A 25.0 ML Sample Of 0.150 M Acetic Acid Is Titrated With A 0.150 M NaOH Solution. What Is The PH After

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Understanding the titration process between acetic acid and sodium hydroxide (NaOH) is fundamental in analytical chemistry, especially when calculating the pH of a solution at various stages of titration. In this article, we will explore how to determine the pH after adding a certain volume of NaOH to a fixed volume of acetic acid, focusing on step-by-step calculations, chemical principles, and key concepts involved in acid-base titrations.

Introduction to Acetic Acid and Sodium Hydroxide Titration

Titration is a laboratory technique used to determine the concentration of an unknown solution by adding a titrant of known concentration until the reaction reaches its equivalence point. In this case, acetic acid (CH_3COOH), a weak acid, reacts with sodium hydroxide (NaOH), a strong base.

Key features of acetic acid:
- Weak acid with a typical dissociation constant (K_a) of approximately 1.8×10^{-5}.
- Partial dissociation in aqueous solution, which influences the pH before the equivalence point.

Key features of NaOH:
- Strong base that dissociates completely in water.
- Used to neutralize acids in titrations.

Understanding the Titration Process

During titration, the pH of the solution changes depending on the amount of titrant added:
- Initial stage: Acidic pH, dominated by acetic acid.
- Before equivalence point: Mixture of acetic acid and acetate ions.
- At equivalence point: Complete neutralization, solution contains only acetate ions.
- Beyond equivalence point: Excess NaOH determines the pH.

In our specific problem, we are asked to find the pH after a certain volume of NaOH has been added to the acetic acid solution, but the exact volume isn't specified. To illustrate, we will analyze three key points:
1. Before any NaOH is added.
2. At the half-equivalence point.
3. At the equivalence point.

Calculating the Initial Conditions

Given data:
- Volume of acetic acid sample, $V_1 = 25.0$ mL $= 0.025$ L
- Concentration of acetic acid, $C_1 = 0.150$ M
- Concentration of NaOH, $C_2 = 0.150$ M

Initial moles of acetic acid:
$$n_{acid} = C_1 \times V_1 = 0.150\, \text{mol/L} \times 0.025\, \text{L} = 3.75 \times 10^{-3}\, \text{mol}$$

Determining the Volume of NaOH Added

The moles of NaOH added after some volume V_2 (in mL or L):
$$n_{NaOH} = C_2 \times V_2$$

The reaction:

$$\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COO}^- + \text{H}_2\text{O}$$

- At any point, the moles of acetic acid remaining:
$$n_{acid,remaining} = n_{acid,initial} - n_{NaOH}$$
- Moles of acetate ions formed:
$$n_{acetate} = n_{NaOH}$$

Calculating pH Before the Equivalence Point

When NaOH has been added but not enough to reach the equivalence point, the solution contains a mixture of unreacted acetic acid and acetate ions (its conjugate base). The pH can be calculated using the Henderson-Hasselbalch equation:

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

Where:
- pK_a of acetic acid ≈ 4.76
- $[\text{A}^-]$ = concentration of acetate ions
- $[\text{HA}]$ = concentration of acetic acid

Case Study: Calculating pH at Different Points

2.1. Before Any NaOH Is Added

At zero volume of NaOH added:
- The solution contains only acetic acid.
- Using the initial molarity and volume:

$$

$$
\text{Initial concentration of acetic acid} = 0.150\, \text{M}
$$

Since acetic acid is a weak acid, its dissociation equilibrium:

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

Using the expression for Ka:

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

Assuming initial dissociation:

$$
[\text{H}^+] = x
$$
$$
[\text{CH}_3\text{COOH}] \approx 0.150 - x \approx 0.150
$$
$$
[\text{CH}_3\text{COO}^-] \approx x
$$

Since Ka is small, $x \ll 0.150$, so:

$$
x = \sqrt{K_a \times C_{initial}} = \sqrt{1.8 \times 10^{-5} \times 0.150} \approx 1.65 \times 10^{-3}\, \text{M}
$$

pH:

$$
pH = -\log x \approx -\log (1.65 \times 10^{-3}) \approx 2.78
$$

Therefore, initial pH ≈ 2.78.

2.2. At the Half-Equivalence Point

This occurs when half of the acetic acid has been neutralized:

$$
n_{NaOH} = 0.5 \times n_{acid,initial} = 1.88 \times 10^{-3}\, \text{mol}
$$

Corresponding volume of NaOH:

$$
V_{half} = \frac{n_{NaOH}}{C_2} = \frac{1.88 \times 10^{-3}}{0.150} \approx 0.0125\, \text{L} = 12.5\, \text{mL}
$$

At this point:
- $[HA] = [A^-]$, since half of the acid has been converted.
- pH = pKa (by the Henderson-Hasselbalch equation).

$$
pH = pK_a = 4.76
$$

This is a characteristic feature of weak acid-strong base titrations.

2.3. At the Equivalence Point

The equivalence point is reached when:

$$
n_{NaOH} = n_{acid,initial} = 3.75 \times 10^{-3}\, \text{mol}
$$

Corresponding volume:

$$
V_{equiv} = \frac{3.75 \times 10^{-3}}{0.150} = 0.025\, \text{L} = 25.0\, \text{mL}
$$

At this point:
- The solution contains only acetate ions, which form a basic solution.
- The pH is determined by the hydrolysis of acetate:

$$
\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \leftrightarrow \text{CH}_3\text{COOH} + \text{OH}^-
$$

Using the K_b for acetate:

$$
K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} \approx 5.56 \times 10^{-10}
$$

Concentration of acetate:

$$
[\text{A}^-] = \frac{n_{NaOH}}{\text{total volume}} = \frac{3.75 \times 10^{-3}}{0.025\, \text{L}} = 0.150\, \text{M}
$$

Hydrolysis of acetate:

$$
K_b = \frac{[\text{OH}^-]^2}{[\text{A}^-]}
$$
$$
[\text{OH}^-] = \sqrt{K_b \times [\text{A}^-]} = \sqrt{5.56 \times 10^{-10} \times 0.150} \approx 9.13 \times 10^{-6}\, \text{M}
$$

pOH:

$$
pOH = -\log 9.13 \times 10^{-6} \approx 5.04
$$

pH

Frequently Asked Questions

What is the pH of a 25.0 mL acetic acid solution after titrating with 0.150 M NaOH when the titration is halfway to equivalence?

At the halfway point, half of the acetic acid has been neutralized, so the pH equals the pKa of acetic acid, which is approximately 4.76.

How do you calculate the pH after adding a certain volume of NaOH to a 0.150 M acetic acid solution?

You determine the moles of acetic acid and NaOH, find the remaining acetic acid or formed acetate, and use the Henderson-Hasselbalch equation to calculate the pH based on the ratio of acetate to acetic acid.

What is the pH after adding enough NaOH to reach the equivalence point in the titration of acetic acid?

At the equivalence point, all acetic acid has been neutralized to acetate ions, and the solution's pH depends on the hydrolysis of acetate; typically, it will be slightly basic, around pH 8.7.

How does the initial concentration of acetic acid affect the pH after partial titration with NaOH?

A higher initial concentration results in a greater amount of acid present, which influences the pH after partial titration; the pH will be lower before equivalence and increase as more base is added.

What is the significance of the buffer region in the titration of acetic acid with NaOH, and how does it relate to pH changes?

The buffer region occurs around the half-equivalence point, where the solution contains significant amounts of acetic acid and acetate, resulting in relatively stable pH despite adding more base, typically near the pKa (~4.76).

Additional Resources

A 25.0 mL Sample of 0.150 M Acetic Acid Is Titrated With a 0.150 M NaOH Solution: An In-Depth Investigation into Acid-Base Titration Dynamics and pH Calculations

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Introduction

Acid-base titrations are fundamental procedures in analytical chemistry, providing insights into the concentration and properties of acids and bases through the measured neutralization reaction. The titration of

acetic acid (a weak acid) with sodium hydroxide (a strong base) exemplifies a classic acid-base interaction, illustrating key concepts such as weak acid dissociation, equivalence point dynamics, and pH calculation methodology.

This comprehensive review dissects the process of titrating a 25.0 mL sample of 0.150 M acetic acid with 0.150 M NaOH, focusing on determining the pH after a specific volume of titrant is added. The analysis involves stoichiometric calculations, equilibrium considerations, and the application of Henderson-Hasselbalch and other relevant equations to elucidate the pH at various titration stages.

Background and Theoretical Foundations

Acetic Acid and Its Properties

Acetic acid (CH_3COOH) is a weak monoprotic acid characterized by a relatively high acid dissociation constant ($K_a \approx 1.8 \times 10^{-5}$). Its dissociation in water can be represented as:

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

The weak acid dissociation equilibrium and the resulting pH are influenced by the initial concentration, temperature, and ionic strength of the solution.

Sodium Hydroxide as a Strong Base

NaOH is a strong base that dissociates completely in aqueous solution:

$$\mathrm{NaOH} \rightarrow \mathrm{Na^+} + \mathrm{OH^-}$$

This complete dissociation simplifies calculations, as the hydroxide ion concentration directly equals the molarity of NaOH added.

Titration Dynamics

In titrating acetic acid with NaOH, the following stages are typically observed:

1. Initial point: pH reflects the weak acid's dissociation.
2. Before equivalence point: pH is above 7, with the solution containing unreacted acetic acid and some acetate ions.
3. At equivalence point: all acetic acid is neutralized, and the solution contains only acetate ions and water.
4. Beyond equivalence point: excess NaOH determines the pH.

Understanding the pH at any volume of titrant added involves calculating the remaining concentrations of acetic acid and acetate, considering the dissociation equilibrium, and applying relevant equations.

Experimental Parameters and Initial Calculations

Initial Data:

- Volume of acetic acid sample, $V_{acid} = 25.0\, \text{mL} = 0.025\, \text{L}$
- Concentration of acetic acid, $C_{acid} = 0.150\, \text{M}$
- Molarity of NaOH, $C_{base} = 0.150\, \text{M}$

Moles of Acetic Acid Initially Present:

$$n_{acid} = C_{acid} \times V_{acid} = 0.150\, \text{mol/L} \times 0.025\, \text{L} = 3.75 \times 10^{-3}\, \text{mol}$$

Moles of NaOH Added:

Let V_{NaOH} be the volume of NaOH added at the point of interest. The moles of NaOH added are:

$$n_{NaOH} = C_{base} \times V_{NaOH}$$

The goal is to determine the pH after a specific volume of NaOH solution is added, which is often expressed as a titration curve.

Step-by-Step pH Calculation Methodology

1. Determining the Moles of NaOH Added

Suppose we are interested in the pH after adding V milliliters of NaOH. For the purpose of this investigation, let's analyze the pH after adding 10.0 mL of NaOH:

$$V_{NaOH} = 10.0\, \text{mL} = 0.010\, \text{L}$$

$$n_{NaOH} = 0.150\, \text{mol/L} \times 0.010\, \text{L} = 1.50 \times 10^{-3}\, \text{mol}$$

2. Comparing Moles of Acid and Base

- Remaining moles of acetic acid:

$$n_{acid, remaining} = n_{acid} - n_{NaOH} = 3.75 \times 10^{-3} - 1.50 \times 10^{-3} = 2.25 \times 10^{-3}\, \text{mol}$$

- Moles of acetate formed:

$$n_{acetate} = n_{NaOH} = 1.50 \times 10^{-3}\, \text{mol}$$

Since the base has reacted with acid, the solution contains a mixture of unreacted acetic acid and acetate ions.

3. Calculating Concentrations After Titration

Total volume of solution:

$$V_{total} = V_{acid} + V_{NaOH} = 25.0\, \text{mL} + 10.0\, \text{mL} = 35.0\, \text{mL} = 0.035\, \text{L}$$

- Concentration of remaining acetic acid:

$$[\mathrm{CH_3COOH}] = \frac{n_{acid, remaining}}{V_{total}} = \frac{2.25 \times 10^{-3}}{0.035} \approx 0.0643\, \text{M}$$

- Concentration of acetate:

$$[\mathrm{CH_3COO^-}] = \frac{1.50 \times 10^{-3}}{0.035} \approx 0.0429\, \text{M}$$

4. Applying the Henderson-Hasselbalch Equation

Given the mixture of weak acid and its conjugate base, the pH can be estimated using:

$$\mathrm{pH} = \mathrm{p}K_a + \log \left(\frac{[\mathrm{A^-}]}{[\mathrm{HA}]} \right)$$

Where:

- $\mathrm{p}K_a = -\log(1.8 \times 10^{-5}) \approx 4.74$

Plugging in the concentrations:

$$\mathrm{pH} = 4.74 + \log \left(\frac{0.0429}{0.0643} \right)$$

$$\mathrm{pH} = 4.74 + \log(0.666)$$

$$\mathrm{pH} \approx 4.74 - 0.177$$

$$\boxed{\mathrm{pH} \approx 4.56}$$

Thus, after adding 10.0 mL of NaOH, the pH is approximately 4.56.

Analysis of Titration Stages

Before the Equivalence Point

In the initial stages, the pH is governed mainly by the weak acid dissociation. As titration progresses and more base is added, the pH gradually increases, following the weak acid buffering region.

Near the Equivalence Point

The equivalence point occurs when moles of NaOH added equal initial moles of acetic acid:

$$V_{eq} = \frac{n_{acid}}{C_{NaOH}} = \frac{3.75 \times 10^{-3}}{0.150} = 0.025\, \text{L} = 25.0\, \text{mL}$$

At this point, all acetic acid has been converted into acetate, and the pH is dictated by the hydrolysis of acetate ions.

At the Equivalence Point

The pH is determined by the hydrolysis of the acetate ion:

$$\mathrm{CH_3COO^-} + \mathrm{H_2O} \rightleftharpoons \mathrm{CH_3COOH} + \mathrm{OH^-}$$

Using the hydrolysis constant K_b:

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} \approx 5.56 \times 10^{-10}$$

Concentration of acetate at equivalence:

$$[\mathrm{A^-}] = \frac{3.75 \times 10^{-3}}{0.025} = 0.15\, \text{M}$$

Hydrolysis of acetate:

$$x = [\mathrm{OH^-}]$$

$$K_b = \frac{x^2}{[\mathrm{A^-}]} \Rightarrow x = \sqrt{K_b \times [\mathrm{A^-}]}$$

$$x = \sqrt{5.56 \times 10

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