If 37.5 Ml Of 0.100 M Naoh Is Added To 10.0 Ml Of 0.100 M Ch3cooh, What Will Be The Ph Of The Resulting

Introduction

If 37.5 mL of 0.100 M NaOH is added to 10.0 mL of 0.100 M CH3COOH, what will be the pH of the resulting solution? This question involves understanding acid-base titration principles, molar calculations, and the concept of the pH of buffer solutions. To determine the final pH, we need to analyze the reaction between acetic acid (CH3COOH) and sodium hydroxide (NaOH), calculate the moles involved, and assess whether the solution is acidic, basic, or neutral after the reaction. This comprehensive analysis will provide insights into the chemistry of weak acids and strong bases, the concept of equivalence points, and how to calculate pH in mixed solutions.

Understanding the Components of the Problem

Acetic Acid (CH3COOH)

- It is a weak acid with a molarity of 0.100 M.
- Volume of the acetic acid solution is 10.0 mL.
- Molar mass of acetic acid is approximately 60.05 g/mol.
- Its acid dissociation constant (Ka) is about 1.8 × 10-5.

Sodium Hydroxide (NaOH)

- It is a strong base with a molarity of 0.100 M.
- Volume of NaOH solution added is 37.5 mL.
- Molar mass of NaOH is approximately 40.00 g/mol.

Key Concepts Needed

- Moles of reactants determined from molarity and volume.
- Reaction between acetic acid and NaOH: neutralization process.
- Equivalence point, excess reactants, and buffer formation.
- Calculation of pH based on the remaining species after reaction.

Step 1: Calculating Moles of Reactants

Calculate moles of acetic acid (CH_3COOH)

- Moles = Molarity × Volume (in liters)
- Volume in liters = 10.0 mL = 0.010 L
- Moles of CH_3COOH = 0.100 mol/L × 0.010 L = 0.001 mol

Calculate moles of NaOH added

- Volume in liters = 37.5 mL = 0.0375 L
- Moles of NaOH = 0.100 mol/L × 0.0375 L = 0.00375 mol

Step 2: Analyzing the Reaction

Reaction Equation

- $NaOH + CH_3COOH \rightarrow CH_3COONa + H_2O$

This reaction shows a 1:1 molar ratio between NaOH and acetic acid.

Determining the Extent of Reaction

- Moles of NaOH (0.00375 mol) is greater than moles of acetic acid (0.001 mol).
- All acetic acid will be neutralized, and excess NaOH will remain in solution.

Calculating Remaining NaOH

- Moles of NaOH used to neutralize acetic acid: 0.001 mol
- Remaining NaOH: 0.00375 mol - 0.001 mol = 0.00275 mol

Step 3: Final Composition of the Solution

After Reaction

- All acetic acid has been reacted; none remains.
- Excess NaOH remains in solution at 0.00275 mol.

Final Volume of the Solution

- Total volume = volume of acetic acid + volume of NaOH added
- 10.0 mL + 37.5 mL = 47.5 mL = 0.0475 L

Concentration of Remaining NaOH

- Concentration = moles / volume in liters = 0.00275 mol / 0.0475 L ≈ 0.0579 M

Step 4: Calculating the pH of the Final Solution

Understanding the pH in the Presence of Excess NaOH

- Since excess NaOH is present, the solution is basic.
- pH determined by the concentration of OH- ions from remaining NaOH.

Calculating pOH

- pOH = –log[OH-]
- [OH-] ≈ 0.0579 M
- pOH ≈ –log(0.0579) ≈ 1.24

Calculating pH

- pH = 14 – pOH ≈ 14 – 1.24 ≈ 12.76

Alternative Scenario: Exact Equivalence Point

If instead, the amount of NaOH added were exactly enough to neutralize all acetic acid, the solution would be at the equivalence point, resulting in a neutral solution with pH ≈ 7.0.

However, in this specific case, since excess NaOH remains, the pH is significantly basic, approximately 12.76, confirming the basic nature of the solution due to unreacted NaOH.

Summary of Key Points

- Initial moles of acetic acid = 0.001 mol
- Initial moles of NaOH added = 0.00375 mol

- NaOH excess after neutralization = 0.00275 mol
- Total solution volume = 47.5 mL
- Concentration of leftover NaOH ≈ 0.0579 M
- pH of resulting solution ≈ 12.76

Conclusion

Adding 37.5 mL of 0.100 M NaOH to 10.0 mL of 0.100 M acetic acid results in a solution with a high pH due to the excess strong base. The calculated pH of approximately 12.76 indicates a strongly basic solution. This analysis underscores the importance of molar calculations and understanding acid-base reactions to predict the pH of complex mixtures accurately.

This exercise exemplifies fundamental concepts in titration and buffer chemistry, illustrating how the amount and concentration of reactants influence the final pH, a critical aspect in many laboratory and industrial processes involving acids and bases.

Frequently Asked Questions

What is the initial concentration and volume of NaOH added to acetic acid in this reaction?

The initial concentration of NaOH is 0.100 M, and the volume added is 37.5 mL.

What is the initial concentration and volume of acetic acid in the solution?

The initial concentration of acetic acid (CH3COOH) is 0.100 M, and the volume is 10.0 mL.

How do you determine the moles of NaOH added to the solution?

Calculate moles of NaOH by multiplying its concentration by volume in liters: 0.100 mol/L × 0.0375 L = 0.00375 mol.

How many moles of acetic acid are initially present in the solution?

Calculate moles of acetic acid: 0.100 mol/L × 0.010 L = 0.001 mol.

What is the net reaction between NaOH and acetic acid in this problem?

NaOH reacts with acetic acid to form sodium acetate and water: $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$.

Will all the acetic acid be neutralized by the added NaOH?

No, since the moles of NaOH (0.00375 mol) are greater than the moles of acetic acid (0.001 mol), the acid will be fully neutralized with excess NaOH remaining.

How do you determine the moles of excess NaOH after neutralization?

Subtract the moles of acetic acid from the moles of NaOH: 0.00375 mol - 0.001 mol = 0.00275 mol of NaOH remaining.

What is the pH of the resulting solution after all acetic acid is neutralized?

The solution contains excess NaOH, so the pH is determined by the leftover base: $pOH = -\log[OH^-]$, and $pH = 14 - pOH$.

How do you calculate the final pH of the solution considering the excess NaOH?

Calculate $[OH^-]$ = 0.00275 mol / total volume in liters (10.0 mL + 37.5 mL = 47.5 mL = 0.0475 L): $[OH^-] \approx$ 0.00275 mol / 0.0475 L $\approx$ 0.0579 M. Then, $pOH \approx 4.24$, so $pH \approx 14 - 4.24 \approx 9.76$.

What is the approximate pH of the resulting solution after mixing 37.5 mL of 0.100 M NaOH with 10.0 mL of 0.100 M acetic acid?

The pH is approximately 9.76, indicating a slightly basic solution due to excess NaOH remaining after neutralization.

Additional Resources

If 37.5 mL of 0.100 M NaOH Is Added to 10.0 mL of 0.100 M CH3COOH, What Will Be the pH of the Resulting Solution?

Introduction

The interaction between a weak acid and a strong base is a fundamental concept in acid-base chemistry, with wide-ranging applications from titration procedures to biological systems. This investigation delves into a specific scenario: what is the resulting pH when 37.5 mL of 0.100 M sodium hydroxide (NaOH) is added to 10.0 mL of 0.100 M acetic acid (CH_3COOH)? The answer to this question involves understanding the principles of neutralization reactions, molar calculations, and the properties of weak acids and strong bases.

Background: Acid-Base Chemistry Fundamentals

Acetic acid (CH_3COOH) is a weak acid with a known dissociation constant (Ka) of approximately 1.8×10^{-5}. It partially dissociates in aqueous solution:

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

Sodium hydroxide (NaOH) is a strong base that dissociates completely:

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

When these two substances are mixed, the hydroxide ions (OH^-) from NaOH neutralize the hydrogen ions (H^+) from acetic acid, resulting in the formation of acetate ions (CH_3COO^-) and water.

Step 1: Establishing the Initial Moles

Calculating initial moles of acetic acid (CH_3COOH):
- Volume = 10.0 mL = 0.0100 L
- Concentration = 0.100 M

$$\text{Moles of } \mathrm{CH_3COOH} = 0.100\, \mathrm{mol/L} \times 0.0100\, \mathrm{L} = 1.00 \times 10^{-3}\, \mathrm{mol}$$

Calculating initial moles of NaOH:
- Volume = 37.5 mL = 0.0375 L
- Concentration = 0.100 M

$$\text{Moles of } \mathrm{NaOH} = 0.100\, \mathrm{mol/L} \times 0.0375\, \mathrm{L} = 3.75 \times 10^{-3}\, \mathrm{mol}$$

Step 2: Reaction Stoichiometry and Extent

Since NaOH is a strong base, it will react fully with acetic acid in a 1:1 molar ratio:

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

Determining the limiting reagent:
- Moles of acetic acid = 1.00×10^{-3} mol
- Moles of NaOH = 3.75×10^{-3} mol

Because the amount of NaOH exceeds that needed to neutralize all acetic acid, the acid will be completely neutralized, and excess NaOH will remain.

Calculating remaining NaOH:

$$
\text{Excess } \mathrm{OH^-} = 3.75 \times 10^{-3} - 1.00 \times 10^{-3} = 2.75 \times 10^{-3}\, \mathrm{mol}
$$

Remaining NaOH after neutralization:
- Total volume after mixing:

$$
V_{total} = 10.0\, \mathrm{mL} + 37.5\, \mathrm{mL} = 47.5\, \mathrm{mL} = 0.0475\, \mathrm{L}
$$

- Concentration of excess NaOH:

$$
[\mathrm{OH^-}] = \frac{2.75 \times 10^{-3}\, \mathrm{mol}}{0.0475\, \mathrm{L}} \approx 0.0579\, \mathrm{M}
$$

Step 3: Determining the Final pH

Since all acetic acid is neutralized, the solution now contains only water, acetate ions (from the neutralized acid), and excess NaOH.

Key points:
- The solution is basic due to residual NaOH.
- The pH is governed by the concentration of OH^- ions.

Calculating pOH:

\[

$$
\mathrm{pOH} = -\log [\mathrm{OH^-}] = -\log (0.0579) \approx 1.24
$$

Calculating pH:

$$
\mathrm{pH} = 14 - \mathrm{pOH} = 14 - 1.24 = 12.76
$$

Therefore, the pH of the resulting solution is approximately 12.76.

Deep Dive: Influence of Acetic Acid and Its Conjugate Base

While the straightforward calculation shows a high pH dominated by excess NaOH, a more thorough analysis considers the presence of the acetate ions and their potential to hydrolyze:

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

However, since all acetic acid is neutralized, the acetate ions are not in excess; the excess hydroxide ions primarily determine the pH.

Note: If the initial amounts were more comparable or if the acid was in excess, the resulting pH would be closer to the neutral or slightly acidic range, necessitating application of the Henderson-Hasselbalch equation.

Factors Affecting the Resulting pH

Several factors could influence the outcome:

1. Concentration and Volume Accuracy: Minor deviations in measurements can significantly affect molar calculations.
2. Temperature: pKa of acetic acid and ionization equilibria are temperature-dependent.
3. Purity of Reagents: Impurities could alter the acid-base balance.
4. Reaction Completeness: Under ideal conditions, the reactions go to completion; in practice, kinetic factors may play a role.

Practical Implications and Applications

Understanding the pH change resulting from adding a strong base to a weak acid is crucial in:

- Titration procedures: Determining endpoint pH values.
- Buffer design: Engineering solutions with specific pH ranges.
- Biological systems: Enzyme activity and cellular homeostasis depend on pH regulation.
- Environmental chemistry: Managing acidity in natural waters.

Summary of Key Findings

- The initial moles of acetic acid and NaOH were calculated based on given concentrations and volumes.
- Excess NaOH remained after complete neutralization of acetic acid.
- The resulting solution's pH was determined primarily by the residual hydroxide ion concentration.
- The final pH of the mixture is approximately 12.76, indicating a strongly basic solution.

Conclusion

The analytical approach to this problem underscores the importance of molar calculations, reaction stoichiometry, and acid-base principles in predicting solution pH. In this specific scenario, the addition of 37.5 mL of 0.100 M NaOH to 10.0 mL of 0.100 M acetic acid results in a strongly basic solution with an approximate pH of 12.76, primarily due to excess hydroxide ions. Such insights are vital for chemists involved in titrations, buffer design, and environmental chemistry, providing a clear example of how quantitative analysis informs qualitative understanding in acid-base reactions.

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