The Ph Of A 0.26m Solution Of Pentanoic Acid Hc5h9o2 Is Measured To Be 2.71. Calculate The Acid Dissociation

The pH of a 0.26 M solution of pentanoic acid (HC5H9O2) is measured to be 2.71. Calculate the acid dissociation constant (Ka) of pentanoic acid and explore its significance in understanding the acid's strength. In this comprehensive guide, we will walk through the step-by-step process of calculating Ka, explain the concepts involved, and discuss the importance of pH and acid dissociation in chemical analysis.

Understanding the Basics: pH, Molarity, and Acid Dissociation

What is pH?

pH is a measure of the acidity or alkalinity of a solution, ranging from 0 to 14. A pH below 7 indicates an acidic solution, while a pH above 7 indicates a basic (alkaline) solution. A pH of 7 is considered neutral.

Implication of the pH Value 2.71

- Indicates a strongly acidic solution.
- Means the concentration of hydrogen ions (H^+) in the solution is high.

Molarity of the Solution (0.26 M)

This is the initial concentration of pentanoic acid (HC5H9O2) before any dissociation occurs. It helps in determining the extent of dissociation based on the pH measurement.

Pentanoic Acid (HC5H9O2): A Brief Overview

Pentanoic acid, also known as valeric acid, is a carboxylic acid with the molecular formula HC5H9O2. It is a weak acid, meaning it does not fully dissociate in aqueous solutions. Understanding its acid dissociation constant (Ka) helps in characterizing its acidity and reactivity.

Structure and Properties

- Contains a carboxyl group (-COOH) responsible for its acidic properties.
- Weak acid with a typical dissociation behavior in water.
- Used in flavors, fragrances, and as a chemical intermediate.

Calculating the Acid Dissociation Constant (Ka)

Step 1: Write the Dissociation Equation

The dissociation of pentanoic acid in water can be represented as:

```
HC5H9O2 (aq) ⇌ H⁺ (aq) + C5H9O2⁻ (aq)
```

Step 2: Determine the Concentration of Hydrogen Ions (H^+)

Using the pH value, calculate the concentration of H^+ ions:

1. Recall the formula:

   ```
   pH = -log[H⁺]
   ```

2. Rearranged to solve for $[H^+]$:

   ```
   [H⁺] = 10^(-pH)
   ```

3. Plugging in the pH value:

   ```
   [H⁺] = 10^(-2.71) ≈ 1.95 × 10^(-3) M
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Step 3: Establish the Equilibrium Concentrations

Initial concentration of pentanoic acid (before dissociation): 0.26 M

At equilibrium:

- $[HC5H9O2] = 0.26 - x \approx 0.26 - 1.95 \times 10^{-3} \approx 0.258$ M
- $[H^+] = x \approx 1.95 \times 10^{-3}$ M
- $[C5H9O2^-] = x \approx 1.95 \times 10^{-3}$ M

Since the dissociation is weak, the change 'x' is small relative to initial concentration, justifying the approximation.

Step 4: Write the Expression for Ka

The acid dissociation constant (Ka) is given by:

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Ka = ([H⁺][C5H9O2⁻]) / [HC5H9O2]
```

Step 5: Calculate Ka

- Substitute the equilibrium concentrations:

  ```
  Ka = (1.95 × 10^(-3) × 1.95 × 10^(-3)) / 0.258
  ```

- Compute numerator:

  ```
  (1.95 × 10^(-3))^2 = 3.80 × 10^(-6)
  ```

- Calculate Ka:

  ```
  Ka ≈ 3.80 × 10^(-6) / 0.258 ≈ 1.47 × 10^(-5)
  ```

Interpreting the Results and Significance of Ka

What Does the Ka Value Tell Us?

- Ka = 1.47 × 10^(-5) indicates that pentanoic acid is a weak acid.
- The smaller the Ka, the weaker the acid. Conversely, a larger Ka indicates a stronger acid.
- Values of Ka typically range from 10^(-1) to 10^(-9) for weak acids.

Comparison with Other Acids

- Strong acids like HCl have Ka values close to 1, indicating complete dissociation.
- Weak acids like pentanoic acid have Ka values much less than 1, reflecting partial dissociation.

Factors Affecting Acid Dissociation and pH

1. Nature of the Acid

- Carboxylic acids are weak acids because their conjugate bases are relatively stable.
- Substituents on the carbon chain can influence acidity.

2. Concentration of the Solution

- Higher concentrations can shift dissociation equilibrium.
- In dilute solutions, acids tend to dissociate more.

3. Temperature

- Increasing temperature can either increase or decrease Ka depending on the enthalpy change of dissociation.
- Typically, dissociation of weak acids is endothermic, so higher temperatures increase Ka.

Practical Applications of Ka and pH Measurements

1. Pharmaceutical Industry

Understanding the acidity of compounds helps in drug formulation and stability.

2. Environmental Chemistry

Analyzing acidity in water sources aids in assessing pollution levels and ecological health.

3. Food Chemistry

pH and acidity influence flavor, preservation, and safety of food products.

Summary and Final Thoughts

By measuring the pH of a solution and knowing its molarity, we can determine the acid dissociation constant (Ka), a key parameter in understanding acid strength. Our calculations show that the Ka of pentanoic acid is approximately 1.47 × 10^(-5), confirming its status as a weak acid. This value not only provides insights into the acid's behavior in aqueous solutions but also informs various practical applications across chemistry, biology, and industry.

Additional Tips for Accurate pH and Ka Calculations

- Ensure pH measurements are precise, as small errors can significantly affect Ka calculations.
- Use accurate initial concentrations and consider activity coefficients in very dilute solutions for more precise results.
- Remember that approximations are valid when 'x' is small compared to initial concentrations; otherwise, use more exact methods.

Conclusion

Calculating the acid dissociation constant from pH and molarity data is a fundamental skill in chemistry that enhances our understanding of acid strength and behavior. In the case of pentanoic acid, the derived Ka value emphasizes its weak acid nature, important for applications in various scientific fields. Mastery of these calculations enables chemists and students to interpret acid-base reactions accurately and apply this knowledge effectively in research and industry.

Frequently Asked Questions

What is the first step in calculating the acid dissociation constant (Ka) for pentanoic acid given its pH and concentration?

The first step is to determine the concentration of hydrogen ions $[H^+]$ from the pH using the formula $[H^+]$ = 10^(-pH).

How do you find the concentration of H^+ ions in this solution?

Using the pH of 2.71, $[H^+]$ = 10^(-2.71) ≈ 1.95 x 10^{-3} M.

What assumption is typically made when calculating Ka for a weak acid like pentanoic acid?

The assumption is that the degree of dissociation is small, so the initial concentration of the acid remains approximately unchanged, and $[H^+] \approx [A^-]$.

How do you determine the concentration of the undissociated acid (HA) after dissociation?

Subtract the concentration of dissociated ions from the initial concentration: $[HA] \approx 0.26$ M - $[H^+] \approx 0.258$ M.

What is the formula for calculating the acid dissociation constant (Ka) from the concentrations?

$Ka = [H^+] [A^-] / [HA]$, where $[H^+]$ and $[A^-]$ are approximately equal for a weak acid.

Calculate the Ka for pentanoic acid based on the given data.

$Ka \approx (1.95 \times 10^{-3}) (1.95 \times 10^{-3}) / 0.258 \approx 1.47 \times 10^{-4}$.

Why is the pH of 2.71 significant in understanding the strength of pentanoic acid?

A pH of 2.71 indicates that pentanoic acid is a weak acid, as it does not fully dissociate in solution, reflected in its relatively low Ka value.

What are common sources of error when calculating the Ka from experimental pH data?

Sources of error include inaccurate pH measurement, assumptions about dissociation degree, and neglecting activity coefficients, which can lead to inaccurate Ka calculations.

Additional Resources

Understanding the pH of a 0.26 M Solution of Pentanoic Acid ($HC_5H_9O_2$): Calculation of Acid Dissociation

When exploring the properties of weak acids like pentanoic acid, understanding how their pH relates to their concentration and degree of dissociation is fundamental in chemistry. The pH of a 0.26 M solution of pentanoic acid ($HC_5H_9O_2$) is measured to be 2.71, and this measurement provides valuable insight into the acid's dissociation behavior in aqueous solution. In this guide, we will delve into the process of calculating the acid dissociation constant (Ka) based on the given pH, concentration, and other relevant parameters, offering a comprehensive understanding of weak acid equilibria.

Introduction to Pentanoic Acid and Acid Dissociation

Pentanoic acid, also known as valeric acid, is a weak carboxylic acid with the chemical formula HC5H9O2. As a weak acid, it does not completely dissociate in water, unlike strong acids such as hydrochloric acid. Instead, it establishes an equilibrium between its molecular form and its ions:

$$\mathrm{HC_5H_9O_2} \leftrightarrow \mathrm{H^+} + \mathrm{C_5H_9O_2^-}$$

Understanding the extent of this dissociation is crucial for predicting the acid's behavior in various chemical environments, such as biological systems, industrial processes, and environmental contexts.

Step 1: Interpreting the Given Data

Before performing calculations, it’s essential to understand the provided data:

- Concentration of pentanoic acid, $[\mathrm{HC5H9O2}]_0$ = 0.26 M
- Measured pH = 2.71

From the pH, we can determine the concentration of hydrogen ions $[\mathrm{H^+}]$ in solution, which is directly related to the degree of acid dissociation.

Step 2: Calculating $[\mathrm{H^+}]$ from pH

The pH of a solution is defined as:

$$\mathrm{pH} = - \log_{10} [\mathrm{H^+}]$$

Rearranged to find $[\mathrm{H^+}]$:

$$[\mathrm{H^+}] = 10^{-\mathrm{pH}}$$

Plugging in the known pH value:

$$[\mathrm{H^+}] = 10^{-2.71}$$

Calculating:

$$[\mathrm{H^+}] \approx 1.95 \times 10^{-3} \ \text{M}$$

This concentration reflects the amount of hydrogen ions produced by the dissociation of pentanoic acid in the solution.

Step 3: Establishing the Dissociation Equilibrium

Let’s define the dissociation process:

$$ \mathrm{HC_5H_9O_2} \leftrightarrow \mathrm{H^+} + \mathrm{C_5H_9O_2^-} $$

Assuming initial concentration of the acid is 0.26 M, and the amount dissociated is x:

- Initial concentrations:
- $[HC_5H_9O_2]$ = 0.26 M
- $[\mathrm{H^+}]$ = 0 M
- $[\mathrm{C_5H_9O_2}^-]$ = 0 M

- At equilibrium:
- $[HC_5H_9O_2]$ = $0.26 - x$
- $[\mathrm{H^+}]$ = x
- $[\mathrm{C_5H_9O_2}^-]$ = x

From the previous step, $[\mathrm{H^+}]$ is approximately 1.95×10^{-3} M, so:

$$ x \approx 1.95 \times 10^{-3} \ \text{M} $$

This value represents the amount of acid that has dissociated.

Step 4: Calculating the Acid Dissociation Constant (Ka)

The acid dissociation constant, K_a, quantifies the strength of a weak acid. It is expressed as:

$$ K_a = \frac{[\mathrm{H^+}][\mathrm{A^-}]}{[\mathrm{HA}]} $$

Where:
- $[\mathrm{H^+}]$ = concentration of hydrogen ions at equilibrium
- $[\mathrm{A^-}]$ = concentration of conjugate base (pentanoate ion)
- $[\mathrm{HA}]$ = concentration of undissociated acid

Substituting the known values:

$$ K_a = \frac{(1.95 \times 10^{-3})(1.95 \times 10^{-3})}{0.26 - 1.95 \times 10^{-3}} $$

Calculating numerator:

$$ (1.95 \times 10^{-3})^2 = 3.80 \times 10^{-6} $$

Calculating denominator:

$$ 0.26 - 1.95 \times 10^{-3} \approx 0.25805 $$

Now, compute K_a:

$$ K_a \approx \frac{3.80 \times 10^{-6}}{0.25805} $$

$$ K_a \approx 1.47 \times 10^{-5} $$

This value indicates the extent of dissociation of pentanoic acid in the solution.

Summary of the Calculation Process

- Given Data:
- Solution concentration: 0.26 M
- pH: 2.71
- Stepwise Approach:
1. Convert pH to $[\mathrm{H^+}]$ using $[\mathrm{H^+}] = 10^{-\mathrm{pH}}$.
2. Assume dissociation x equals $[\mathrm{H^+}]$.
3. Write the equilibrium expression for Ka.
4. Substitute known values and solve for Ka.

Implications and Significance

The calculated K_a value of approximately 1.47×10^{-5} aligns with typical weak carboxylic acids, which generally have K_a values ranging between 10^{-6} and 10^{-3}. This indicates that pentanoic acid is moderately weak, dissociating only partially in aqueous solution.

Understanding K_a allows chemists and students to predict the behavior of pentanoic acid in different contexts, such as:

- Buffer solutions formation
- Acid-base titrations
- Biological systems where fatty acids play roles

- Environmental impact assessments

Additional Considerations

While the calculation provides a good approximation, some factors might influence the precise K_a value:

- Temperature: K_a is temperature-dependent; the calculation assumes standard temperature (~25°C).
- Activity coefficients: For more accurate results, especially at higher concentrations, activity coefficients should be considered.
- Purity of the solution: Impurities or additional ions can affect pH measurements.

Conclusion

By examining the pH of a 0.26 M solution of pentanoic acid ($HC_5H_9O_2$) and applying fundamental principles of acid-base chemistry, we've demonstrated how to calculate the acid dissociation constant K_a. The process illustrates the importance of understanding equilibrium concepts, the relationship between pH and hydrogen ion concentration, and how these factors interplay to reveal the strength of a weak acid. Such calculations are essential tools in both academic and practical chemistry, providing insights into the behavior of organic acids in various environments.

Summary Checklist for Calculating Acid Dissociation:

- Measure or obtain the pH of the solution.
- Convert pH to $[\mathrm{H^+}]$ using $\mathrm{[H^+]} = 10^{-\mathrm{pH}}$.
- Assume $[\mathrm{H^+}]$ equals the amount of acid dissociated, x.
- Write the equilibrium expression for K_a.
- Substitute known values and compute K_a.
- Interpret the value in terms of acid strength and dissociation behavior.

This step-by-step approach ensures clarity and accuracy in analyzing weak acid solutions, fostering a deeper understanding of acid-base equilibria in organic chemistry.

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