Ethanoic Acid Has A Pka Of 4.75. Find The Ph Of The Solution That Results From The Addition Of 40.0 Ml

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Understanding the pH of a solution containing ethanoic acid (acetic acid) is fundamental in many chemical, biological, and industrial applications. With a pKa value of 4.75, ethanoic acid is a weak acid that partially dissociates in aqueous solutions. When a specific volume of ethanoic acid solution is added to water, calculating the resulting pH involves understanding the acid's dissociation equilibrium, molarity, and the concept of pKa. In this guide, we will walk through the detailed process of determining the pH after adding 40.0 mL of ethanoic acid to a solution, providing comprehensive insights into acid-base chemistry.

Understanding the Basics: Ethanoic Acid and pKa

What is Ethanoic Acid?

- Ethanoic acid, also known as acetic acid, has the chemical formula CH_3COOH.
- It is classified as a weak acid because it does not fully dissociate in water.
- Commonly found in vinegar, it imparts a characteristic sour taste and pungent smell.

pKa and Its Significance

- The pKa value of 4.75 indicates the acidity strength of ethanoic acid.
- Lower pKa values correspond to stronger acids; higher values indicate weaker acids.
- pKa is related to the acid dissociation constant (Ka) by the relation:

$$
pKa = -\log_{10}(Ka)
$$

- For ethanoic acid:

$$
Ka = 10^{-pKa} = 10^{-4.75} \approx 1.78 \times 10^{-5}
$$

Determining the pH: Step-by-Step Approach

To find the pH after adding 40.0 mL of ethanoic acid, we need to consider several factors:
- The initial concentration of the acid.
- The total volume of the solution after addition.
- The dissociation equilibrium of ethanoic acid.
- The use of the Henderson-Hasselbalch equation or direct equilibrium calculations.

Note: Since the concentration of the acid isn't specified in the problem statement, we will assume a typical concentration or demonstrate how to solve the problem given various concentration values.

Assumptions and Given Data

- Volume of ethanoic acid solution added: 40.0 mL
- Concentration of ethanoic acid (assuming): 0.1 M (for demonstration purposes)
- Initial volume of water or solution before addition: 100 mL (assumption)
- Total final volume: sum of initial volume plus 40.0 mL added

Note: If specific initial concentrations or volumes are provided, the calculations can be adjusted accordingly.

Calculations: Determining the Final pH

1. Calculating Moles of Ethanoic Acid Added

- Moles of acid added:

$$
n_{acid} = C_{acid} \times V_{acid}
$$

Convert volume to liters:

$$
V_{acid} = 40.0\, \text{mL} = 0.040\, \text{L}
$$

Assuming $C_{acid} = 0.1\, \text{M}$:

$$

n_{acid} = 0.1\, \text{mol/L} \times 0.040\, \text{L} = 0.004\, \text{mol}
\]

2. Calculating the Total Volume of the Solution

- Initial volume of water or solution: 100 mL = 0.100 L
- Total volume after addition:

\[
V_{total} = 0.100\, \text{L} + 0.040\, \text{L} = 0.140\, \text{L}
\]

3. Determining the Concentration of Ethanoic Acid in Final Solution

- Final concentration:

\[
C_{final} = \frac{n_{acid}}{V_{total}} = \frac{0.004\, \text{mol}}{0.140\, \text{L}} \approx 0.0286\, \text{M}
\]

4. Establishing the Equilibrium: Acid Dissociation

- For weak acids, the dissociation in water can be represented as:

\[
CH_3COOH \leftrightarrow H^+ + CH_3COO^-
\]

- The dissociation constant:

\[
Ka = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}
\]

- Assuming initial concentration of undissociated acid is approximately \(C_{final} \), and \(x \) is the amount dissociated:

\[
[H^+] = x
\]

\[
[CH_3COO^-] = x
\]

\[

$$
[CH_3COOH] \approx C_{final} - x \approx C_{final}
$$

(since x is small compared to the initial concentration)

- Substituting into the expression:

$$
Ka \approx \frac{x^2}{C_{final}}
$$

$$
x = \sqrt{Ka \times C_{final}} = \sqrt{1.78 \times 10^{-5} \times 0.0286}
$$

$$
x \approx \sqrt{5.09 \times 10^{-7}} \approx 7.13 \times 10^{-4}\, \text{M}
$$

- The pH is then:

$$
pH = -\log_{10} [H^+] = -\log_{10} (x) \approx -\log_{10} (7.13 \times 10^{-4}) \approx 3.15
$$

Result: The approximate pH of the solution after adding 40.0 mL of 0.1 M ethanoic acid to 100 mL of water is around 3.15.

Impact of Concentration and Volume Variations

The above calculations depend heavily on the initial concentration of the ethanoic acid solution and the initial volume of water. To generalize:

- Higher initial concentration: Results in a higher initial molarity, leading to a higher [H+] after dissociation, thus a lower pH.
- Larger volume of water: Dilutes the acid, reducing its molarity and increasing the pH.
- Different acid concentrations: Changing the initial molarity affects the dissociation extent and equilibrium pH.

Using the Henderson-Hasselbalch Equation for

Buffer Solutions

In cases where the solution contains both acetic acid and its conjugate base (acetate), the pH can be calculated using the Henderson-Hasselbalch equation:

$$
pH = pKa + \log_{10}\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)
$$

Where:
- $[\text{A}^-]$: concentration of acetate ion.
- $[\text{HA}]$: concentration of undissociated acetic acid.

This is particularly useful in buffer solutions or when the solution contains both forms of the acid.

Practical Applications and Relevance

Understanding how to calculate the pH after adding a specific volume of ethanoic acid is crucial in various contexts:

1. **Food Industry:** Controlling acidity in vinegar production and food preservation.

2. **Laboratory Analysis:** Preparing buffer solutions with desired pH levels.

3. **Environmental Chemistry:** Assessing the acidity of water bodies affected by organic acids.

4. **Pharmaceuticals:** Designing drug formulations where pH stability is critical.

Summary and Key Takeaways

- Ethanoic acid's pKa of 4.75 indicates moderate acidity, typical of weak acids.
- Determining the pH of a solution after adding a known volume involves calculating molarity, considering dissociation equilibria, and applying the appropriate equations.
- Assumptions about initial concentrations significantly influence the calculations; precise data are essential for accurate results.
- The Henderson-Hasselbalch equation is useful when both acid and conjugate base are present.

- Practical knowledge of acid-base chemistry is vital across diverse scientific and industrial fields.

Conclusion

Calculating the pH of a solution containing ethanoic acid after adding a specific volume is a fundamental skill in chemistry. By understanding the dissociation process, using the pKa value, and applying equilibrium principles, one can accurately determine the resulting pH. While this demonstration assumes certain initial conditions, the methodology can be adapted with precise data to achieve accurate and meaningful results. Mastery of these concepts

Frequently Asked Questions

What is the pH of a solution formed by adding 40.0 mL of ethanoic acid with a pKa of 4.75?

To determine the pH, additional information such as the concentration of the acid or the amount of solute is needed. Without concentration details, the exact pH cannot be calculated.

How does the pKa value of 4.75 for ethanoic acid relate to its acidity?

A pKa of 4.75 indicates that ethanoic acid is a weak acid, with a moderate tendency to donate protons in solution, influencing the pH accordingly.

What assumptions are made when calculating pH from the pKa of ethanoic acid?

Assumptions typically include that the acid is weak and only partially dissociates, the solution behaves ideally, and the initial concentration of the acid is known or can be estimated.

How can the volume of 40.0 mL of ethanoic acid be used to find the pH of the solution?

The volume alone is insufficient; the concentration (molarity) of the acid solution must be known to calculate the pH using the acid dissociation equilibrium and the pKa value.

What is the significance of the pKa value in determining the pH of a weak acid solution?

The pKa provides a measure of the acid's strength and helps calculate the pH when the concentration of the acid is known, using the Henderson-Hasselbalch equation or equilibrium calculations.

If the concentration of ethanoic acid is known, how can the pH be calculated from its pKa?

Using the Henderson-Hasselbalch equation: pH = pKa + log([A−]/[HA]), or by setting up an equilibrium expression to solve for [H+], leading to the pH.

Why is it important to know the exact concentration of ethanoic acid when calculating pH?

Because the pH depends on the hydrogen ion concentration derived from the acid's dissociation, which is directly related to its molarity; without this, the pH cannot be precisely determined.

Additional Resources

Ethanoic acid has a pKa of 4.75. Find the pH of the solution that results from the addition of 40.0 mL of this acid. Understanding how to determine the pH of a solution containing a weak acid like ethanoic acid is fundamental in chemistry, especially in areas such as titration, buffer solutions, and acid-base chemistry. This article provides a comprehensive guide to calculating the pH based on the given data, explaining the principles behind the calculations, and offering step-by-step instructions to arrive at the correct answer.

Introduction to Ethanoic Acid and pKa

Ethanoic acid, also known as acetic acid, is a weak organic acid with the chemical formula CH_3COOH. It is a common component of vinegar and plays a significant role in organic chemistry and industrial applications. Its acid strength is characterized by its pKa value of 4.75, which indicates the pH at which half of the acid molecules are dissociated into ions in solution.

The pKa value is crucial because it defines the acid's dissociation equilibrium:

$$ \text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+ $$

At the pKa, the concentrations of undissociated acid and its conjugate base are equal. Knowing the pKa allows chemists to predict the pH of solutions and how the acid will behave under different conditions.

Understanding the Problem

The core question is: What is the pH of a solution formed by adding 40.0 mL of ethanoic acid? To answer this, we need to know:

- The concentration of the ethanoic acid solution
- The volume of the solution
- The dissociation constant (or pKa)
- The initial molar amount of acid

Since the problem statement provides only the volume of ethanoic acid added (40.0 mL) and the pKa (4.75), we must assume or infer additional data:

- The concentration of the ethanoic acid solution (commonly provided or assumed)
- The nature of the solution (is it pure acid, or a dilute solution?)

For this guide, we will assume a typical concentration of ethanoic acid, such as 0.10 M (molarity), which is common in laboratory scenarios. If actual concentration differs, the calculation steps remain the same, but the initial molar amount would change accordingly.

Step-by-Step Calculation Approach

Step 1: Determine the moles of ethanoic acid added

Using the volume and concentration:

$$
\text{Moles of } CH_3COOH = \text{Concentration} \times \text{Volume}
$$

- Volume: 40.0 mL = 0.0400 L
- Concentration: 0.10 M

$$
\text{Moles} = 0.10 \, \text{mol/L} \times 0.0400\, \text{L} = 0.00400\, \text{mol}
$$

Step 2: Calculate the initial concentration of the acid in the solution

If the solution volume after addition remains 40.0 mL (assuming no volume change due to mixing), the molarity is:

$$
\text{Initial concentration} = \frac{\text{moles}}{\text{volume in liters}} = \frac{0.00400\, \text{mol}}{0.0400\, \text{L}} = 0.10\, \text{M}
$$

This confirms the initial concentration as 0.10 M.

Step 3: Write the dissociation equilibrium expression

The dissociation of ethanoic acid:

$$
\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+
$$

Let:

- x = concentration of H^+ ions at equilibrium
- C_0 = initial concentration of ethanoic acid (0.10 M)

Since it's a weak acid, and assuming complete initial dissociation is negligible, the equilibrium concentration of undissociated acid is approximately $C_0 - x$.

At equilibrium:

- $[\text{H}^+]$ = x
- $[\text{CH}_3\text{COO}^-]$ = x
- $[\text{CH}_3\text{COOH}]$ = $C_0 - x$

The acid dissociation constant:

$$
K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = 10^{-pKa}
$$

Given $pKa = 4.75$:

$$
K_a = 10^{-4.75} \approx 1.78 \times 10^{-5}
$$

Step 4: Set up the equilibrium expression

$$
K_a = \frac{x^2}{C_0 - x}
$$

Assuming $x \ll C_0$, which is typical for weak acids, simplifying:

$$
K_a \approx \frac{x^2}{C_0}
$$

Solve for x:

$$
x = \sqrt{K_a \times C_0}
$$

Plug in numbers:

$$
x = \sqrt{1.78 \times 10^{-5} \times 0.10} = \sqrt{1.78 \times 10^{-6}} \approx 1.33 \times 10^{-3} \text{ M}
$$

Step 5: Calculate the pH

Since x represents the $[\text{H}^+]$:

$$
\text{pH} = -\log [\text{H}^+] = -\log (1.33 \times 10^{-3}) \approx 2.88
$$

Final Result

Based on the assumptions made, the pH of the solution after adding 40.0 mL of 0.10 M ethanoic acid is approximately 2.88.

Factors Affecting the pH and Further Considerations

1. Concentration of the Acid

- If the actual concentration differs (more concentrated or dilute), the pH calculation will change accordingly.
- For more concentrated solutions, the pH will be lower (more acidic).
- For more dilute solutions, the pH will be higher (less acidic).

2. Volume and Dilution Effects

- If the solution volume increases significantly after addition (for example, mixing with water), the initial concentration decreases, affecting pH.
- Always consider total volume after mixing when performing precise calculations.

3. Assumptions in the Calculation

- The approximation $x \ll C_0$ simplifies calculations but may not hold for very dilute solutions.

- For higher accuracy, especially in dilute solutions, use the quadratic formula to solve the equilibrium expression:

$$
K_a = \frac{x^2}{C_0 - x}
$$

which leads to:

$$
x^2 + K_a x - K_a C_0 = 0
$$

and solving for x.

Practical Applications of This Calculation

Understanding how to determine pH from the pKa and initial concentration is vital in:

- Buffer solution design: Knowing how weak acids behave allows chemists to prepare buffers with desired pH.
- Titration analysis: Calculating the pH at various stages of titration.
- Industrial processes: Managing pH in manufacturing and environmental systems.
- Biological systems: pH regulation in physiological conditions.

Summary of Key Points

- The pKa value indicates the strength of a weak acid; for ethanoic acid, it's 4.75.
- To find the pH, determine the initial molarity and use the acid dissociation constant.
- Approximate calculations often assume $x \ll C_0$, but for precise work, solve the quadratic equation.
- The pH of a typical 0.10 M ethanoic acid solution is approximately 2.88, given standard assumptions.

Final Thoughts

Accurately calculating the pH of weak acid solutions like ethanoic acid hinges on understanding the dissociation equilibrium, the pKa, and initial concentrations. This process illustrates the importance of fundamental concepts in acid-base chemistry and provides a basis for more complex calculations involving buffers, titrations, and real-world chemical systems.

Always verify assumptions and consider the context of your solution to ensure precise and meaningful results.

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