The Ka Value For Acetic Acid, CH3COOH(aq), Is 1.8x10^-5. Calculate The Ph Of A 2.80 M Acetic Acid Solution.PH=Calculate

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Understanding the pH of a solution is fundamental in chemistry, especially when dealing with acids and bases. Acetic acid (CH_3COOH), a weak acid commonly found in vinegar, has a known acid dissociation constant (Ka) of 1.8×10^{-5}. Knowing its Ka value allows chemists to determine the acidity of solutions at various concentrations, which is crucial in fields ranging from food chemistry to environmental science. In this article, we will explore how to calculate the pH of a 2.80 M acetic acid solution step by step, providing insights into the underlying principles of acid dissociation and equilibrium calculations.

Understanding Acetic Acid and Its Dissociation

Acetic acid, CH_3COOH, is classified as a weak acid because it does not completely dissociate in aqueous solution. Instead, it establishes an equilibrium between its undissociated form and its ions:

$$\mathrm{CH_3COOH} \leftrightarrow \mathrm{CH_3COO^-} + \mathrm{H^+}$$

The degree to which acetic acid dissociates is characterized by its acid dissociation constant, Ka, which for acetic acid is 1.8×10^{-5}. This value indicates that at equilibrium, the concentration of hydrogen ions (H^+) is relatively low compared to the initial concentration of acetic acid.

Calculating the pH of a 2.80 M Acetic Acid Solution

To determine the pH of a 2.80 M acetic acid solution, follow these key steps:

1. Identify Known Values:
- Initial concentration of acetic acid, $[\mathrm{CH_3COOH}]_0 = 2.80\, M$
- Acid dissociation constant, $K_a = 1.8 \times 10^{-5}$

2. Set Up the Equilibrium Expression:

Let x represent the concentration of $\mathrm{H^+}$ ions at equilibrium.

Since acetic acid dissociates as:

$$\mathrm{CH_3COOH} \leftrightarrow \mathrm{CH_3COO^-} + \mathrm{H^+}$$

At equilibrium:
- $[\mathrm{H^+}] = x$
- $[\mathrm{CH_3COO^-}] = x$
- $[\mathrm{CH_3COOH}] = 2.80 - x$

The equilibrium expression for K_a is:

$$K_a = \frac{[\mathrm{CH_3COO^-}][\mathrm{H^+}]}{[\mathrm{CH_3COOH}]}$$

Substituting the known values:

$$1.8 \times 10^{-5} = \frac{x \times x}{2.80 - x} = \frac{x^2}{2.80 - x}$$

3. Make an Assumption to Simplify:

Because K_a is small, assume $x \ll 2.80$, so:

$$2.80 - x \approx 2.80$$

This simplifies the calculation to:

$$1.8 \times 10^{-5} = \frac{x^2}{2.80}$$

4. Solve for x:

$$x^2 = (1.8 \times 10^{-5}) \times 2.80$$

$$x^2 = 5.04 \times 10^{-5}$$

$$x = \sqrt{5.04 \times 10^{-5}}$$

$$x \approx 0.00709\, M$$

This is the concentration of $\mathrm{H^+}$ ions in the solution.

Calculating pH from Hydrogen Ion Concentration

The pH of the solution is calculated using:

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

Substitute the value of x:

$$ \mathrm{pH} = -\log (0.00709) $$

$$ \mathrm{pH} \approx -(-2.15) $$

$$ \mathrm{pH} \approx 2.15 $$

Final Answer

The pH of a 2.80 M acetic acid solution is approximately 2.15.

Additional Considerations and Accuracy

While the assumption that $x \ll 2.80$ simplifies calculations, for higher concentrations or more precise results, it's advisable to solve the quadratic equation directly:

$$ x^2 = K_a (2.80 - x) $$

which leads to:

$$ x^2 + K_a x - K_a \times 2.80 = 0 $$

Solving this quadratic provides a more accurate value of x. However, for typical weak acid concentrations like 2.80 M, the approximation used here is sufficiently accurate.

Implications of the pH Value in Practical

Contexts

Understanding the pH of acetic acid solutions is important across various applications:

- Food Industry: Vinegar, which contains acetic acid, has a pH around 2-3, aligning with our calculation.
- Chemical Manufacturing: Precise pH measurements are crucial for reactions sensitive to acidity.
- Environmental Science: Acetic acid concentrations in natural waters can influence microbial activity and water chemistry.
- Laboratory Analysis: Accurate pH calculations are essential for preparing solutions and calibrating pH meters.

Summary

Calculating the pH of a weak acid such as acetic acid involves understanding its dissociation equilibrium, applying the Ka value, and solving for hydrogen ion concentration. With the Ka value of 1.8×10^{-5} and an initial concentration of 2.80 M, the pH is approximately 2.15. This calculation demonstrates the weak acid nature of acetic acid and showcases standard methods used in analytical chemistry for pH determination.

FAQs

- Q: Why can we assume $x \ll 2.80$ in this calculation?
Because K_a is small, indicating minimal dissociation relative to the initial concentration, making the approximation valid.

- Q: How would the pH change if the concentration of acetic acid were lower?
Lower concentration would result in a higher pH (less acidic), as fewer hydrogen ions are produced.

- Q: Is this method applicable to other weak acids?
Yes, similar principles apply, but specific Ka values must be used for each acid.

Conclusion

Understanding how to calculate the pH of weak acid solutions like acetic acid is essential in both academic and practical contexts. The key lies in leveraging the acid dissociation constant and applying equilibrium principles. With a Ka of 1.8×10^{-5}, a 2.80 M acetic acid solution exhibits a pH of approximately 2.15, reaffirming its classification as a weak acid with moderate acidity. Mastery of these calculations enhances one’s ability to analyze and manipulate chemical solutions accurately, fostering a deeper comprehension of acid-base chemistry in real-world applications.

Frequently Asked Questions

What is the Ka value for acetic acid?

The Ka value for acetic acid (CH3COOH) is 1.8 x 10^-5.

How do you calculate the pH of a 2.80 M acetic acid solution?

Use the expression for weak acid dissociation: Ka = [H+]^2 / [HA], then solve for [H+] and calculate pH = -log[H+].

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

The initial concentration is 2.80 M, as given in the problem.

How do you set up the equilibrium expression for acetic acid dissociation?

Ka = [H+][CH3COO-] / [CH3COOH], assuming x is the concentration of H+ at equilibrium, then Ka = x^2 / (initial concentration - x).

What is the approximate value of [H+] in this calculation?

Since acetic acid is a weak acid, [H+] is approximately √(Ka initial concentration), which can be used to estimate pH.

How do you calculate the pH after finding [H+]?

Calculate pH = -log[H+], using the [H+] value obtained from the equilibrium expression.

What is the approximate pH of a 2.80 M acetic acid solution?

The pH is approximately 2.40, based on the calculation using the Ka value and initial concentration.

Why does acetic acid have a higher pH compared to strong acids at the same concentration?

Because acetic acid is a weak acid with a Ka of 1.8 x 10^-5, it dissociates less completely, resulting in a higher pH.

Can the quadratic formula be used to find the exact pH in this problem?

Yes, for more precise results, you can set up and solve the quadratic equation derived from the equilibrium expression to find [H+].

Additional Resources

The Ka Value For Acetic Acid, CH3COOH(aq), Is 1.8x10^-5. Calculate The pH Of A 2.80 M Acetic Acid Solution. pH=Calculate

Understanding the acidity of acetic acid, especially through its Ka value, is essential in chemistry for predicting how it behaves in aqueous solutions. Acetic acid, a weak acid with the chemical formula CH_3COOH, is commonly found in vinegar and is widely studied in both academic and industrial contexts. Its acid dissociation constant, Ka, quantifies its strength by indicating how much it dissociates into ions in water. This article delves into the significance of the Ka value for acetic acid, how to calculate the pH of a given concentration, and the broader implications of these calculations in chemistry.

Understanding the Acid Dissociation Constant (Ka) for Acetic Acid

What is Ka?

The acid dissociation constant, Ka, is a quantitative measure of the strength of an acid in solution. It expresses the equilibrium concentration of ions produced when an acid dissociates in water. For acetic acid, the dissociation can be represented as:

$CH_3COOH\ (aq) \rightleftharpoons CH_3COO^-\ (aq) + H^+\ (aq)$

The equilibrium expression for this dissociation is:

$$Ka = [CH_3COO^-][H^+] / [CH_3COOH]$$

Given that acetic acid has a Ka of 1.8×10^{-5}, it indicates relatively weak acidity, meaning it does not fully dissociate in water.

Significance of the Ka Value

- Weak Acid Indicator: The small value of Ka (less than 1) confirms that acetic acid is a weak acid.
- Dissociation Extent: The smaller the Ka, the less the acid dissociates, and vice versa.
- pH Prediction: Knowing Ka allows chemists to estimate the pH of solutions at various concentrations.

Features of Acetic Acid Based on Ka

- Partial Dissociation: Only a small fraction of acetic acid molecules release protons into the solution.
- Buffering Capacity: Acetic acid can act as a buffer, resisting pH changes when small amounts of acids or bases are added.
- Industrial Uses: Its weak acidity makes it suitable for food preservation, pharmaceuticals, and chemical synthesis.

Calculating the pH of a 2.80 M Acetic Acid Solution

Step 1: Write the Dissociation Equation and Expression

As previously stated:

$$CH_3COOH\ (aq) \rightleftharpoons CH_3COO^-\ (aq) + H^+\ (aq)$$

The initial concentration of acetic acid is 2.80 M, and initially, the concentrations of acetate and hydrogen ions are zero.

Step 2: Set Up an ICE Table

	Initial (M)	Change (M)	Equilibrium (M)
CH_3COOH	2.80	-x	2.80 - x

CH_3COO^-	0	+x	x
H^+	0	+x	x

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

Step 3: Write the Equilibrium Expression

Using the Ka value:

$Ka = [CH_3COO^-][H^+] / [CH_3COOH] = 1.8\times10^{-5}$

Substituting the equilibrium concentrations:

$1.8\times10^{-5} = (x)(x) / (2.80 - x)$

Since Ka is small, we can assume x is much less than 2.80, simplifying to:

$1.8\times10^{-5} \approx x^2 / 2.80$

Step 4: Solve for x (H^+ concentration)

Rearranged:

$x^2 \approx 1.8\times10^{-5} \times 2.80$

$x^2 \approx 5.04\times10^{-5}$

$x \approx \sqrt{(5.04\times10^{-5})} \approx 7.09\times10^{-3}$ M

This is the concentration of hydrogen ions in the solution.

Step 5: Calculate the pH

pH is defined as:

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

$pH = -\log(7.09\times10^{-3})$

$pH \approx 2.15$

Therefore, the pH of a 2.80 M acetic acid solution is approximately 2.15.

Implications of the pH Calculation

Understanding Acid Strength in Practice

The calculated pH indicates a moderately acidic solution, though not as acidic as strong acids like HCl, which dissociate completely. This pH value aligns with typical vinegar solutions, which usually have pH values around 2.4 to 3.0.

Relevance in Industrial and Food Applications

- Food Industry: pH controls are critical for safety, flavor, and preservation.
- Chemical Synthesis: Understanding acidity guides reaction conditions and catalyst choices.
- Buffer Systems: Acetic acid's weak dissociation allows it to buffer solutions effectively, maintaining stable pH levels.

Pros and Cons of Using Ka for Acid Strength Assessment

Pros:
- Provides a quantitative measure of acid strength.
- Enables precise pH calculations for various concentrations.
- Useful in designing buffer solutions and understanding reaction equilibria.
- Applicable across a wide range of weak acids.

Cons:
- Assumes ideal behavior; real solutions may deviate due to activity coefficients.
- Small Ka values may require approximation methods, which can introduce errors at higher concentrations.
- Does not account for temperature effects; Ka varies with temperature.

Features of the Calculation Method

- Simplification Assumption: x is small relative to initial concentration, simplifying calculations.
- Applicability: Suitable for weak acids with small Ka values.
- Limitations: At higher concentrations or larger Ka, more complex quadratic equations may be necessary.

Conclusion and Broader Context

Understanding the Ka value for acetic acid is fundamental in predicting its behavior in aqueous solutions. The calculated pH of approximately 2.15 for a

2.80 M solution reflects its status as a weak acid, with partial dissociation that influences various practical applications. This exercise exemplifies how fundamental principles of equilibrium chemistry enable chemists to determine solution properties accurately, informing everything from industrial processes to food preservation.

The approach used here can be generalized to other weak acids, making the understanding of Ka and pH calculations a cornerstone of analytical chemistry. While the calculations involve assumptions for simplicity, they serve as powerful tools for approximating real-world phenomena, emphasizing the importance of theoretical knowledge in practical applications.

In summary, the Ka value of 1.8×10^{-5} for acetic acid not only characterizes its acidity but also guides the quantitative determination of solution pH, providing insights into its chemical behavior and applications across various fields.

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