (acetic acid) with a strong base (sodium hydroxide). Its significance lies in its amphiprotic behavior in aqueous solutions; it can influence the pH through hydrolysis of the acetate ion:

$$ \text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^- $$

The equilibrium of this hydrolysis reaction determines whether the solution is basic, acidic, or neutral.

Properties and Dissociation of Sodium Acetate

- Dissociation in Water: Sodium acetate dissociates completely into Na^+ and CH_3COO^- ions:

$$ \text{CH}_3\text{COONa} \rightarrow \text{Na}^+ + \text{CH}_3\text{COO}^- $$

- Hydrolysis of the Acetate Ion: The acetate ion reacts with water, leading to a basic solution:

$$ \text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^- $$

- Buffer System: Because acetic acid and acetate form a conjugate acid-base pair, the solution acts as a buffer, resisting drastic changes in pH upon addition of small amounts of acids or bases.

Calculating the Concentration of Sodium Acetate in Solution

Mass to Molarity Conversion

The first step in calculating the pH involves determining the molarity (concentration) of sodium acetate in solution:

- Given Data:
- Mass of sodium acetate, $m = 1.90\, \text{g}$
- Volume of solution, $V = 71.0\, \text{mL} = 0.071\, \text{L}$
- Molar mass of sodium acetate, M_{CH_3COONa}:

Calculating molar mass:

$$
\begin{aligned}
M_{CH_3COONa} &= M_{C} + 3 \times M_{H} + 2 \times M_{O} + M_{Na} \\
&= 12.01\, \text{g/mol} + 3 \times 1.008\, \text{g/mol} + 2 \times 16.00\, \text{g/mol} + 22.99\, \text{g/mol} \\
&= 12.01 + 3.024 + 32.00 + 22.99 \\
&\approx 70.02\, \text{g/mol}
\end{aligned}
$$

- Moles of sodium acetate:

$$
n = \frac{m}{M} = \frac{1.90\, \text{g}}{70.02\, \text{g/mol}} \approx 0.02714\, \text{mol}
$$

- Molarity of sodium acetate:

$$
\text{Molarity} (M) = \frac{n}{V} = \frac{0.02714\, \text{mol}}{0.071\, \text{L}} \approx 0.382\, \text{M}
$$

This concentration forms the basis for subsequent pH calculations.

Understanding Hydrolysis and Equilibrium Constants

The Acid-Base Equilibrium of Acetate Ion

The hydrolysis of acetate ions is characterized by the base dissociation constant K_b, which relates to the acid dissociation constant K_a of acetic acid:

$$
K_a \times K_b = K_w
$$

where:

- K_w = 1.0×10^{-14} at 25°C
- K_a for acetic acid = 1.8×10^{-5}

- Calculating K_b:

$$
K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} \approx 5.56 \times 10^{-10}
$$

This small K_b indicates that the hydrolysis of acetate is weak, but sufficient to produce a basic solution.

The Hydrolysis Reaction and Equilibrium Expression

The hydrolysis equilibrium:

$$
\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-
$$

has an associated equilibrium expression:

$$
K_b = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]}
$$

Assuming initial acetate concentration $[\text{CH}_3\text{COO}^-]_0$ and small degree of hydrolysis \(x

\):

$$
K_b = \frac{x^2}{[\text{CH}_3\text{COO}^-]_0}
$$

The resulting hydroxide ion concentration $[\text{OH}^-]$ determines the pH.

Calculating the pH of the Sodium Acetate Solution

Step 1: Establishing Hydrolysis Equilibrium

- Initial concentration of acetate: $[\text{CH}_3\text{COO}^-]_0 \approx 0.382\, \text{M}$
- Assumption: Since K_b is small, the degree of hydrolysis x is small relative to initial concentration, allowing the approximation:

$$
K_b = \frac{x^2}{[\text{CH}_3\text{COO}^-]_0} \Rightarrow x = \sqrt{K_b \times [\text{CH}_3\text{COO}^-]_0}
$$

- Calculating x:

$$
x = \sqrt{5.56 \times 10^{-10} \times 0.382} \approx \sqrt{2.12 \times 10^{-10}} \approx 1.46 \times 10^{-5}\, \text{M}
$$

This is the hydroxide ion concentration:

$$
[\text{OH}^-] \approx 1.46 \times 10^{-5}\, \text{M}
$$

Step 2: Calculating pOH and pH

- pOH:

$$
\text{pOH} = -\log [\text{OH}^-] = -\log(1.46 \times 10^{-5}) \approx 4.84
$$

- pH:

$$
\text{pH} = 14 - \text{pOH} = 14 - 4.84 = 9.16
$$

This indicates that the solution is mildly basic.

Additional Factors Influencing pH

Temperature Dependence

The calculations are based on standard temperature conditions (25°C). Changes in temperature affect K_a, K_b, and K_w, thus slightly altering the pH.

Impact of Ionic Strength

High ionic strength can influence activity coefficients, potentially modifying the equilibrium position slightly, but for dilute solutions like this, the effect is minimal.

Precision and Experimental Variations

In practical settings, factors such as impurities, measurement precision, and the purity of reagents can affect the pH value.

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