Calculate The Ph At 25c Of A 0.24m Solution Of Sodium Propionate Nac2h5co2. Note That Propionic Acid

Calculate The pH At 25°C Of A 0.24 M Solution Of Sodium Propionate ($NaC_2H_5CO_2$). Note That Propionic Acid

Understanding how to calculate the pH of a sodium propionate solution at 25°C involves delving into concepts of acid-base chemistry, the properties of salts, and the behavior of weak acids and their conjugate bases in aqueous solutions. Sodium propionate ($NaC_2H_5CO_2$) is the sodium salt of propionic acid (C_2H_5COOH), a weak carboxylic acid. When dissolved in water, sodium propionate dissociates completely, and its solution exhibits basic properties due to the hydrolysis of the conjugate base, propionate ion ($C_2H_5CO_2^-$).

This article provides a comprehensive guide to calculating the pH of a 0.24 M sodium propionate solution at 25°C, with an emphasis on understanding the chemistry involved, the relevant equilibrium expressions, and the step-by-step calculation process.

Understanding the Chemistry of Sodium Propionate in Water

1. Dissociation of Sodium Propionate

When sodium propionate dissolves in water, it dissociates completely:

- $NaC_2H_5CO_2$ (sodium propionate) → Na^+ (aq) + $C_2H_5CO_2^-$ (propionate ion)

Since sodium ions are spectator ions that do not affect the pH, the focus is on the propionate ion's hydrolysis.

2. Hydrolysis of the Propionate Ion

The conjugate base of a weak acid (propionic acid) undergoes hydrolysis:

$C_2H_5CO_2^- + H_2O \rightleftharpoons C_2H_5COOH + OH^-$

This equilibrium produces hydroxide ions (OH^-), making the solution basic.

Key Concepts for pH Calculation

1. Acid-Base Equilibria

The hydrolysis of the propionate ion is governed by its base dissociation constant, K_b. To find the pH, we need to determine the concentration of OH^- ions produced and then convert that into pH.

2. Relationship Between Acid and Base Constants

Since propionic acid is a weak acid, its conjugate base's strength is characterized by its K_b, which is related to the acid dissociation constant (K_a) via:

$$K_b = K_w / K_a$$

where K_w is the ionization constant of water at 25°C (1.0×10^{-14}).

3. Utilizing pKa and pKa Values

The acid dissociation constant (K_a) for propionic acid is approximately 1.32×10^{-5} at 25°C, corresponding to a pKa of:

$$pKa = -\log(K_a) \approx 4.88$$

This value is essential for calculating K_b of the propionate ion.

Step-by-Step Calculation of pH

1. Find K_b for Propionate Ion

Using the relationship:

$$K_b = K_w / K_a$$

$$K_b = (1.0 \times 10^{-14}) / (1.32 \times 10^{-5}) \approx 7.58 \times 10^{-10}$$

2. Set Up Hydrolysis Expression

Let x be the concentration of OH^- ions produced during hydrolysis:

$$C_2H_5CO_2^- + H_2O \rightleftharpoons C_2H_5COOH + OH^-$$

Initial concentration of propionate ion = 0.24 M
Change in concentration = -x (for $C_2H_5CO_2^-$) and +x (for OH^- and C_2H_5COOH)

At equilibrium:

- $[C_2H_5CO_2^-] = 0.24 - x \approx 0.24$ (assuming x <- $[OH^-] = x$

The hydrolysis constant expression:

$K_b = [OH^-]^2 / [C_2H_5CO_2^-] \approx x^2 / 0.24$

3. Solve for x (OH^- concentration)

Rearranged:

$x^2 = K_b \times 0.24$
$x^2 = (7.58 \times 10^{-10}) \times 0.24 \approx 1.82 \times 10^{-10}$

Therefore:

$x = \sqrt{(1.82 \times 10^{-10})} \approx 1.35 \times 10^{-5}$ M

This is the concentration of OH^- ions in the solution.

4. Calculate pOH and pH

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

Since pH + pOH = 14 at 25°C:

$pH = 14 - 4.87 \approx 9.13$

Result: The pH of the 0.24 M sodium propionate solution at 25°C is approximately 9.13, indicating a mildly basic solution.

Additional Considerations and Factors

1. Temperature Effects

The calculation assumes a standard temperature of 25°C. Changes in temperature can alter K_a and K_b values, thereby affecting the pH.

2. Ionic Strength and Activity Coefficients

In dilute solutions, activity coefficients are close to 1, but at higher concentrations, they can influence the pH. For typical laboratory conditions, this approximation is acceptable.

3. Accurate pKa Values

Using precise pKa values from reliable sources ensures accurate calculations. Propionic acid's pKa is around 4.87–4.88.

Summary of the Calculation Process

- Identify the relevant equilibrium: hydrolysis of the conjugate base.
- Calculate the base dissociation constant (K_b) from K_a and K_w.
- Set up the hydrolysis equilibrium expression and solve for OH^- concentration.
- Convert OH^- concentration to pOH, then determine pH.

This systematic approach provides a solid foundation for understanding how salts like sodium propionate influence solution pH and demonstrates the application of acid-base principles in real-world chemistry scenarios.

Conclusion

Calculating the pH of a sodium propionate solution at 25°C involves understanding the chemistry of weak acids and their conjugate bases, applying equilibrium constants, and performing careful mathematical analysis. For a 0.24 M solution of sodium propionate, the pH is approximately 9.13, reflecting its basic nature due to hydrolysis of the propionate ion. Mastery of these concepts is essential for students and professionals working in analytical chemistry, environmental science, and related fields where pH plays a critical role.

Keywords: pH calculation, sodium propionate, propionic acid, hydrolysis, weak base, equilibrium, K_b, pKa, acid-base chemistry, solution pH at 25°C

Frequently Asked Questions

What is the purpose of calculating the pH of a 0.24 M sodium propionate solution at 25°C?

Calculating the pH helps determine the solution's acidity or alkalinity, which is important for understanding its chemical behavior, especially since sodium propionate is a salt derived from a weak acid and weak base.

How does sodium propionate (NaC2H5CO2) influence the pH of its aqueous solution?

Sodium propionate acts as the salt of a weak acid (propionic acid) and a strong base (NaOH), typically resulting in a slightly basic solution with pH above 7.

What is the relationship between propionic acid and sodium propionate in pH calculations?

Propionic acid is the weak acid conjugate of sodium propionate; understanding their equilibrium helps in calculating the pH of the solution based on the weak acid's dissociation constant (Ka).

What is the value of the acid dissociation constant (Ka) for propionic acid at 25°C?

The Ka for propionic acid at 25°C is approximately 1.3×10^{-5}, which is essential for pH calculations involving its conjugate base.

How do you set up the calculation to find the pH of the sodium propionate solution?

Use the Henderson-Hasselbalch equation or perform an equilibrium calculation considering the hydrolysis of the propionate ion, using known values of Ka and the concentration of the salt.

What is the approximate pH of a 0.24 M sodium propionate solution at 25°C?

The pH is approximately 8.9 to 9.0, reflecting its basic nature due to hydrolysis of the propionate ion in water.

Why is it important to consider the hydrolysis of the propionate ion when calculating pH?

Because the propionate ion ($C_2H_5COO^-$) reacts with water to produce OH^- ions, influencing the solution's pH and making it basic.

Can you explain the steps to calculate the pH of the sodium propionate solution using the hydrolysis approach?

Yes, first write the hydrolysis reaction of C2H5COO$^-$ with water, determine the hydrolysis constant (Kb), then use the initial concentration to find $[OH^-]$, and finally calculate pH from pOH. Alternatively, use the Henderson-Hasselbalch equation with the known pKa of propionic acid.

Additional Resources

Calculate the pH at 25°C of a 0.24 M solution of sodium propionate ($NaC_2H_5CO_2$). Note that propionic acid is a weak acid with interesting buffering properties, and understanding how to determine the pH of its salt solutions provides valuable insight into acid-base chemistry, buffer systems, and solution equilibria. In this comprehensive guide, we will walk through the steps, principles, and calculations needed to accurately determine the pH of a sodium propionate solution at standard laboratory temperature.

Introduction to Sodium Propionate and Its Acid-Base Properties

Sodium propionate ($NaC_2H_5CO_2$) is the sodium salt of propionic acid (C_2H_5COOH). When dissolved in water, sodium propionate dissociates completely into sodium (Na^+) ions and propionate ($C_2H_5CO_2^-$) ions:

$$ \text{NaC}_2\text{H}_5\text{CO}_2 \rightarrow \text{Na}^+ + \text{C}_2\text{H}_5\text{CO}_2^- $$

Since Na^+ is a spectator ion, the pH of the solution primarily depends on the hydrolysis of the propionate ion, which is the conjugate base of propionic acid.

Why Focus on Propionic Acid?

Propionic acid, like other weak acids, does not ionize completely in water. Its conjugate base, propionate, has a base strength that influences the pH of solutions containing its salts. Understanding this relationship is essential for fields like buffer solution design, pharmaceutical formulation, and biochemical processes.

Fundamental Concepts Needed for pH Calculation

Before diving into calculations, let's review key concepts:

1. Acid-Base Equilibria

- Weak acid dissociation:

$$ \text{C}_2\text{H}_5\text{COOH} \leftrightarrow \text{H}^+ + \text{C}_2\text{H}_5\text{CO}_2^- $$

- Conjugate base hydrolysis (for sodium propionate):

$$ \text{C}_2\text{H}_5\text{CO}_2^- + \text{H}_2\text{O} \leftrightarrow \text{C}_2\text{H}_5\text{COOH} + \text{OH}^- $$

2. Acid Dissociation Constant (Ka)

Ka reflects the strength of propionic acid:

$$ K_a = 1.34 \times 10^{-5} \quad \text{at 25°C} $$

3. Base Dissociation Constant (Kb)

For the conjugate base:

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

where K_w is the ionization constant of water (1.0×10^{-14}).

4. pKa and Relationship to pH

- $pK_a = -\log K_a$
- The pH of buffer systems can be estimated using the Henderson-Hasselbalch equation when applicable.

Step-by-Step Calculation of pH

Step 1: Recognize the Nature of the Solution

Since sodium propionate is a salt of a weak acid, its aqueous solution will be basic due to hydrolysis of the propionate ion:

$$ \text{C}_2\text{H}_5\text{CO}_2^- + \text{H}_2\text{O} \leftrightarrow \text{C}_2\text{H}_5\text{COOH} + \text{OH}^- $$

Therefore, the pH will be greater than 7.

Step 2: Determine the Hydrolysis Equilibrium

For a salt of a weak acid, the hydrolysis equilibrium can be characterized using the base dissociation constant K_b:

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

Calculating K_b:

$$K_b = \frac{1.0 \times 10^{-14}}{1.34 \times 10^{-5}} \approx 7.46 \times 10^{-10}$$

Step 3: Set Up the Hydrolysis Expression

Let the initial concentration of sodium propionate be $C = 0.24\, \text{M}$.

At equilibrium, let x be the concentration of OH^- produced:

$$\text{C}_2\text{H}_5\text{CO}_2^- + \text{H}_2\text{O} \leftrightarrow \text{C}_2\text{H}_5\text{COOH} + \text{OH}^-$$

The hydrolysis expression:

$$K_b = \frac{[\text{C}_2\text{H}_5\text{COOH}][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{CO}_2^-]}$$

Assuming x is small relative to initial concentration:

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

Solve for x:

$$x^2 = K_b \times C$$

$$x^2 = 7.46 \times 10^{-10} \times 0.24$$

$$x^2 \approx 1.79 \times 10^{-10}$$

$$x \approx \sqrt{1.79 \times 10^{-10}} \approx 1.34 \times 10^{-5} \, \text{M}$$

This is the concentration of OH^-.

Step 4: Calculate pOH and pH

- $\text{pOH} = -\log [\text{OH}^-]$

$$\text{pOH} = -\log (1.34 \times 10^{-5}) \approx 4.87$$

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

$$\text{pH} = 14 - 4.87 = 9.13$$

Thus, the pH of the 0.24 M sodium propionate solution at 25°C is approximately 9.13.

Additional Considerations and Refinements

1. Validity of Approximations

- The assumption that $x \ll C$ is reasonable here, but if more precision is needed, a quadratic equation can be solved for greater accuracy.
- Temperature effects: The values used are standard at 25°C; deviations can alter Ka, Kb, and thus pH.

2. Buffering Capacity

Although this solution is primarily governed by hydrolysis, if you add weak acids or bases, the pH will shift according to the Henderson-Hasselbalch equation:

$$\text{pH} = pK_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]} \right)$$

3. Impact of Ionic Strength

High ionic strength can influence activity coefficients and slightly alter pH readings, but for dilute solutions like this, the effect is minimal.

Summary of Key Steps

1. Identify the species responsible for pH: Propionate ion ($C_2H_5CO_2^-$).
2. Calculate K_b: $K_b = K_w / K_a$.
3. Set up hydrolysis equilibrium: Assume x for $[OH^-]$.
4. Solve for x: Use $K_b \approx x^2 / C$.
5. Calculate pOH and pH: Convert $[OH^-]$ to pOH, then to pH.

Final Remarks

Calculating the pH of a sodium propionate solution at 25°C involves understanding the chemistry of weak acids and their conjugate bases, applying equilibrium principles, and making reasonable assumptions to simplify calculations. This method demonstrates how fundamental constants and concepts like Ka, Kb, and hydrolysis equilibria converge to give an accurate estimate of solution pH. Whether you're a student, researcher, or professional chemist, mastering these calculations is essential for designing and interpreting experiments involving buffer systems and weak electrolyte solutions.

In summary, the pH of a 0.24 M sodium propionate solution at 25°C is approximately 9.13, reflecting its basic nature resulting from hydrolysis of the conjugate base. This example underscores the importance of weak acid/base equilibria in chemistry and provides a practical framework for solving similar pH calculation problems.

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