The A Of A Monoprotic Weak Acid Is 0.00336. What Is The Percent Ionization Of A 0.199 M Solution Of This

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Understanding the degree of ionization of weak acids in aqueous solutions is fundamental in chemistry, particularly in acid-base chemistry. Given the acid dissociation constant (Ka) value, which in this case is 0.00336, and the initial concentration of the acid, 0.199 M, we can determine the percent ionization of the acid in solution. This article delves into the concepts, calculations, and significance of this process, providing a comprehensive understanding of how to approach such problems.

Introduction to Acid Ionization and Ka

What Is a Monoprotic Weak Acid?

A monoprotic weak acid is an acid that can donate one proton (H^+) per molecule and does not fully ionize in aqueous solution. Examples include acetic acid (CH_3COOH) and formic acid ($HCHO_2$). Unlike strong acids, which dissociate completely, weak acids establish an equilibrium between their undissociated and dissociated forms.

The Acid Dissociation Constant (Ka)

Ka quantifies the extent of ionization of a weak acid in solution. It is defined as:

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

where:

- $[\mathrm{H}^+]$ is the concentration of hydrogen ions,
- $[\mathrm{A}^-]$ is the concentration of conjugate base,
- $[\mathrm{HA}]$ is the concentration of the undissociated acid.

A smaller Ka indicates a weaker acid, with less ionization, whereas a larger Ka indicates a stronger acid.

Given Data and Objective

- Ka (or "A" in the problem statement) = 0.00336
- Initial concentration of the acid, C_0 = 0.199 M
- Goal: Calculate the percent ionization of this acid in solution.

Understanding Percent Ionization

Definition of Percent Ionization

Percent ionization expresses the fraction of the original acid molecules that have dissociated into ions, expressed as a percentage:

$$\text{Percent Ionization} = \left(\frac{[\mathrm{H}^+]}{C_0} \right) \times 100$$

Since for a monoprotic acid, $[\mathrm{H}^+]$ is approximately equal to $[\mathrm{A}^-]$, the calculation involves determining $[\mathrm{H}^+]$ at equilibrium.

Significance of Percent Ionization

Knowing the percent ionization helps in understanding the acidity strength and behavior of the solution, which is important in various applications like titrations, biological systems, and industrial processes.

Step-by-Step Calculation of Percent Ionization

1. Establishing the ICE Table

To analyze the equilibrium, set up an ICE (Initial, Change, Equilibrium) table:

Species	Initial (M)	Change (M)	Equilibrium (M)
HA (acid)	0.199	-x	0.199 - x
H^+ (protons)	0	+x	x
A^- (conjugate base)	0	+x	x

Since the acid is weak, x (the amount ionized) is small compared to initial concentration, but we need to verify this through calculation.

2. Write the Expression for Ka

Using the equilibrium concentrations:
$$ \mathrm{Ka} = \frac{[\mathrm{H}^+][\mathrm{A}^-]}{[\mathrm{HA}]} = \frac{x \times x}{0.199 - x} = \frac{x^2}{0.199 - x} $$

Given that $\mathrm{Ka} = 0.00336$, the equation becomes:
$$ 0.00336 = \frac{x^2}{0.199 - x} $$

3. Making an Approximation

Because Ka is small, we can assume $x \ll 0.199$, leading to:
$$ 0.199 - x \approx 0.199 $$

This simplifies the equation to:
$$ 0.00336 \approx \frac{x^2}{0.199} $$
$$ x^2 \approx 0.00336 \times 0.199 $$
$$ x^2 \approx 0.00066864 $$
$$ x \approx \sqrt{0.00066864} $$
$$ x \approx 0.02586 \text{ M} $$

This is the concentration of $[\mathrm{H}^+]$ at equilibrium.

4. Calculating Percent Ionization

Now, using the initial concentration:
$$ \text{Percent Ionization} = \left(\frac{x}{C_0} \right) \times 100 = \left(\frac{0.02586}{0.199} \right) \times 100 $$

Calculating:
$$ \frac{0.02586}{0.199} \approx 0.130 $$
$$ \text{Percent Ionization} \approx 0.130 \times 100 = 13.0\% $$

Therefore, the percent ionization of the 0.199 M solution of this weak acid is approximately 13.0%.

Discussion of Results and Validity of Approximations

Is the Approximation Valid?

The initial assumption that $x \ll 0.199$ was used to simplify calculations. To verify, compare x to the initial concentration:
$$ \frac{0.02586}{0.199} \approx 0.13 $$
Since 13% of the initial concentration dissociates, the approximation is reasonably valid because x is about 13% of the initial value, just at

the threshold where the approximation remains acceptable.

For higher accuracy, a quadratic solution can be considered:

$$ 0.00336 = \frac{x^2}{0.199 - x} $$
$$ 0.00336 (0.199 - x) = x^2 $$
$$ 0.00066864 - 0.00336 x = x^2 $$
$$ x^2 + 0.00336 x - 0.00066864 = 0 $$

Using quadratic formula:
$$ x = \frac{-0.00336 \pm \sqrt{(0.00336)^2 - 4 \times 1 \times (-0.00066864)}}{2} $$
$$ x = \frac{-0.00336 \pm \sqrt{0.00001129 + 0.00267456}}{2} $$
$$ x = \frac{-0.00336 \pm \sqrt{0.00268585}}{2} $$
$$ x = \frac{-0.00336 \pm 0.05183}{2} $$

Discarding the negative root:
$$ x = \frac{0.05183 - 0.00336}{2} = \frac{0.04847}{2} = 0.02424 \text{ M} $$

This value is close to the previous estimate (0.02586 M), confirming the approximation's validity.

Implications and Applications

Understanding Acid Strength

The percent ionization indicates how much of the acid dissociates in solution. A 13% ionization reflects a weak acid with significant but not complete dissociation.

Practical Uses

- In titrations, knowing the ionization helps in selecting appropriate titrant concentrations.
- In biological systems, weak acids' behavior influences pH regulation.
- In industrial processes, controlling ionization impacts reaction efficiencies.

Summary and Key Takeaways

- The Ka value provides a quantitative measure of acid strength and degree of ionization.
- For a 0.199 M solution of a weak acid with Ka = 0.00336, the percent

ionization is approximately 13.0%.

- The assumption $x \ll C_0$ simplifies calculations and is valid here, but quadratic solutions confirm accuracy.
- Understanding ionization is essential for applications across chemistry, biology, and industry.

Conclusion

Determining the percent ionization of a weak acid involves understanding the equilibrium established in solution. By applying the Ka expression, making reasonable approximations, and verifying with quadratic solutions, chemists can accurately assess acid strength and behavior. The calculated 13% ionization for this particular weak acid solution highlights its moderate dissociation in water, providing insights into its chemical properties and practical applications.

This detailed analysis demonstrates the importance of fundamental concepts like equilibrium expressions and assumptions in solving real-world chemistry problems, enabling scientists and

Frequently Asked Questions

What is the definition of percent ionization of a weak acid?

Percent ionization of a weak acid is the ratio of the concentration of ionized acid to the initial concentration of the acid, multiplied by 100.

Given the acid dissociation constant (Ka) of 0.00336, how can we determine the degree of ionization?

We can set up an expression for Ka = $[H^+][A^-]/[HA]$, then use the initial concentration and the degree of ionization to find the ionization percentage.

What is the initial concentration of the weak acid in the solution?

The initial concentration of the weak acid is 0.199 M.

How do we calculate the concentration of ionized acid (H+) in this problem?

We assume the degree of ionization is small, so $[H^+] \approx x$, and use the expression $K_a = x^2 / (\text{initial concentration} - x)$.

What is the formula to find the percent ionization of the weak acid?

Percent ionization = (concentration of ionized acid / initial concentration) × 100 = $(x / 0.199) \times 100$.

How do you solve for x (the concentration of H+) using the given Ka and initial concentration?

Set up the equation $K_a = x^2 / (\text{initial concentration} - x)$, then solve for x, assuming x is small compared to initial concentration.

What approximate value of x (H+ concentration) can be used to find the percent ionization in this case?

Since Ka is small, x is approximately $\sqrt{K_a \times \text{initial concentration}} = \sqrt{0.00336 \times 0.199} \approx 0.0258$ M.

What is the final percent ionization of the solution based on the calculation?

Percent ionization $\approx (0.0258 / 0.199) \times 100 \approx 12.97\%$, approximately 13%.

Why is it valid to assume x is small compared to the initial concentration in this calculation?

Because the value of Ka is small, indicating only a small fraction of the acid dissociates, making the approximation valid for simplifying calculations.

Additional Resources

The A Of A Monoprotic Weak Acid Is 0.00336. What Is The Percent Ionization Of A 0.199 M Solution Of This

In the realm of acid-base chemistry, understanding the degree to which a weak acid ionizes in solution is fundamental. For chemists and students alike, calculating the percent ionization provides insight into the acid's strength and behavior in various contexts. Today, we explore the process of determining the percent ionization of a monoprotic weak acid with a given acid dissociation constant (Ka) of 0.00336, when dissolved at a concentration of 0.199 molar (M). This exploration not only clarifies the calculation involved but also sheds light on the significance of percent ionization in chemical analysis and practical applications.

Understanding the Basics: What Is Percent Ionization?

Before delving into calculations, it is essential to grasp what percent ionization actually means in chemical terms.

Definition and Significance

Percent ionization refers to the proportion of an acid molecule that dissociates into its ions in a solution, expressed as a percentage. For a monoprotic acid, which donates one proton (H+), the general dissociation can be written as:

\[HA \rightleftharpoons H^+ + A^- \]

Here, percent ionization quantifies how much of the initial acid (HA) has dissociated into hydrogen ions (H+) and conjugate base (A−).

Understanding this percentage is critical because:

- It helps characterize the strength of the acid; a higher percent ionization indicates a stronger weak acid.
- It influences pH calculations, buffer capacity, and reactivity.
- It guides the design of chemical processes, pharmaceutical formulations, and environmental assessments.

The Acid Dissociation Constant (Ka): The Key Parameter

The given value, Ka = 0.00336, is central to our calculations.

What Does Ka Represent?

Ka, or the acid dissociation constant, measures the equilibrium position of the dissociation of a weak acid in water. It is expressed as:

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

- A smaller Ka indicates a weaker acid (less dissociation).
- A larger Ka indicates a stronger weak acid (more dissociation).

In our case, Ka = 0.00336 suggests a weak acid with moderate dissociation.

Relation to Percent Ionization

Percent ionization is directly related to the concentrations of ions and the initial concentration of the acid, as well as the Ka value.

Key Concept:
If the initial concentration of the acid is C_0, and the amount that dissociates is x, then:

$$ \text{Percent Ionization} = \left(\frac{x}{C_0} \right) \times 100\% $$

Setting Up the Calculation: Step-by-Step Approach

To determine the percent ionization, we need to establish the equilibrium concentrations based on the initial concentration and the dissociation constant.

Given Data:

- $C_0 = 0.199\, \text{M}$ (initial concentration of the acid)
- $Ka = 0.00336$

Assumptions and Simplifications

- The acid is monoprotic and dissociates as $HA \rightleftharpoons H^+ + A^-$.
- The initial concentration of H^+ and A^- is zero before dissociation.
- The amount of dissociation x (in molarity) is small relative to C_0, allowing us to use the approximation $C_0 - x \approx C_0$ for some calculations.

Define Variables:

- x = molarity of H^+ and A^- at equilibrium, i.e., the amount dissociated.
- $[HA]_{eq} = C_0 - x$

Given the small value of Ka, the approximation $C_0 - x \approx C_0$ often holds, simplifying calculations.

Mathematical Calculation of Ionization Degree

Let's proceed with the mathematical framework to find x, and subsequently, the percent ionization.

Expression for Ka:

$$
Ka = \frac{[H^+][A^-]}{[HA]} \approx \frac{x \times x}{C_0} = \frac{x^2}{C_0}
$$

Since $x \ll C_0$, this approximation is valid.

Calculating x:

$$
x^2 = Ka \times C_0
$$
$$
x^2 = 0.00336 \times 0.199 = 0.00066864
$$
$$
x = \sqrt{0.00066864} \approx 0.02586\, \text{M}
$$

This x value represents the concentration of H^+ ions at equilibrium.

Calculating Percent Ionization:

$$
\text{Percent Ionization} = \left(\frac{x}{C_0} \right) \times 100\%
$$
$$
= \left(\frac{0.02586}{0.199} \right) \times 100\%

\]

$$
\approx 0.130 \times 100\% = 13.0\%
$$

Thus, approximately 13.0% of the initial acid molecules dissociate in the solution.

Interpreting the Results and Their Implications

The calculated percent ionization offers several insights into the nature of this weak acid.

Strength of the Acid

A 13% ionization indicates a weak acid, but not extremely weak—meaning it dissociates enough to influence pH significantly. For context:

- Strong acids like HCl ionize nearly 100% in aqueous solution.
- Weak acids typically ionize less than 10–20%, depending on their Ka.

Given this, our acid exhibits moderate ionization, which aligns with its Ka value.

Impact on pH and Chemical Behavior

The degree of ionization directly affects the pH of the solution:

$$
pH = -\log [H^+]
$$

$$
pH = -\log 0.02586 \approx 1.59
$$

This pH indicates a solution that is acidic but not strongly so, consistent with a weak acid.

Applications and Practical Significance

Knowing the percent ionization helps in various applications:

- Buffer solutions: The degree of dissociation influences buffer capacity.
- Pharmaceuticals: Drug stability and absorption depend on ionization in

bodily fluids.
- Environmental chemistry: Acid dissociation affects soil and water acidity, impacting ecosystems.
- Industrial processes: Acid strength determines corrosion rates and catalytic activity.

Limitations and Considerations in the Calculation

While the calculation provides a good estimate, several factors can influence the precision:

- The approximation $C_0 - x \approx C_0$ may introduce minor errors if x is not negligible relative to C_0. In this case, x is about 13% of C_0, so a more precise quadratic approach could be employed for higher accuracy.
- Temperature effects: The Ka value can vary with temperature, influencing ionization degree.
- Ionic strength and activity coefficients: In real solutions, these can alter effective concentrations and dissociation behavior.

For rigorous scientific work, solving the quadratic equation without approximation ensures higher accuracy.

Conclusion: The Significance of the Calculated Percent Ionization

By analyzing the given Ka value and initial concentration, we've determined that approximately 13% of the monoprotic weak acid molecules dissociate in solution. This percentage reflects a moderate strength acid, capable of significantly impacting pH and chemical reactivity. Such calculations are vital for chemists and engineers designing processes, formulating pharmaceuticals, or assessing environmental impacts. Understanding the relationship between Ka, concentration, and percent ionization enables precise control and prediction of solution behavior, underscoring the importance of foundational principles in practical chemistry.

In essence, the 13% ionization figure encapsulates the delicate balance of dissociation equilibrium, illustrating how even weak acids can exert substantial influence in aqueous environments. Whether in laboratory research or industrial applications, mastering these calculations empowers scientists `to manipulate and harness chemical properties effectively.`

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