Calculate The Ph In The Tirtration Of 25 Ml Of 0.10 Acetic Acid By Sodium Hydrodixe After The Addition

Calculate The pH In The Titration Of 25 Ml Of 0.10 Acetic Acid By Sodium Hydroxide After The Addition

Understanding how to calculate the pH during titration processes is fundamental in analytical chemistry. In particular, the titration of acetic acid with sodium hydroxide is a classic example used to illustrate acid-base reactions, buffer solutions, and pH calculations. This article will guide you step-by-step on how to determine the pH after a specific volume of sodium hydroxide has been added to 25 mL of 0.10 M acetic acid. Whether you're a student preparing for exams or a chemist performing titrations in the lab, this detailed explanation will help you grasp the concepts and perform accurate calculations.

Understanding the Basics of Acid-Base Titration

The Acid-Base Reaction

In the titration of acetic acid (a weak acid) with sodium hydroxide (a strong base), the reaction proceeds as follows:

$$ \text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COONa} + \text{H}_2\text{O} $$

This reaction is a typical acid-base neutralization where acetic acid reacts with sodium hydroxide to produce sodium acetate (a salt) and water.

The Concept of Equivalence Point

The equivalence point occurs when the moles of sodium hydroxide added equal the initial moles of acetic acid present. At this point, the solution contains only the salt (sodium acetate) and water, and the pH depends on the hydrolysis of the salt.

Initial Data and Calculations

Before beginning calculations, gather all relevant data:

- Volume of acetic acid solution: 25 mL
- Concentration of acetic acid: 0.10 M
- Volume of sodium hydroxide added: This varies; for this example, assume a specific volume (e.g., 10 mL). For a comprehensive understanding, calculations at various points are provided.

Step 1: Calculate Moles of Acetic Acid

Convert the volume to liters:

$$ V_{acid} = 25 \text{ mL} = 0.025 \text{ L} $$

Calculate moles:

$$ \text{moles of acetic acid} = C_{acid} \times V_{acid} = 0.10 \times 0.025 = 0.0025 \text{ mol} $$

Step 2: Determine Moles of Sodium Hydroxide Added

Suppose you add x mL of NaOH solution. Convert to liters and calculate moles:

$$ V_{NaOH} = x \text{ mL} = \frac{x}{1000} \text{ L} $$

$$ \text{moles of NaOH} = C_{NaOH} \times V_{NaOH} = 0.10 \times \frac{x}{1000} $$

At any point, knowing the volume of NaOH added allows you to determine how much acid has reacted and what remains.

Calculating pH After Partial Neutralization

The pH depends on the extent of neutralization. We analyze different stages:

Case 1: Before the Equivalence Point (Partial Neutralization)

When less than the equivalence point volume of NaOH has been added, some acetic acid remains unreacted. The solution contains a mixture of weak acid and its conjugate base, forming a buffer.

Step 1: Determine Remaining Moles of Acetic Acid and Formed Sodium Acetate

- Moles of acetic acid remaining:

$$ \text{Remaining acid} = \text{initial} - \text{reacted} = 0.0025 - \text{moles of NaOH added} $$

- Moles of sodium acetate formed:

$$ \text{NaOAc} = \text{moles of NaOH added} $$

Step 2: Calculate Concentrations of Acid and Conjugate Base
Total volume after addition:

$$ V_{total} = V_{acid} + V_{NaOH} $$

Concentrations:

$$ [\text{HA}] = \frac{\text{remaining acid moles}}{V_{total}} $$

$$ [\text{A}^-] = \frac{\text{moles of sodium acetate}}{V_{total}} $$

Step 3: Use the Henderson-Hasselbalch Equation
The pH of a buffer solution:

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

- The pKa of acetic acid:

$$ pK_a = 4.76 $$

Calculate pH accordingly.

Case 2: At the Equivalence Point

When the moles of NaOH added equal the initial moles of acetic acid:

$$ \text{moles NaOH} = 0.0025 \text{ mol} $$

- Total volume:

$$ V_{eq} = V_{acid} + V_{NaOH} $$

- The solution contains only sodium acetate, which hydrolyzes in water:

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

- To find pH, calculate the concentration of acetate and its hydrolysis.

Step 4: Hydrolysis of Sodium Acetate and pH

Calculation

The hydrolysis constant:

$$ K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{1.74 \times 10^{-5}} \approx 5.75 \times 10^{-10} $$

Concentration of acetate:

$$ [\text{A}^-] = \frac{\text{moles of acetate}}{V_{eq}} $$

Set up equilibrium expression:

$$ \text{A}^- + \text{H}_2\text{O} \leftrightarrow \text{HA} + \text{OH}^- $$

Assuming hydrolysis produces x mol/L of OH^-:

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

Calculate pOH:

$$ \text{pOH} = -\log x $$

and then:

$$ \text{pH} = 14 - \text{pOH} $$

Calculations at Various Points During Titration

To comprehensively understand the titration curve, calculations can be performed at different volumes of NaOH added — for example, at 5 mL, 10 mL, 15 mL, and at the equivalence point.

- At 5 mL NaOH added: Less than half neutralized; solution acts as a buffer.
- At 10 mL NaOH added: Near half-equivalence point; pH close to pKa.
- At 15 mL NaOH added: Past the half-equivalence point; pH rises rapidly.
- At 25 mL NaOH added: At the equivalence point; pH determined by hydrolysis.
- Beyond 25 mL: Excess NaOH; pH dictated by remaining hydroxide.

Practical Example: Calculating pH After Adding 10 mL NaOH

Suppose you add 10 mL of NaOH:

- Moles of NaOH:

$$ 0.10 \times 0.010 = 0.001 \text{ mol} $$

- Remaining acetic acid:

$$ 0.0025 - 0.001 = 0.0015 \text{ mol} $$

- Sodium acetate formed:

$$ 0.001 \text{ mol} $$

- Total volume:

$$ 25 \text{ mL} + 10 \text{ mL} = 35 \text{ mL} = 0.035 \text{ L} $$

- Concentrations:

$$ [\text{HA}] = \frac{0.0015}{0.035} \approx 0.0429 \text{ M} $$

$$ [\text{A}^-] = \frac{0.001}{0.035} \approx 0.0286 \text{ M} $$

- Use Henderson-Hasselbalch:

$$ \text{pH} = 4.76 + \log \left(\frac{0.0286}{0.0429} \right) = 4.76 + \log(0.666) $$

$$ \text{pH} \approx 4.76 - 0.177 = 4.58 $$

This indicates that after adding 10 mL of NaOH, the solution has a pH of approximately 4.58.

Conclusion: Key Takeaways for Accurate

Frequently Asked Questions

How do you calculate the pH after titrating 25 mL of 0.10 M acetic acid with sodium hydroxide?

To calculate the pH after titration, determine the moles of acetic acid and NaOH reacted, find the remaining acetic acid or formed acetate

ions, and then use the Henderson-Hasselbalch equation or direct ion concentration calculations based on the titration stage.

What is the initial pH of 25 mL of 0.10 M acetic acid before titration?

Initial pH can be calculated using the acid dissociation constant (Ka) of acetic acid, leading to a pH around 2.88 for this concentration.

At what point during titration does the mixture reach the equivalence point for 25 mL of 0.10 M acetic acid?

The equivalence point occurs after adding 25 mL of NaOH solution with molarity 0.10 M, neutralizing all acetic acid to form acetate ions, where the pH typically rises above 7 due to hydrolysis.

How to determine the pH after adding a certain volume of NaOH during titration?

Calculate the remaining moles of acetic acid and the moles of sodium acetate formed, then apply the Henderson-Hasselbalch equation or compute the hydrolysis of acetate to find the pH.

What is the pH after adding 10 mL of 0.10 M NaOH to the acetic acid solution?

After adding 10 mL of NaOH, some acetic acid has been neutralized, resulting in a buffer solution. The pH can be calculated using the concentrations of acetic acid and acetate ions in solution, typically around 3.0 to 4.0.

How does the pH change during the titration of acetic acid with NaOH?

Initially, the pH is low (~2.9). As NaOH is added, pH increases gradually, forming a buffer region, then rising sharply near the equivalence point, and finally leveling off in the basic range after complete neutralization.

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

At the half-equivalence point, half of the acetic acid is neutralized, and pH equals the pKa of acetic acid, approximately 4.76.

How do you calculate the pH at the equivalence point in this titration?

At the equivalence point, all acetic acid is converted to acetate ions, and the pH is determined by hydrolysis of the acetate ion, typically around 8.7.

What role does the acetic acid's Ka play in calculating pH during titration?

Ka helps determine the degree of ionization of acetic acid at various stages, enabling calculation of hydrogen ion concentration and pH during the titration process.

Can you provide a step-by-step method to calculate pH after adding any volume of NaOH during the titration?

Yes. First, calculate moles of acetic acid and NaOH added. Determine remaining acetic acid or formed acetate. Use the Henderson-Hasselbalch equation for buffer regions or hydrolysis equations near the equivalence point to find pH.

Additional Resources

pH Calculation in Titration of Acetic Acid with Sodium Hydroxide: An Expert Guide

Titration is a fundamental laboratory technique used to determine the concentration of an unknown solution by reacting it with a solution of known concentration. Among various titrations, the acid-base titration involving acetic acid (a weak acid) and sodium hydroxide (a strong base) is particularly educational, offering insights into concepts like equivalence points, buffer regions, and pH calculations. This article provides a comprehensive guide on how to calculate the pH during the titration of 25 mL of 0.10 M acetic acid (CH_3COOH) with sodium hydroxide (NaOH), focusing on the process after each addition, especially after the equivalence point.

Understanding the Fundamentals of the Titration Process

Before delving into pH calculations, it is essential to understand the underlying chemistry of the titration process between acetic acid and sodium hydroxide.

Reaction Overview

The primary chemical reaction is:

$$\mathrm{CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O}$$

- Acetic acid is a weak acid, meaning it does not dissociate completely in solution.
- Sodium hydroxide is a strong base, dissociating fully in solution.
- The titration involves gradually adding NaOH to the acetic acid solution until the acid is neutralized.

Key Concepts in Titration

- Initial pH: The pH of the acetic acid solution before any base is added.
- Equivalence Point: The point where moles of NaOH added equal the moles of acetic acid originally present.
- Half-Equivalent Point: The point where half

of the acetic acid has been neutralized; the pH at this point equals the pKa of acetic acid.
- Beyond Equivalence Point: When excess NaOH is present, the solution's pH is dominated by the excess base.

Calculating Initial pH of the Acetic Acid Solution

The initial pH is determined by the concentration of acetic acid and its acid dissociation constant (K_a).

Data Summary

- Volume of acetic acid, $V_{acid} = 25\, \text{mL} = 0.025\, \text{L}$
- Concentration of acetic acid, $C_{acid} = 0.10\, \text{M}$
- Moles of acetic acid, $n_{acid} = C_{acid} \times V_{acid} = 0.10\, \text{mol/L} \times 0.025\, \text{L} = 2.5 \times 10^{-3} \, \text{mol}$

The dissociation of acetic acid is:

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

with $K_a = 1.8 \times 10^{-5}$.

Calculating the Initial pH

Assuming x is the concentration of H^+:

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

Since acetic acid is weak and initial concentration is known:

$$ K_a = \frac{x^2}{C_{acid} - x} \approx \frac{x^2}{C_{acid}} \quad \text{(because } x \ll C_{acid} \text{)} $$

$$ x = \sqrt{K_a \times C_{acid}} = \sqrt{1.8 \times 10^{-5} \times 0.10} $$

$$ x = \sqrt{1.8 \times 10^{-6}} \approx 1.34 \times 10^{-3}\, \text{M} $$

Thus,

$$pH_{initial} = -\log [H^+] = -\log (1.34 \times 10^{-3}) \approx 2.87$$

Initial pH ≈ 2.87

Progressing Through the Titration: Step-by-Step pH Calculations

The pH at any point during titration depends on the amount of NaOH added relative to the initial amount of acetic acid. We analyze key points:

- Before equivalence point
- At equivalence point
- After equivalence point

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1. Before the Equivalence Point

This is when some NaOH has been added, but not enough to neutralize all acetic acid.

Scenario: Suppose 10 mL of NaOH (0.10 M) is added

- Volume of NaOH added: $V_{NaOH} = 10\, \text{mL} = 0.010\, \text{L}$
- Moles of NaOH added:

$$n_{NaOH} = 0.10\, \text{mol/L} \times 0.010\, \text{L} = 1.0 \times 10^{-3}\, \text{mol}$$

- Moles of acetic acid initially: $2.5 \times 10^{-3}\, \text{mol}$

Remaining acetic acid:

$$n_{acid, remaining} = 2.5 \times 10^{-3} - 1.0 \times 10^{-3} = 1.5 \times 10^{-3}\, \text{mol}$$

- Moles of acetate formed:

$$n_{acetate} = 1.0 \times 10^{-3}\, \text{mol}$$

Total volume of solution:

$$V_{total} = 25\, \text{mL} + 10\, \text{mL} = 35\, \text{mL} = 0.035\, \text{L}$$

- Concentration of remaining acetic acid:

$$ [CH_3COOH] = \frac{1.5 \times 10^{-3}}{0.035} \approx 0.0429\, \text{M} $$

- Concentration of acetate:

$$ [CH_3COO^-] = \frac{1.0 \times 10^{-3}}{0.035} \approx 0.0286\, \text{M} $$

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pH Calculation using Buffer Equation (Henderson-Hasselbalch):

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

where

- $pK_a = -\log K_a = -\log 1.8 \times 10^{-5} \approx 4.74$

Plugging in:

$$ pH = 4.74 + \log \left(\frac{0.0286}{0.0429}\right) $$
$$ pH = 4.74 + \log (0.666) $$
$$ pH = 4.74 - 0.177 $$
$$ pH \approx 4.56 $$

pH after adding 10 mL NaOH ≈ 4.56

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2. At the Equivalence Point

The equivalence point occurs when:

$$ n_{NaOH} = n_{acid} = 2.5 \times 10^{-3}\, \text{mol} $$

Calculating the volume of NaOH needed:

$$ V_{eq} = \frac{n_{acid}}{C_{NaOH}} = \frac{2.5 \times 10^{-3}}{0.10} = 0.025\, \text{L} = 25\, \text{mL} $$

At this point:

- All acetic acid has been neutralized to form sodium acetate.
- The solution contains a conjugate base (acetate ion), which hydrolyzes in water:

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

This hydrolysis determines the pH.

Calculating pH at the Equivalence Point

- Moles of acetate:

\[n_{acetate} = 2.5 \times 10^{-3}\, \text{mol} \]

- Volume of solution:

\[V_{total} = 25\, \text{mL} + 25\, \text{mL} = 50\, \text{mL} = 0.05\, \text{L} \]

- Concentration of acetate:

\[[\mathrm{CH_3COO}^-] = \frac{2.5 \times 10^{-3}}{0.05} = 0.05\, \text{M} \]

- Hydrolysis constant:

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

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