Calculate The Concentrations Of All Species In A 1.40 M Nach3coo (sodium Acetate) Solution. The Ionization

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Understanding the ionization and resulting species in a sodium acetate (NaCH3COO) solution is fundamental in chemistry, especially in the context of aqueous solutions and buffer systems. Sodium acetate is a salt formed from the weak acid acetic acid (CH3COOH) and the strong base sodium hydroxide (NaOH). When dissolved in water, it dissociates and interacts with water molecules, leading to a mixture of ions and molecules in equilibrium. Accurately calculating the concentrations of all species involved requires a detailed understanding of acid-base equilibria, dissociation constants, and the principles of solution chemistry.

This comprehensive article will guide you through the process of calculating the concentrations of all chemical species in a 1.40 M sodium acetate solution, focusing on ionization, dissociation, and equilibrium considerations. By the end, you'll understand how to approach such calculations systematically, ensuring clarity and confidence in analyzing similar chemical systems.

Understanding Sodium Acetate and Its Solution Chemistry

What Is Sodium Acetate?

Sodium acetate (NaCH3COO) is an ionic compound composed of sodium cations (Na^+) and acetate anions (CH_3COO^-). It is typically produced through the neutralization of acetic acid with sodium hydroxide:

$$ \text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{NaCH}_3\text{COO} + \text{H}_2\text{O} $$

In aqueous solution, sodium acetate dissociates completely because sodium salts of strong bases are highly soluble and ionize fully:

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

However, the acetate ion (CH_3COO^-) is the conjugate base of a weak acid (acetic acid), and it can undergo hydrolysis in water, affecting the pH and equilibrium concentrations.

Key Concepts for Calculating Species Concentrations

Before delving into calculations, it’s essential to grasp some fundamental concepts:

- Dissociation of Sodium Acetate: Complete in water, providing Na^{+} and $CH3COO^{-}$ ions.
- Hydrolysis of Acetate Ion: The conjugate base reacts with water, producing hydroxide ions (OH^{-}) and acetic acid (CH3COOH).
- Equilibrium Constants: The acid dissociation constant (Ka) of acetic acid and the base hydrolysis constant (Kb) of acetate are crucial.
- pH and pOH: Determine the extent of hydrolysis and the concentrations of hydroxide ions.

Step-by-Step Calculation Process

To find the concentrations of all species in the solution, follow these steps:

1. Identify the initial concentration of sodium acetate (given as 1.40 M).
2. Determine the hydrolysis of acetate ions.
3. Set up equilibrium expressions based on hydrolysis.
4. Calculate the pH of the solution.
5. Compute the equilibrium concentrations of all species.

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1. Initial Concentration of Sodium Acetate

Since sodium acetate dissociates completely:

- $[Na^+]_{initial} = 1.40\, \text{M}$
- $[CH_3COO^-]_{initial} = 1.40\, \text{M}$

The initial concentration of acetic acid is zero in this salt solution, but hydrolysis will produce some acetic acid in equilibrium.

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2. Hydrolysis of Acetate Ion

The acetate ion reacts with water:

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

This reaction is governed by the equilibrium constant K_b, which relates to the acid dissociation constant K_a of acetic acid:

$$ K_a \times K_b = K_w $$

Where:

- K_a for acetic acid = 1.8×10^{-5}
- K_w (ionization constant of water) = 1.0×10^{-14}

Calculating K_b:

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

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3. Setting Up Equilibrium Expressions

Let x be the concentration of OH^- produced (which also equals the amount of acetic acid formed at equilibrium), then:

- $[CH_3COO^-]$ after hydrolysis: $1.40 - x$
- $[CH_3COOH]$ formed: x
- $[OH^-]$: x

Using the expression for hydrolysis:

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

Substitute:

$$ 5.56 \times 10^{-10} = \frac{x \times x}{1.40 - x} $$

Given the very small K_b, x is expected to be small compared to 1.40, so approximate:

$$ 1.40 - x \approx 1.40 $$

Thus:

$$ 5.56 \times 10^{-10} \approx \frac{x^2}{1.40} $$

Solving for x:

$$ x^2 = 5.56 \times 10^{-10} \times 1.40 $$
$$ x^2 = 7.78 \times 10^{-10} $$
$$ x = \sqrt{7.78 \times 10^{-10}} $$
$$ x \approx 2.79 \times 10^{-5}\, \text{M} $$

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4. Determining pH and pOH

Since x represents the concentration of OH^-, the pOH is:

$$ \text{pOH} = -\log [\text{OH}^-] = -\log (2.79 \times 10^{-5}) \approx 4.55 $$

And the pH:

$$ \text{pH} = 14 - \text{pOH} = 14 - 4.55 = 9.45 $$

This indicates the solution is mildly basic, consistent with the hydrolysis of acetate ions.

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5. Final Concentrations of All Species

Species	Concentration (M)	Explanation
$[Na^+]$	1.40	From initial dissociation, remains unchanged.
$[CH_3COO^-]$	$1.40 - 2.79\times10^{-5} \approx 1.39997$	Slight decrease due to hydrolysis.
$[CH_3COOH]$	2.79×10^{-5}	From hydrolysis.
$[OH^-]$	2.79×10^{-5}	Hydrolysis product.
$[H^+]$	$1 \times 10^{-14} / [OH^-]$ ≈ 3.58×10^{-10}	From water autoionization.

Summary of Key Results

- Sodium ions (Na^+): 1.40 M — unchanged, as sodium salts dissociate fully.
- Acetate ions (CH_3COO^-): approximately 1.39997 M — slightly reduced by hydrolysis.
- Acetic acid (CH_3COOH): approximately 2.79×10^{-5} M — produced by hydrolysis.

- Hydroxide ions (\(OH^-\)): approximately 2.79×10^{-5} M.
- pH of solution: approximately 9.45, indicating mild basicity.

Implications and Applications

Understanding the equilibrium concentrations in sodium acetate solutions is vital for various applications, including:

- Buffer solutions: Sodium acetate creates buffers with acetic acid, stabilizing pH during chemical reactions.
- Industrial processes: Precise control of ion concentrations ensures optimal product quality.
- Laboratory experiments: Accurate calculations of species concentrations inform titration and analytical procedures.

Conclusion

Calculating the concentrations of all species in a 1.40 M sodium acetate solution involves understanding dissociation, hydrolysis, and equilibrium principles. By applying the known constants \(K_a

Frequently Asked Questions

What is the initial concentration of sodium acetate in the solution?

The initial concentration of sodium acetate (NaC2H3COO) is 1.40 M, as given in the problem.

How does the ionization of sodium acetate affect the concentrations of acetate and other species in solution?

Sodium acetate dissociates completely into Na+ and C2H3COO− ions; however, the acetate ion can undergo hydrolysis, affecting the concentrations of acetic acid and acetate in equilibrium.

What is the equilibrium expression for the hydrolysis of acetate ion in water?

The equilibrium expression is $K_b = [CH_3COO^-][OH^-] / [CH_3COOH]$, or equivalently, using the acid dissociation constant: $K_a K_b = K_w$.

How do you calculate the pH of a 1.40 M sodium acetate solution considering ionization?

You set up an equilibrium for acetate hydrolysis, use known K_a for acetic acid, and solve for OH^- concentration, then find pH from pOH = -log[OH^-].

What are the concentrations of all species in solution after ionization reaches equilibrium?

At equilibrium, the concentrations of Na^+ and $C_2H_3COO^-$ are approximately 1.40 M, while the acetate ion's hydrolysis produces a small amount of CH_3COOH and OH^-, whose concentrations are determined via equilibrium calculations.

Why is it important to consider the hydrolysis of acetate when calculating species concentrations?

Because acetate is a weak base, its hydrolysis affects the pH and the equilibrium concentrations of acetic acid and hydroxide ions, which are essential for accurate calculations.

How does the ionization of sodium acetate influence the solution's pH?

The hydrolysis of acetate produces hydroxide ions, making the solution slightly basic and increasing the pH above 7.

Additional Resources

Calculate The Concentrations Of All Species In A 1.40 M Nach3coo (sodium Acetate) Solution. The Ionization

Sodium acetate ($NaCH_3COO$) is a common salt resulting from the neutralization of acetic acid (CH_3COOH) with sodium hydroxide (NaOH). When dissolved in water, sodium acetate dissociates completely, but the acetate ion (CH_3COO^-) also exhibits weak basic properties by undergoing hydrolysis. Understanding the precise concentrations of all species in a sodium acetate solution is fundamental in fields such as analytical chemistry, biochemistry, and environmental science. This article explores the detailed process of calculating the concentrations of all species in a 1.40 M sodium acetate solution, including the ionization and hydrolysis equilibria involved.

Understanding the Composition of Sodium Acetate Solution

Before delving into the calculations, it's crucial to grasp the chemistry behind sodium acetate in aqueous solution.

- Dissociation of Sodium Acetate:
Sodium acetate is a soluble salt that dissociates completely in water:
$NaCH_3COO\ (s) \rightarrow Na^+\ (aq) + CH_3COO^-\ (aq)$

- Hydrolysis of the Acetate Ion:
Although sodium acetate dissociates fully, the acetate ion (CH_3COO^-) is a weak base and can accept a proton from water:
$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$

This hydrolysis leads to a basic solution, and the equilibrium establishes a certain pH, affecting the concentrations of all species.

Step 1: Initial Concentrations of Dissolved Species

Given the initial molarity of sodium acetate ($NaCH_3COO$) as 1.40 M:

- Na^+ concentration:
Since sodium acetate dissociates completely,
$[Na^+] = 1.40$ M

- CH_3COO^- concentration:
Likewise, initially,
$[CH_3COO^-] = 1.40$ M

At this stage, we need to account for the hydrolysis of CH_3COO^-, which will shift some acetate ions back to acetic acid (CH_3COOH), changing the concentrations of species in solution.

Step 2: The Hydrolysis Equilibrium and Its Constants

The key equilibrium governing the basicity of the solution is:

$$\mathrm{CH_3COO^-} + \mathrm{H_2O} \rightleftharpoons \mathrm{CH_3COOH} + \mathrm{OH^-}$$

The equilibrium constant for this process is the base hydrolysis constant K_b of acetate, which relates to the acid dissociation constant K_a of acetic acid:

$$K_b = \frac{K_w}{K_a}$$

Where:

- K_w is the ionization constant of water (1.0×10^{-14} at 25°C)
- K_a for acetic acid is approximately 1.8×10^{-5}

Calculating K_b:

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

This small K_b indicates that acetate is a weak base.

Step 3: Establishing the Equilibrium Concentrations

Let’s define the change in concentration due to hydrolysis:

- x: the amount of acetate that hydrolyzes, producing x mol/L of acetic acid and x mol/L of hydroxide ions.

Initial concentrations:

Species	Concentration	Change	Equilibrium concentration
CH_3COO^-	1.40 M	- x	1.40 - x
CH_3COOH	0 M	+ x	x
OH^-	0 M	+ x	x

Since sodium acetate dissociates completely and hydrolysis is the only process altering acetate concentration, these are the main species at equilibrium.

Using the hydrolysis equilibrium expression:

$$
K_b = \frac{[\mathrm{CH_3COOH}][\mathrm{OH^-}]}{[\mathrm{CH_3COO^-}]}
$$

Substituting the equilibrium concentrations:

$$
K_b = \frac{x \times x}{1.40 - x} \approx \frac{x^2}{1.40}
$$

Because K_b is very small, x is expected to be much less than 1.40 M, allowing us to approximate $1.40 - x \approx 1.40$:

$$
x^2 \approx K_b \times 1.40
$$

Calculating x:

$$
x^2 = 5.56 \times 10^{-10} \times 1.40 \approx 7.78 \times 10^{-10}
$$
$$
x = \sqrt{7.78 \times 10^{-10}} \approx 2.79 \times 10^{-5} \, \text{M}
$$

This small value confirms the assumption; the hydrolysis causes only a slight shift.

Step 4: Determining the Final Species Concentrations

With $x \approx 2.79 \times 10^{-5}$ M, the equilibrium concentrations are:

- Acetate ion, CH_3COO^-:
$$
1.40 - x \approx 1.40 - 2.79 \times 10^{-5} \approx 1.39997\, \text{M}
$$
Essentially unchanged from the initial.

- Acetic acid, CH_3COOH:
\[

$$
x \approx 2.79 \times 10^{-5}\, \text{M}
$$

- Hydroxide ion, OH^-:

$$
x \approx 2.79 \times 10^{-5}\, \text{M}
$$

- Sodium ion, Na^+:

Remains at 1.40 M, as sodium does not participate in acid-base chemistry here.

Step 5: Calculating pH and pOH of the Solution

The pH of the solution is primarily determined by the concentration of hydroxide ions:

$$
\mathrm{pOH} = - \log [\mathrm{OH^-}] \approx - \log (2.79 \times 10^{-5}) \approx 4.55
$$

From pOH, the pH is:

$$
\mathrm{pH} = 14 - \mathrm{pOH} = 14 - 4.55 = 9.45
$$

The solution is mildly basic, consistent with hydrolysis of acetate ions.

Step 6: Summary of All Species Concentrations

Species	Concentration (M)	Notes
Na^+	1.40	Fully dissociated; remains unchanged
CH_3COO^-	approximately 1.39997	Slightly decreased from initial
CH_3COOH	approximately 2.79×10^{-5}	From hydrolysis of acetate
OH^-	approximately 2.79×10^{-5}	From hydrolysis
H_3O^+	Calculated via pH: $10^{(-9.45)} \approx 3.55 \times 10^{-10}$	Very low, indicating basicity

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Implications and Applications

Understanding the equilibrium and species concentrations in sodium acetate solutions is vital in various contexts:

- Buffer Preparation:
Sodium acetate is commonly used in buffer solutions at around pH 4.75; knowing the detailed concentrations helps optimize buffer capacity.

- Biochemical Reactions:
Many enzymatic reactions are pH-dependent; accurate pH calculations ensure optimal conditions.

- Environmental Chemistry:
The behavior of acetate in water bodies influences nutrient cycles and pollution management.

- Industrial Processes:
Precise knowledge of ion concentrations ensures quality control in manufacturing.

Conclusion

Calculating the concentrations of all species in a sodium acetate solution involves understanding dissociation, hydrolysis equilibria, and applying equilibrium constants judiciously. In a 1.40 M solution, the acetate ion remains largely undisturbed, with only a tiny fraction undergoing hydrolysis, producing a mildly basic environment with a pH around 9.45. The detailed approach outlined demonstrates how fundamental principles of chemistry—equilibrium constants, assumptions based on their magnitudes, and iterative reasoning—allow chemists to precisely characterize solution compositions. Such insights are not only academically interesting but are also critical in practical applications across scientific disciplines.

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