Consider The Titration Of A 25.0-ml Sample Of 0.110 M Hc2h3o3 (ka=1.8105) With 0.120 M Naoh.

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Titration experiments are fundamental in analytical chemistry, enabling precise determination of unknown concentrations of acids and bases. In this article, we will explore the detailed process of titrating a 25.0-mL sample of weak acid, specifically acetic acid (Hc2h3o3), with a strong base, sodium hydroxide (NaOH). We'll examine the calculations involved, the concept of acid dissociation constant (Ka), and the significance of titration curves, all aimed at providing a comprehensive understanding of this classic laboratory procedure.

Understanding the Components of the Titration

1. The Acid: Acetic Acid (HC2H3O3)

Acetic acid, also known as vinegar acid, is a weak monoprotic acid with the chemical formula HC2H3O3. Its properties are characterized by:

- Concentration: 0.110 M
- Volume of sample: 25.0 mL
- Acid dissociation constant (Ka): 1.81×10^{-5}

The Ka value reflects the degree of dissociation of acetic acid in water, indicating it only partially ionizes, which influences the titration curve and pH calculations.

2. The Base: Sodium Hydroxide (NaOH)

NaOH is a strong base, meaning it dissociates completely in aqueous solution:

- Concentration: 0.120 M

- Role in titration: To neutralize the acetic acid and determine its concentration precisely.

Calculating the Moles of Acid and Base

1. Moles of Acetic Acid in the Sample

Given the volume and molarity, the moles of acetic acid are calculated as:

Moles of $HC_2H_3O_3$ = Concentration × Volume (in liters)
= 0.110 mol/L × 0.0250 L
= 0.00275 mol

2. Moles of NaOH Required for Complete Neutralization

Since NaOH is a strong base and reacts in a 1:1 molar ratio with acetic acid:

Moles of NaOH needed = Moles of acetic acid
= 0.00275 mol

The volume of NaOH required can then be calculated:

Volume of NaOH = Moles / Molarity
= 0.00275 mol / 0.120 mol/L
≈ 0.02292 L or 22.92 mL

This is the theoretical volume of NaOH needed to reach the equivalence point—the point where moles of acid equal moles of base.

Understanding the Titration Curve and pH Calculations

1. The pH Before Titration

Initially, the solution contains only acetic acid, a weak acid. The pH can be estimated using the acid dissociation equation:

$HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$

Using the Ka expression:

$Ka = [H^+][C_2H_3O_2^-] / [HC_2H_3O_2]$

Since the acid is weak, the initial concentration of H^+ can be approximated by:

$[H^+] \approx \sqrt{(Ka \times [acid])}$
$= \sqrt{(1.81 \times 10^{-5} \times 0.110)}$
$\approx \sqrt{(1.99 \times 10^{-6})}$
≈ 0.00141 M

The initial pH:

$pH = -\log[H^+] \approx -\log(0.00141) \approx 2.85$

This indicates an acidic solution before titration begins.

2. The pH at the Equivalence Point

At the equivalence point, all acetic acid has reacted to form acetate ions ($C_2H_3O_2^-$). Since acetate is the conjugate base of a weak acid, the solution will be slightly basic. The hydrolysis of acetate ions influences the pH:

$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_3 + OH^-$

Using the base hydrolysis constant (Kb):

$Kb = Kw / Ka = 1.0 \times 10^{-14} / 1.81 \times 10^{-5} \approx 5.52 \times 10^{-10}$

The concentration of acetate at equivalence:

[Acetate] = moles / total volume
= 0.00275 mol / (25.0 mL + volume of NaOH added)

At the equivalence point, the total volume is approximately:

25.0 mL + 22.92 mL ≈ 47.92 mL or 0.04792 L

Concentration of acetate:

≈ 0.00275 mol / 0.04792 L ≈ 0.0574 M

The hydrolysis of acetate produces OH^-, and the pOH is calculated:

$[OH^-] = \sqrt{(K_b \times [\text{Acetate}])} \approx \sqrt{(5.52 \times 10^{-10} \times 0.0574)} \approx \sqrt{(3.17 \times 10^{-11})} \approx 5.63 \times 10^{-6}$ M

pOH:

$pOH = -\log(5.63 \times 10^{-6}) \approx 5.25$

And the pH:

$pH = 14 - pOH \approx 8.75$

This slightly basic pH confirms the nature of the titration involving a weak acid and a strong base.

Interpreting the Titration Curve and End Point

1. The Shape of the Titration Curve

The titration curve of acetic acid with NaOH typically exhibits:

- Initial pH around 2.85
- A gradual rise in pH as NaOH is added

- A sharp inflection point at the equivalence point (around pH 8.75)
- Post-equivalence, the pH stabilizes at a more basic level

2. Detecting the End Point

The equivalence point is where the amount of NaOH added equals the amount of acetic acid in the sample, often detected using a pH indicator like phenolphthalein, which changes color around pH 8.2–10.0. Proper titration requires:

- Slow addition of NaOH near the expected equivalence point
- Careful observation of color change
- Recording the volume of NaOH used at the end point

Applications and Significance of Titration in Analytical Chemistry

1. Determining Unknown Concentrations

Titration allows chemists to accurately determine the concentration of an unknown solution by comparing it to a standard solution with a known concentration, as demonstrated in the acetic acid and NaOH example.

2. Quality Control and Purity Testing

Industries employ titration to assess the purity of substances, ensuring compliance with quality standards, especially in pharmaceuticals, food, and beverages.

3. Environmental Monitoring

Titration techniques are used to analyze water samples for pollutants like acids, bases, and other chemicals, contributing to environmental protection efforts.

Conclusion: The Value of Understanding Titration Chemistry

The titration of acetic acid with sodium hydroxide exemplifies fundamental principles of acid-base chemistry, including molarity calculations, pH estimation, and the interpretation of titration curves. By understanding the dissociation constants and the behavior of weak acids and their conjugates, chemists can accurately determine unknown concentrations and analyze sample compositions. Mastery of titration techniques is essential for students and professionals in analytical chemistry, environmental science, and quality control, emphasizing the importance of precision and theoretical knowledge in laboratory practice. Whether for academic purposes or industrial applications, the insights gained from this titration exemplify the critical role of chemical equilibria and stoichiometry in real-world scenarios.

Frequently Asked Questions

What is the main goal when titrating 25.0 mL of 0.110 M HC2H3O3 with 0.120 M NaOH?

The main goal is to determine the concentration of acetic acid (HC2H3O3) in the sample by neutralizing it with a known concentration of NaOH and calculating the titration endpoint and related parameters.

How do you calculate the volume of NaOH required to completely neutralize 25.0 mL of 0.110 M HC2H3O3?

Use the molarity and volume relationship: M1V1 = M2V2. So, V2 = (M1 × V1) / M2 = (0.110 mol/L × 25.0 mL) / 0.120 mol/L ≈ 22.92 mL of NaOH.

What is the significance of the Ka value (1.81×10^-5) for acetic acid in this titration?

The Ka value indicates the acid strength and helps determine the pH at various points during titration, especially at the half-equivalence point where pH equals pKa.

How do you find the pH at the equivalence point for this titration?

Since acetic acid is a weak acid titrated with a strong base, the equivalence point pH will be basic. It can be calculated using the hydrolysis of the acetate ion, considering the Kb of the conjugate base, or by an approximation based on the concentration of the salt formed.

What is the pH at the half-equivalence point during this titration?

At the half-equivalence point, the pH equals the pKa of acetic acid, which is approximately pKa = -log(1.81×10^-5) ≈ 4.74.

How does the initial pH of the acetic acid solution compare to the pH at the equivalence point?

The initial pH of the acetic acid solution is acidic, around 2.9, while the pH at the equivalence point is basic, approximately 8.7, due to the formation of the acetate ion and its hydrolysis.

What indications can be used to determine the endpoint of this titration in a laboratory setting?

The endpoint can be detected using a pH indicator that changes color near pH 8.2–8.6, such as phenolphthalein, which turns pink at the equivalence point for this titration.

Additional Resources

Consider The Titration Of A 25.0-mL Sample Of 0.110 M $HC_2H_3O_3$ With 0.120 M NaOH is a classic example that showcases fundamental principles of acid-base chemistry, titration techniques, and equilibrium calculations. This problem provides an excellent opportunity to explore how weak acids interact with strong bases, how to determine the equivalence point, and how to interpret titration curves. In this comprehensive guide, we'll walk through each step of this titration process, from understanding the chemistry involved to performing detailed calculations, ensuring that you grasp both the theoretical concepts and practical applications.

Introduction to Titration and the Chemistry of $HC_2H_3O_3$

Titration is a laboratory technique used to determine the concentration of an unknown solution by reacting it with a solution of known concentration. In this scenario, the titration of a 25.0-mL sample of 0.110 M $HC_2H_3O_3$ (acetic acid) with 0.120 M NaOH involves a weak acid (acetic acid) reacting with a strong base (sodium hydroxide).

The Chemistry of Acetic Acid ($HC_2H_3O_3$)

Acetic acid ($HC_2H_3O_3$, often written as CH_3COOH) is a weak monoprotic acid. Its dissociation in water can be represented as:

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

The acid dissociation constant (Ka) for acetic acid is given as:

$$ K_a = 1.81 \times 10^{-5} $$

This value indicates that acetic acid does not fully dissociate in water, which influences the titration curve and the pH changes during titration.

Step 1: Understanding the Titration Process

The titration process involves gradually adding the strong base, NaOH, to the acetic acid solution until the equivalence point is reached—the point where moles of base added equal moles of acid initially present.

Key Terms and Concepts:

- Initial solution: 25.0 mL of 0.110 M $HC_2H_3O_3$
- Titrant (NaOH): 0.120 M NaOH
- Equivalence point: When all acetic acid has been neutralized
- End point: The observed point where the indicator changes color (not necessarily the same as the equivalence point but very close)

Step 2: Calculating the Moles of Acetic Acid in the Initial Sample

The first step is to determine how many moles of acetic acid are present before titration begins:

$$
\text{Moles of HC}_2\text{H}_3\text{O}_3 = \text{Concentration} \times \text{Volume}
$$

Converting volume to liters:

$$
25.0\, \text{mL} = 0.0250\, \text{L}
$$

Calculating moles:

$$
\text{Moles of HC}_2\text{H}_3\text{O}_3 = 0.110\, \text{mol/L} \times
$$

0.0250\, \text{L} = 2.75 \times 10^{-3}\, \text{mol}
\]

This is the total amount of acetic acid in the initial sample.

Step 3: Determining the Volume of NaOH Needed to Reach the Equivalence Point

At the equivalence point, the moles of NaOH added will equal the initial moles of acetic acid, because the reaction proceeds as:

\[\text{HC}_2\text{H}_3\text{O}_3 + \text{OH}^- \rightarrow \text{C}_2\text{H}_3\text{O}_3^- + \text{H}_2\text{O} \]

Using the concentration of NaOH:

\[
\text{Volume of NaOH} = \frac{\text{Moles of NaOH}}{\text{Concentration of NaOH}} = \frac{2.75 \times 10^{-3}\, \text{mol}}{0.120\, \text{mol/L}} \approx 0.02292\, \text{L}
\]

Expressed in milliliters:

\[
0.02292\, \text{L} \times 1000\, \text{mL/L} \approx 22.92\, \text{mL}
\]

Thus, approximately 22.92 mL of 0.120 M NaOH are needed to reach the equivalence point.

Step 4: Analyzing the pH at Various Stages of the Titration

Understanding the pH behavior at different points is crucial for interpreting titration curves.

a) Initial pH (before titration):

Calculate the initial pH of the acetic acid solution, considering its weak acid nature.

Using the expression:

\[
K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{HC}_2\text{H}_3\text{O}_3]}
\]

Assuming initial concentration $C = 0.110\, \text{M}$, and letting x be the concentration of H^+:

$$
K_a = \frac{x^2}{C - x} \approx \frac{x^2}{C}
$$

since $x \ll C$.

Calculating:

$$
x = \sqrt{K_a \times C} = \sqrt{1.81 \times 10^{-5} \times 0.110} \approx \sqrt{1.99 \times 10^{-6}} \approx 0.00141\, \text{M}
$$

pH:

$$
\text{pH} = -\log [\text{H}^+] = -\log(0.00141) \approx 2.85
$$

b) pH at the equivalence point:

At the equivalence point, all acetic acid has been neutralized, resulting in a solution of the conjugate base (acetate ion, $\text{C}_2\text{H}_3\text{O}_3^-$). The pH here is determined by hydrolysis of acetate:

$$
\text{C}_2\text{H}_3\text{O}_3^- + \text{H}_2\text{O} \leftrightarrow \text{HC}_2\text{O}_3 + \text{OH}^-
$$

The pH can be calculated using the K_b of acetate:

$$
K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.81 \times 10^{-5}} \approx 5.52 \times 10^{-10}
$$

The concentration of acetate at the equivalence point:

$$
\text{Concentration} = \frac{\text{moles}}{\text{total volume}}
$$

Total volume:

$$
V_{total} = 25.0\, \text{mL} + 22.92\, \text{mL} \approx 47.92\, \text{mL} =
$$

$$
0.04792\, \text{L}
$$

Concentration of acetate:

$$
\frac{2.75 \times 10^{-3}\, \text{mol}}{0.04792\, \text{L}} \approx 0.0574\, \text{M}
$$

Hydrolysis of acetate:

$$
K_b = \frac{[\text{OH}^-]^2}{[\text{C}_2\text{H}_3\text{O}_3^-]} \Rightarrow [\text{OH}^-] = \sqrt{K_b \times C}
$$

$$
[\text{OH}^-] = \sqrt{5.52 \times 10^{-10} \times 0.0574} \approx \sqrt{3.17 \times 10^{-11}} \approx 5.63 \times 10^{-6}\, \text{M}
$$

pOH:

$$
-\log(5.63 \times 10^{-6}) \approx 5.25
$$

pH:

$$
14 - 5.25 = 8.75
$$

Thus, the pH at the equivalence point is approximately 8.75.

c) pH after adding some base (before the equivalence point):

The pH during titration can be calculated by considering the remaining acid and the amount of base added, using the Henderson-Hasselbalch equation:

$$
\text{pH} = pK_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]} \right)
$$

Step 5: Performing a Detailed Titration Calculation at a Given Volume

Suppose we want to find the pH after adding 10 mL of Na

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