A 0.15M Solution Of A Monoprotic Acid HA Has A Percent Ionization Of 2.1%. Determine The Acid Dissociation

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Understanding the ionization of acids in aqueous solutions is fundamental in chemistry, especially in the study of acid-base equilibria. When dealing with a monoprotic acid such as HA, knowing how much of the acid dissociates in water provides insights into its strength and behavior. This article will guide you through the process of calculating the acid dissociation constant (K_a) based on the given molarity and percent ionization, emphasizing key concepts, step-by-step calculations, and practical implications.

Introduction to Acid Ionization and K_a

What Is a Monoprotic Acid?

A monoprotic acid is an acid that donates only one proton (H^+) per molecule during ionization. Examples include hydrochloric acid (HCl), acetic acid (CH_3COOH), and formic acid (HCOOH). Their ionization in water can be represented as:

$$ \mathrm{HA} \rightleftharpoons \mathrm{H}^+ + \mathrm{A}^- $$

The extent to which HA dissociates influences the pH of the solution and is quantified by the acid dissociation constant, K_a.

What Is the Acid Dissociation Constant, K_a?

K_a is a measure of the strength of an acid in aqueous solution. It is defined as:

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

A larger K_a indicates a stronger acid that dissociates more extensively, while a smaller K_a indicates a weaker acid.

Given Data and What It Means

In this problem:

- Initial concentration of HA, $C_{initial} = 0.15\, \mathrm{M}$
- Percent ionization, $\%\ \text{ionization} = 2.1\%$

Percent ionization refers to the ratio of dissociated acid to the initial concentration, expressed as a percentage:

$$ \%\ \text{ionization} = \frac{\text{Concentration of ionized HA}}{\text{Initial concentration of HA}} \times 100 $$

Using this information allows us to compute the actual concentrations of ions at equilibrium, which are crucial for calculating K_a.

Step-by-Step Calculation of the Acid Dissociation Constant

Step 1: Calculate the Concentration of Ionized Acid

The concentration of ionized HA (i.e., the concentration of H^+ and A^- ions at equilibrium) can be calculated as:

$$ [\mathrm{H}^+] = [\mathrm{A}^-] = \left(\frac{\%\ \text{ionization}}{100} \right) \times C_{initial} $$

Plugging in the values:

$$ [\mathrm{H}^+] = \left(\frac{2.1}{100} \right) \times 0.15\, \mathrm{M} = 0.00315\, \mathrm{M} $$

This is the concentration of hydrogen ions and conjugate base formed at equilibrium.

Step 2: Determine the Remaining Concentration of HA

Since some of the initial HA dissociates, the remaining concentration is:

$$ [\mathrm{HA}]_{equilibrium} = C_{initial} - [\mathrm{H}^+] $$

$$
[\mathrm{HA}]_{equilibrium} = 0.15\, \mathrm{M} - 0.00315\, \mathrm{M} = 0.14685\, \mathrm{M}
$$

For simplicity, we can approximate this as 0.147 M.

Step 3: Write the Expression for K_a

Using the equilibrium concentrations:

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

Since $[\mathrm{H}^+] = [\mathrm{A}^-]$:

$$ K_a = \frac{(0.00315)^2}{0.147} $$

Calculating numerator:

$$ (0.00315)^2 = 9.9225 \times 10^{-6} $$

Dividing by 0.147:

$$ K_a = \frac{9.9225 \times 10^{-6}}{0.147} \approx 6.75 \times 10^{-5} $$

Therefore, the acid dissociation constant, K_a, is approximately 6.75×10^{-5}.

Interpreting the Result and Its Significance

The computed K_a indicates a weak acid, as values less than 1 denote weak acids that do not dissociate extensively in water. The value of approximately 6.75×10^{-5} suggests that only a small fraction of HA molecules dissociate at equilibrium, consistent with the given percent ionization of 2.1%.

Implications:

- pH Calculation: Using $[\mathrm{H}^+]$, the pH of the solution can be estimated:

$$ pH = -\log [\mathrm{H}^+] = -\log (0.00315) \approx 2.5 $$

- Acid Strength: The K_a value places the acid in the weak acid category, useful for predicting its behavior in various chemical contexts.

- Buffer Capacity: Knowledge of K_a and concentration helps in designing buffer solutions involving this acid.

Additional Considerations and Practical Applications

Effect of Concentration on Ionization

According to Le Châtelier's principle, increasing the concentration of the acid shifts the equilibrium, affecting ionization. Understanding this helps chemists control pH in industrial processes and laboratory settings.

Comparison with Strong Acids

Strong acids, such as HCl or H_2SO_4, have K_a values approaching 1 or higher, meaning they dissociate completely. Weak acids like HA with small K_a values dissociate only slightly, which influences their reactivity and the pH of solutions.

Experimental Determination of K_a

While calculations provide theoretical values, experimental methods include:

- Titration with a base to find the equivalence point
- Spectrophotometric methods to measure ion concentrations
- Using pH meters to determine hydrogen ion activity

Summary and Key Takeaways

- A 0.15 M solution of a monoprotic acid with 2.1% ionization results in a K_a of approximately \(6.75 \times 10^{-5}\).
- The calculation involves determining the concentration of dissociated ions and applying the equilibrium expression.
- The small K_a value confirms the weak nature of the acid and influences pH and buffer capacity.
- Understanding acid ionization and K_a is essential for applications in chemistry, biology, medicine, and industrial processes.

Final Thoughts

Mastering the calculation of acid dissociation constants from percent ionization and initial concentrations is foundational for chemists. It enables the prediction of solution behavior, pH, and interactions in complex systems. Whether designing pharmaceuticals, analyzing environmental samples, or conducting research, understanding these principles enhances both theoretical knowledge and practical skills in chemistry.

Remember: Always double-check calculations and consider experimental variability. Theoretical values serve as guides, but real-world measurements provide the most accurate insights into acid behavior.

Frequently Asked Questions

What is the formula to calculate the acid dissociation constant (Ka) for a monoprotic acid based on percent ionization?

Ka = (C × % ionization / 100)^2 / (C - (C × % ionization / 100)), where C is the initial concentration.

Given a 0.15 M solution of a monoprotic acid with 2.1% ionization, how do you determine the concentration of ionized acid (HA) at equilibrium?

The ionized concentration is 0.15 M × 2.1% = 0.00315 M.

What is the initial concentration of HA in the solution?

The initial concentration is 0.15 M, as given in the problem statement.

How do you convert percent ionization to the expression used in calculating Ka?

Percent ionization is converted to the ionized concentration by dividing the initial concentration by 100 and multiplying by the percentage: Ionized concentration = C × (% ionization / 100).

What is the step-by-step process to find the acid dissociation constant (Ka) in this problem?

First, calculate the ionized concentration (0.00315 M). Then, approximate the equilibrium concentration of HA as nearly equal to the initial concentration (since ionization is small). Finally, use the formula Ka = (ionized concentration)^2 / (initial concentration - ionized concentration) to find Ka.

Calculate the value of Ka for the acid in this problem.

Ka = (0.00315)^2 / (0.15 - 0.00315) ≈ 9.92 × 10^-6.

Why is it acceptable to approximate the denominator as the initial concentration in this calculation?

Because the percent ionization is very small (2.1%), the change in concentration due to ionization is negligible compared to the initial concentration, simplifying calculations.

Additional Resources

Acid Dissociation: Unlocking the Secrets of Monoprotic Acid Ionization

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Introduction: The Significance of Acid Dissociation in Chemistry

In the realm of chemistry, understanding how acids behave in aqueous solutions is fundamental to both theoretical and practical applications. Among the various parameters that describe acid behavior, acid dissociation (often represented by the dissociation constant, K_a) is a critical measure. It quantifies the extent to which an acid releases protons into solution, thus dictating its strength, reactivity, and role in chemical reactions.

This article examines a specific case: a 0.15 M solution of a monoprotic acid (HA) exhibiting a percent ionization of 2.1%. By analyzing this scenario, we'll determine the acid dissociation constant, K_a, and explore the broader implications of acid ionization, providing insights into how chemists interpret and utilize such data.

Understanding the Problem: Key Concepts and Definitions

Before diving into calculations, it's essential to clarify the foundational concepts involved:

1. Monoprotic Acid (HA)

A monoprotic acid is an acid that donates only one proton (H^+) per molecule during dissociation. Examples include hydrochloric acid (HCl) and acetic acid (CH_3COOH). In aqueous solution, the dissociation can be represented as:

$$
\mathrm{HA} \rightleftharpoons \mathrm{H}^+ + \mathrm{A}^-
$$

2. Concentration of the Acid

The initial concentration of the acid, before any dissociation occurs, is given as 0.15 M.

3. Percent Ionization

Percent ionization indicates what fraction of the original acid molecules dissociate into ions at equilibrium. It's calculated as:

$$
\text{Percent Ionization} = \left(\frac{\text{Concentration of ionized acid}}{\text{Initial acid concentration}} \right) \times 100
$$

In this case, it's 2.1%, meaning a very small fraction of the acid molecules dissociate.

4. Acid Dissociation Constant (K_a)

This equilibrium constant measures the strength of an acid:

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

A higher K_a indicates a stronger acid, dissociating more readily.

Step-by-Step Calculation of the Acid Dissociation Constant (K_a)

To determine K_a, we systematically analyze the data provided. The key steps involve calculating the equilibrium concentrations of ions and undissociated acid, then applying the formula for K_a.

1. Calculate the Concentration of Ionized Acid ($[\mathrm{A}^-]$ and $[\mathrm{H}^+]$)

Given:

- Initial concentration of HA, $[\mathrm{HA}]_{initial} = 0.15\, \mathrm{M}$
- Percent ionization = 2.1%

The concentration of acid that dissociates is:

$$
[\mathrm{A}^-] = [\mathrm{H}^+] = \frac{\text{Percent ionization}}{100} \times [\mathrm{HA}]_{initial}
$$

Calculating:

$$
[\mathrm{A}^-] = \frac{2.1}{100} \times 0.15\, \mathrm{M} = 0.0021 \times 0.15\, \mathrm{M} = 0.00315\, \mathrm{M}
$$

Similarly, the concentration of free H^+ ions at equilibrium is also 0.00315 M, since each dissociation produces one H^+ and one A^-.

2. Determine the Remaining Concentration of Undissociated Acid ($[\mathrm{HA}]$)

The remaining concentration of HA is:

$$
[\mathrm{HA}]_{equilibrium} = [\mathrm{HA}]_{initial} - [\mathrm{A}^-]
$$

$$
[\mathrm{HA}]_{equilibrium} = 0.15\, \mathrm{M} - 0.00315\, \mathrm{M} = 0.14685\, \mathrm{M}
$$

3. Calculate K_a

Using the formula:

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

Substituting the values:

$$

$$
K_a = \frac{(0.00315)(0.00315)}{0.14685}
$$

Calculating numerator:

$$
(0.00315)^2 = 9.9225 \times 10^{-6}
$$

Now, compute K_a:

$$
K_a = \frac{9.9225 \times 10^{-6}}{0.14685} \approx 6.76 \times 10^{-5}
$$

Interpreting the Results: What Does K_a Tell Us?

The calculated $K_a \approx 6.76 \times 10^{-5}$ indicates the acid's relative strength. To contextualize:

- Weak Acid: Since $K_a < 1 \times 10^{-3}$, this acid is considered weak, not dissociating extensively in solution.
- Comparison to Strong Acids: Strong acids like HCl have $K_a \sim 1$, meaning nearly complete dissociation—an order of magnitude higher than our value.
- Implication of Low Percent Ionization: The 2.1% ionization aligns with the low K_a, confirming the acid's weak nature.

This value is typical of weak monoprotic acids such as acetic acid or similar weak acids, which only partially dissociate under standard conditions.

Broader Implications and Applications

Understanding K_a is crucial for multiple domains:

1. Acid Strength and Buffer Design

Chemists utilize K_a to design buffer solutions by selecting acids with suitable dissociation constants to maintain pH stability in various systems, from biological to industrial.

2. pH Calculations

Knowing K_a allows accurate pH predictions of solutions:

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

In this case:

$$
\mathrm{pH} = -\log(0.00315) \approx 2.50
$$

which correlates with the initial concentration and ionization levels.

3. Titration and Analytical Chemistry

Determining K_a helps in titration analysis, especially when assessing the concentration of unknown acids or analyzing reaction equilibria.

4. Environmental and Biological Relevance

Weak acids with low K_a values are prevalent in natural systems, influencing soil chemistry, blood buffering capacity, and pollutant behavior.

Limitations and Considerations in Real-World Scenarios

While the calculation provides a solid estimate, real-world conditions introduce complexities:

- Activity vs. Concentration: Ionic interactions can affect activity coefficients, altering effective dissociation.
- Temperature Dependence: K_a varies with temperature; this calculation assumes room temperature (~25°C).
- Dilution Effects: At very dilute or very concentrated solutions, deviations may occur.
- Experimental Error: Percent ionization measurements depend on precise analytical techniques.

Conclusion: The Power of Quantitative Acid Analysis

By dissecting the scenario of a 0.15 M monoprotic acid with a 2.1% ionization, we've demonstrated how fundamental principles of equilibrium chemistry enable the determination of the acid dissociation constant, K_a. The resulting value, approximately 6.76×10^{-5}, confirms the weak acid nature of the solution and provides a window into its chemical behavior.

This analysis exemplifies how seemingly simple data points—concentration and percent ionization—can yield profound insights into acid strength, reactivity, and solution behavior. Whether designing buffers, predicting pH, or understanding environmental processes, the ability to calculate and interpret K_a remains an indispensable skill for chemists and students alike.

In essence, mastering acid dissociation calculations unlocks a deeper appreciation of the delicate balance governing chemical equilibria, reinforcing the importance of quantitative analysis in chemistry's vast and diverse landscape.

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