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Introduction to Acid-Base Titration and Acetic Acid

Titration is a fundamental laboratory technique used to determine the concentration of an unknown solution by reacting it with a solution of known concentration. When dealing with weak acids like acetic acid (CH_3COOH), understanding the properties such as the acid dissociation constant (Ka) becomes essential to analyze the titration process accurately.

In this article, we explore the titration of 50.0 mL of 0.100 M acetic acid with 0.200 M sodium hydroxide (NaOH), focusing on calculating the pH at various stages, determining the equivalence point, and understanding the significance of the acid dissociation constant (Ka) of acetic acid. This comprehensive analysis provides insights valuable for chemistry students, educators, and laboratory practitioners interested in acid-base equilibria.

Understanding Acetic Acid and Its Properties

What is Acetic Acid?

Acetic acid (CH_3COOH) is a weak, monoprotic acid commonly found in vinegar, where it imparts the sour taste and pungent smell. Its chemical structure consists of a methyl group attached to a carboxyl group, which is responsible for its acidic properties.

Why is Ka Important?

The acid dissociation constant, Ka, quantifies the strength of an acid in solution. It describes the equilibrium between the undissociated acid and its ions:

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

The Ka value influences the pH of the solution and the titration curve's shape. For acetic acid, Ka is approximately 1.8×10^{-5}, indicating it is a weak acid that does not fully dissociate in water.

Details of the Titration Process

Initial Conditions

- Volume of acetic acid solution: 50.0 mL
- Concentration of acetic acid: 0.100 M
- Concentration of NaOH: 0.200 M

The initial moles of acetic acid:

$$
n_{acid} = C_{acid} \times V_{acid} = 0.100\, \text{mol/L} \times 0.050\, \text{L} = 0.005\, \text{mol}
$$

The initial moles of NaOH added depend on the volume titrated, but at this stage, none has been added.

Calculating the Volume of NaOH Required to Reach the Equivalence Point

Since acetic acid is monoprotic, each mole reacts with one mole of NaOH:

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

The volume of NaOH needed to neutralize all acetic acid:

$$
V_{NaOH} = \frac{n_{acid}}{C_{NaOH}} = \frac{0.005\, \text{mol}}{0.200\, \text{mol/L}} = 0.025\, \text{L} = 25.0\, \text{mL}
$$

Thus, at 25.0 mL of NaOH added, the titration reaches the equivalence point.

Understanding the Titration Curve and pH Calculations

The titration curve of acetic acid with NaOH exhibits characteristic features: initial pH, gradual increase during addition, a sharp rise at the equivalence point, and a final pH determined by excess base.

Initial pH of the Acetic Acid Solution

Using the Ka expression:

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

Assuming initial dissociation is small, the initial concentration of $[\text{H}^+]$:

$$
[\text{H}^+] = \sqrt{K_a \times C_{acid}} = \sqrt{1.8 \times 10^{-5} \times 0.100} \approx 1.34 \times 10^{-3}\, \text{M}
$$

Calculating initial pH:

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

pH at Various Points During Titration

- Before the Equivalence Point: The pH is determined by the remaining acetic acid and the hydrolysis of its conjugate base.
- At Half-Equivalence Point: The pH equals the pKa (by Henderson-Hasselbalch equation).
- At the Equivalence Point: The solution contains only sodium acetate, a weak base, so the pH is above 7.
- Beyond Equivalence: Excess NaOH dominates, and pH increases sharply.

Calculating pH at the Half-Equivalence Point

At half-weak acid titration, half of the acetic acid has been neutralized:

$$
V_{half} = \frac{V_{equivalence}}{2} = \frac{25\, \text{mL}}{2} = 12.5\, \text{mL}
$$

Number of moles of acetic acid remaining:

$$
n_{remaining} = n_{acid} - n_{NaOH} = 0.005\, \text{mol} - (0.200\, \text{mol/L} \times 0.0125\, \text{L}) = 0.005\, \text{mol} - 0.0025\, \text{mol} = 0.0025\, \text{mol}
$$

Concentration of remaining acetic acid:

$$
C_{remaining} = \frac{0.0025\, \text{mol}}{0.0125\, \text{L}} = 0.200\, \text{M}
$$

But more straightforwardly, at half-equivalence, the pH equals pKa:

$$
pH = pK_a = -\log (1.8 \times 10^{-5}) \approx 4.74
$$

Determining the pH at the Equivalence Point

At the equivalence point, all acetic acid has been converted to sodium acetate (CH_3COONa), which is a weak base. Its hydrolysis affects the pH:

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

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 the equivalence point:

$$
C_{acetate} = \frac{n_{acid}}{V_{total}} = \frac{0.005\, \text{mol}}{0.050\, \text{L} + 0.025\, \text{L}} = \frac{0.005}{0.075} \approx 0.0667\, \text{M}
$$

Hydrolysis of acetate:

$$
K_b = \frac{[\text{OH}^-]^2}{C_{acetate}}
$$

$$
[\text{OH}^-] = \sqrt{K_b \times C_{acetate}} = \sqrt{5.56 \times 10^{-10} \times 0.0667} \approx 6.09 \times 10^{-6}\, \text{M}
$$

pOH:

$$
pOH = -\log [\text{OH}^-] \approx 5.22
$$

pH:

$$
pH = 14 - pOH \approx 8.78
$$

Thus, at the equivalence point, the solution is slightly basic with pH around 8.78.

Significance of the Ka of Acetic Acid in Titration

Understanding the Ka of acetic acid is fundamental in predicting the behavior of the titration. It influences:

- The initial pH of the acetic acid solution.
- The pH at the half-equivalence point (which equals pKa).

- The shape of the titration curve, especially the buffer region.
- The pH at the equivalence point, which is

Frequently Asked Questions

What is the initial moles of acetic acid in 50.0 mL of 0.100 M solution?

The initial moles of acetic acid are calculated as 0.100 mol/L × 0.050 L = 0.005 mol.

How many moles of NaOH are needed to completely neutralize 50.0 mL of 0.100 M acetic acid?

Since acetic acid is a weak acid, it reacts 1:1 with NaOH; thus, 0.005 mol of NaOH are needed, which corresponds to 0.005 mol / 0.200 mol/L = 0.025 L or 25 mL.

What is the pH of the solution after adding 25 mL of NaOH to the acetic acid?

At this point, the acid is completely neutralized, and the solution contains only water and sodium acetate, which forms a basic solution. The pH can be calculated from the hydrolysis of acetate ions, approximately pH ≈ 8.7.

How can the Ka of acetic acid be determined from the titration data?

By analyzing the pH at the half-equivalence point, where moles of acid equal moles of base, the pKa can be found as pKa = pH at half-equivalence, which then gives Ka = 10^(-pKa).

What is the significance of the equivalence point in this titration?

The equivalence point is where the amount of NaOH added equals the initial amount of acetic acid, resulting in complete neutralization and a sharp change in pH, typically around pH 8.7 for acetic acid titrations.

How does the Ka of acetic acid influence the shape of the titration curve?

The Ka determines the acid strength; a smaller Ka results in a higher pH at the equivalence point and a more gradual titration curve, while a larger Ka results in a steeper curve.

What is the expected pH at the start of the titration before any NaOH is added?

The initial pH can be calculated from the acetic acid concentration using the expression for weak acids: pH ≈ 2.87.

At the half-equivalence point, what is the relationship between pH and pKa?

At half-equivalence, pH equals pKa because the concentrations of acetic acid and acetate are equal, allowing direct determination of Ka from the pH.

Why is understanding the Ka of acetic acid important in titration analysis?

Knowing the Ka helps predict the pH behavior during titration, interpret the titration curve, and accurately determine the concentration of unknown acids or bases.

Additional Resources

50.0 mL of 0.100 M Acetic Acid (CH_3COOH (aq)) is titrated with 0.200 M NaOH (aq). The Ka of acetic acid — this scenario offers a classic example of acid-base titration, illustrating key concepts in chemistry such as molarity, stoichiometry, pH calculations, and acid dissociation constants. Whether you're a student preparing for exams or a professional seeking a refresher, understanding this titration process provides foundational insight into how acids and bases interact in aqueous solutions. This guide will walk you through the detailed steps involved in analyzing this titration, from initial setup to calculating the pH at various points, including the equivalence point, and understanding the significance of the acid dissociation constant (Ka).

Understanding the Basics: Acetic Acid and Sodium Hydroxide

Before diving into calculations, it's essential to understand the properties of the substances involved:

- Acetic Acid (CH_3COOH): A weak acid with a known Ka of approximately 1.8×10^{-5} at 25°C. Its weak acid nature means it does not fully dissociate in water, establishing an equilibrium between undissociated acetic acid and its ions.
- Sodium Hydroxide (NaOH): A strong base that dissociates completely in water, providing hydroxide ions (OH^-) readily.

The Titration Setup

Initial concentrations and volumes:

- Acetic Acid: 50.0 mL of 0.100 M solution
- Sodium Hydroxide: 0.200 M solution

Key concept: The titration involves gradually adding NaOH to the acetic acid solution until the acid is neutralized. The goal is to analyze the pH at various stages, especially at the equivalence point where the acid is completely neutralized.

Step 1: Calculating Moles of Acetic Acid

Determine the initial moles of acetic acid:

$$
\text{Moles of acetic acid} = \text{Molarity} \times \text{Volume (in liters)} = 0.100\, \text{mol/L} \times 0.050\, \text{L} = 0.005\, \text{mol}
$$

Step 2: Determining the Volume of NaOH Needed to Reach Equivalence Point

Since NaOH is a strong base and acetic acid is a weak acid, the reaction can be written as:

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

The molar ratio is 1:1. Therefore,

$$
\text{Volume of NaOH} = \frac{\text{Moles of acetic acid}}{\text{Concentration of NaOH}} = \frac{0.005\, \text{mol}}{0.200\, \text{mol/L}} = 0.025\, \text{L} = 25.0\, \text{mL}
$$

This means adding 25.0 mL of 0.200 M NaOH will neutralize all acetic acid, reaching the equivalence point.

Step 3: Titration Curve and pH Calculations at Key Points

A titration curve plots pH against the volume of titrant added. Critical points include:

- Initial pH (before titrant addition)
- Half-equivalence point
- Equivalence point
- Post-equivalence point

Each of these points offers insights into the solution's properties.

Initial pH of the Acetic Acid Solution

Calculate the initial pH of the 0.100 M acetic acid:

Since acetic acid is a weak acid, we use the acid dissociation constant (Ka):

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

The dissociation equilibrium:

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

Assuming initial concentration:

- Initial concentration of acetic acid: 0.100 M
- Let x be the concentration of $\mathrm{H^+}$ at equilibrium.

Set up the expression:

$$
K_a = \frac{x^2}{0.100 - x} \approx \frac{x^2}{0.100}
$$

(assuming x is small relative to 0.100 M)

Solve for x:

$$
x^2 = K_a \times 0.100 = 1.8 \times 10^{-5} \times 0.100 = 1.8 \times 10^{-6}
$$

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

\]

Calculate initial pH:

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

Step 4: pH at the Half-Equivalence Point

The half-equivalence point occurs when half the acid has been neutralized. At this stage:

- Moles of acetic acid remaining $= 0.0025\, \text{mol}$
- Moles of acetate $= 0.0025\, \text{mol}$

Key property: At the half-equivalence point, $\mathrm{pH} = \mathrm{p}K_a$

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

Thus,

$$
\boxed{\mathrm{pH} \approx 4.74}
$$

This is a crucial point in the titration curve, representing the buffer region where the solution contains significant amounts of both acetic acid and acetate ions.

Step 5: pH at the Equivalence Point

At the equivalence point, all acetic acid has been converted into acetate ions ($\mathrm{CH_3COO^-}$). The solution now contains the conjugate base of acetic acid, which affects the pH due to hydrolysis.

Hydrolysis of acetate:

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

The equilibrium constant for this base hydrolysis:

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

Calculate the hydroxide ion concentration:

- Moles of acetate at equivalence: 0.005 mol
- Total volume: initial 50.0 mL + added 25.0 mL = 75.0 mL (0.075 L)

Concentration of acetate:

\[

$$
[\mathrm{CH_3COO^-}] = \frac{0.005\, \text{mol}}{0.075\, \text{L}} \approx 0.0667\, \text{M}
$$

Set up hydrolysis equilibrium:

$$
K_b = \frac{[\mathrm{OH^-}]^2}{[\mathrm{CH_3COO^-}]}
$$

Let x be the concentration of $\mathrm{OH^-}$ produced:

$$
x^2 = K_b \times [\mathrm{CH_3COO^-}] = 5.56 \times 10^{-10} \times 0.0667 \approx 3.71 \times 10^{-11}
$$

$$
x = \sqrt{3.71 \times 10^{-11}} \approx 6.09 \times 10^{-6}\, \text{M}
$$

Calculate pOH:

$$
\mathrm{pOH} = -\log x \approx -\log (6.09 \times 10^{-6}) \approx 5.22
$$

Calculate pH:

$$
\mathrm{pH} = 14 - \mathrm{pOH} \approx 14 - 5.22 = 8.78
$$

Thus, the pH at the equivalence point is approximately 8.78, indicating a slightly basic solution due to the hydrolysis of acetate ions.

Step 6: Post-Equivalence pH

Adding excess NaOH increases the pH further:

- For example, adding 30 mL of NaOH (extra 5 mL beyond the equivalence point):

Total volume: 80 mL (0.080 L)

Moles of NaOH added: $0.200\, \text{mol/L} \times 0.030\, \text{L} = 0.006\, \text{mol}$

Moles of excess NaOH:

$$
0.006\, \text{mol} - 0.005\, \text{mol} = 0.001\, \text{mol}
$$

50 0 Ml Of 0 100 M Acetic Acid Ch3cooh Aq Is Titrated With 0

200 M Naoh Aq The Ka Of Acetic Acid

50 0 Ml Of 0 100 M Acetic Acid Ch3cooh Aq Is Titrated With 0 200 M Naoh Aq The Ka Of Acetic Acid

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