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What Is The pH Of A Solution Prepared By Mixing 25.00 ML Of 0.10 M CH_3CO_2H With 25.00 ML Of 0.010 M NaOH?

Understanding the pH of a solution resulting from mixing a weak acid with a base is fundamental in chemistry, particularly in acid-base titrations, buffer solution preparation, and various industrial applications. In this article, we will explore how to determine the pH of a solution formed when 25.00 mL of 0.10 M acetic acid (CH_3CO_2H) is mixed with 25.00 mL of 0.010 M sodium hydroxide (NaOH). We will analyze the chemical reactions involved, calculate the moles of reactants, identify the limiting reagent, and ultimately compute the pH of the resulting solution.

Understanding the Components of the Mixture

Acetic Acid (CH_3CO_2H)

- Weak acid
- Concentration: 0.10 M
- Volume: 25.00 mL

Sodium Hydroxide (NaOH)

- Strong base
- Concentration: 0.010 M
- Volume: 25.00 mL

This mixture involves a weak acid and a strong base, which typically results in a solution with a pH above 7 after the neutralization process, depending on the extent of reaction.

Calculating the Moles of Reactants

Determine moles of acetic acid (CH_3CO_2H)

Using the molarity and volume:

$$\text{moles of } CH_3CO_2H = M \times V$$

where:
- $M = 0.10\, \text{mol/L}$
- $V = 25.00\, \text{mL} = 0.02500\, \text{L}$

$$\text{moles of } CH_3CO_2H = 0.10\, \text{mol/L} \times 0.02500\, \text{L} = 0.00250\, \text{mol}$$

Determine moles of sodium hydroxide (NaOH)

Similarly:
- $M = 0.010\, \text{mol/L}$
- $V = 0.02500\, \text{L}$

$$
\text{moles of NaOH} = 0.010\, \text{mol/L} \times 0.02500\, \text{L} = 0.00025\, \text{mol}
$$

Reaction Between Acetic Acid and NaOH

Reaction Equation

$$
\text{CH}_3\text{CO}_2\text{H} + \text{NaOH} \rightarrow \text{CH}_3\text{CO}_2\text{Na} + \text{H}_2\text{O}
$$

Since NaOH is a strong base, it will completely react with acetic acid up to the limit of the available NaOH moles.

Determining the Limiting Reagent

- Moles of acetic acid: 0.00250 mol
- Moles of NaOH: 0.00025 mol

Because NaOH is in a smaller amount (limiting reagent), it will be completely consumed, reacting with an equivalent amount of acetic acid.

Extent of Reaction

- NaOH reacts with an equal molar amount of acetic acid: 0.00025 mol
- Remaining acetic acid: $0.00250\, \text{mol} - 0.00025\, \text{mol} = 0.00225\, \text{mol}$

The reaction consumes all the NaOH, producing the corresponding amount of sodium acetate (CH_3CO_2Na) and water.

Determining the Final Composition of the Solution

Remaining Acetic Acid

- Moles: 0.00225 mol
- Volume: Total volume of the mixture = 25.00 mL + 25.00 mL = 50.00 mL = 0.05000 L

$$
\text{Concentration of remaining acetic acid} = \frac{0.00225\, \text{mol}}{0.05000\, \text{L}} = 0.045\, \text{M}
$$

Produced Sodium Acetate (CH_3CO_2Na)

- Moles: Equal to the reacted NaOH, 0.00025 mol
- Concentration:

$$
\text{Concentration of CH}_3\text{CO}_2\text{Na} = \frac{0.00025\, \text{mol}}{0.05000\, \text{L}} = 0.005\, \text{M}
$$

The mixture now contains:
- Unreacted acetic acid at 0.045 M
- Sodium acetate at 0.005 M

This constitutes a buffer solution, which will influence the pH.

Calculating the pH of the Buffer Solution

Using the Henderson-Hasselbalch Equation

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

where:
- pK_a of acetic acid at 25°C is approximately 4.76.
- $[\text{A}^-]$ is the concentration of acetate ion (from sodium acetate).
- $[\text{HA}]$ is the concentration of acetic acid.

Substituting known values:

$$
\text{pH} = 4.76 + \log \left(\frac{0.005}{0.045} \right)
$$

Calculating the ratio:

$$
\frac{0.005}{0.045} \approx 0.1111
$$

Calculating the log:

$$
\log(0.1111) \approx -0.954
$$

Therefore:

$$
\text{pH} = 4.76 - 0.954 = 3.806
$$

Note: The pH appears somewhat acidic, which might seem counterintuitive since the mixture contains a weak acid and a neutralized salt. The key reason is that the remaining acetic acid predominates, and the buffer capacity is limited by the low concentration of acetate.

Final pH of the Mixture

Based on the calculations, the pH of the resulting solution is approximately 3.81.

This indicates the solution remains mildly acidic due to the excess acetic acid, despite the addition of a base, because only a small amount of NaOH was used relative to the initial amount of acetic acid.

Additional Considerations and Summary

Factors Affecting the pH Calculation

- Temperature: The pK_a value varies slightly with temperature; calculations assume 25°C.
- Concentration accuracy: Small deviations in volume or molarity can affect the pH.
- Activity coefficients: For dilute solutions, activity coefficients are close to 1, but in more concentrated solutions, they might influence the pH.

Summary of Key Steps

1. Calculate initial moles of each reactant.

2. Determine the limiting reagent and extent of reaction.

3. Calculate remaining concentrations of weak acid and formed salt.

4. Use the Henderson-Hasselbalch equation to find the pH.

Conclusion

The pH of a solution prepared by mixing 25.00 mL of 0.10 M acetic acid with 25.00 mL of 0.010 M NaOH is approximately 3.81. This value reflects the partial neutralization of acetic acid, resulting in a buffer solution dominated by unreacted acetic acid and some sodium acetate. Understanding these calculations is essential for controlling solution pH in laboratory and industrial processes, emphasizing the importance of stoichiometry, acid-base equilibria, and buffer chemistry in quantitative analysis.

Frequently Asked Questions

What is the pH of a solution prepared by mixing 25.00 mL of 0.10 M acetic acid (CH_3COOH) with 25.00 mL of 0.010 M NaOH?

The pH can be calculated by first determining the moles of acetic acid and NaOH, then finding the remaining acid or base after neutralization, and finally calculating the pH from the resulting concentration. In this case, the solution will be slightly basic with a pH around 4.75.

How does the initial concentration of acetic acid affect the pH when mixed with a base like NaOH?

A higher initial concentration of acetic acid increases the amount of acid available to neutralize the base, resulting in a lower pH after mixing. Conversely, a lower acid concentration leads to a higher pH after neutralization.

Why do we need to consider the volume change when calculating the pH of the mixed solution?

Because the total volume after mixing (the sum of individual volumes) affects the concentrations of the remaining acid or base, which in turn influences the pH calculation.

What is the significance of the molarity of acetic acid and NaOH in determining the final pH?

The molarity indicates the concentration of each component, which helps determine the number of moles reacting and the extent of neutralization, ultimately affecting the final pH of the solution.

How do you calculate the pH after mixing a weak acid with a strong base?

Calculate the moles of acid and base, determine the neutralization amount, find the remaining concentration of the excess component, and then use the appropriate pH formula: for excess acid, $pH = -\log[H^+]$; for excess base, $pOH = -\log[OH^-]$, then convert to pH.

What role does the acid dissociation constant (Ka) play in determining the pH of the solution?

Ka helps determine the degree of dissociation of the weak acid, which influences the concentration of H^+ ions in equilibrium, and thus helps in calculating the pH of the solution.

If the amount of NaOH added exceeds the amount of

acetic acid present, what will be the approximate pH of the solution?

The solution will be basic, with a pH greater than 7, typically around 8–9, depending on the excess amount of NaOH.

How can buffer solution principles be applied to this mixture of acetic acid and NaOH?

Since acetic acid is a weak acid and NaOH is a strong base, the mixture can act as a buffer near the equivalence point, resisting pH changes; however, in this specific case, it mainly depends on the extent of neutralization.

What is the effect of diluting the mixture after neutralization on the pH?

Dilution generally increases the volume, decreasing the concentration of ions, which can cause the pH to approach neutral or slightly shift depending on the remaining ions in solution.

Why is it important to precisely measure the volumes and concentrations when preparing solutions for pH calculations?

Accurate measurements ensure reliable calculations of moles and concentrations, leading to precise pH determination, which is crucial in chemical experiments and applications requiring exact pH control.

Additional Resources

What Is The pH Of A Solution Prepared By Mixing 25.00 mL Of 0.10 M CH_3CO_2H With 25.00 mL Of 0.010 M NaOH

Understanding the pH of a solution formed by mixing acetic acid (CH_3CO_2H) with sodium hydroxide (NaOH) is essential in many areas of chemistry, including buffer solutions, titrations, and industrial processes. This specific scenario involves mixing equal volumes of a weak acid and a dilute strong base, leading to a solution that exhibits partial neutralization. This article provides a comprehensive analysis of the problem, aimed at elucidating the underlying principles, calculations, and implications involved in determining the pH of such a mixture.

Introduction to Acetic Acid and Sodium Hydroxide

Before delving into the calculations, it’s important to understand the fundamentals of the substances involved.

Acetic Acid (CH_3CO_2H)

- Nature: Weak acid
- Concentration: 0.10 M
- Properties: Partially dissociates in water; equilibrium constant (acid dissociation constant, Ka) is approximately 1.8×10^{-5} at 25°C.
- Role in solution: Provides hydrogen ions (H^+) that determine acidity.

Sodium Hydroxide (NaOH)

- Nature: Strong base
- Concentration: 0.010 M
- Properties: Fully dissociates in water, releasing hydroxide ions (OH^-).
- Role in solution: Provides a source of OH^- ions, which neutralize acids.

Understanding the Mixture: Volume, Moles, and Concentrations

When mixing solutions, the key parameters for calculating pH are the moles of acid and base, their initial concentrations, and the total volume after mixing.

Initial Data Summary

- Volume of acetic acid solution: 25.00 mL = 0.02500 L
- Concentration of acetic acid: 0.10 M
- Moles of acetic acid:

$$n_{acid} = C \times V = 0.10\, \text{mol/L} \times 0.02500\, \text{L} = 2.50 \times 10^{-3}\, \text{mol}$$

- Volume of NaOH solution: 25.00 mL = 0.02500 L
- Concentration of NaOH: 0.010 M
- Moles of NaOH:

$$n_{base} = 0.010\, \text{mol/L} \times 0.02500\, \text{L} = 2.50 \times 10^{-4}\, \text{mol}$$

Total volume after mixing:

- $$V_{total} = 0.02500\, \text{L} + 0.02500\, \text{L} = 0.05000\, \text{L}$$

Reaction Between Acetic Acid and Sodium Hydroxide

The core chemical process is a neutralization reaction:

$$\mathrm{CH_3CO_2H} + \mathrm{OH^-} \rightarrow \mathrm{CH_3CO_2^-} + \mathrm{H_2O}$$

- Since NaOH is a strong base, it will react completely with acetic acid until either the acid or base is exhausted.
- The reaction involves the neutralization of acetic acid, producing acetate ions ($CH_3CO_2^-$).

Calculating the Extent of Neutralization

Given the initial moles:
- Moles of acetic acid: 2.50×10^{-3}
- Moles of NaOH: 2.50×10^{-4}

Since NaOH is present in a smaller amount:
- The base will neutralize part of the acetic acid:

$$\text{Moles of acetic acid neutralized} = 2.50 \times 10^{-4}\, \text{mol}$$

Remaining acetic acid:

$$n_{acid, remaining} = 2.50 \times 10^{-3} - 2.50 \times 10^{-4} = 2.25 \times 10^{-3}\, \text{mol}$$

Moles of acetate ions produced:

$$n_{acetate} = 2.50 \times 10^{-4}\, \text{mol}$$

Determining the Main Species in Solution

After the reaction:
- The solution contains:
- Unreacted acetic acid: 2.25×10^{-3} mol
- Acetate ions: 2.50×10^{-4} mol
- Water (from neutralization)

Total volume: 0.05000 L

Concentrations:
- Acetic acid:
$$[\mathrm{CH_3CO_2H}] = \frac{2.25 \times 10^{-3}}{0.05000} = 0.045\, \text{M}$$
- Acetate ion:
$$[\mathrm{CH_3CO_2^-}] = \frac{2.50 \times 10^{-4}}{0.05000} = 0.005\, \text{M}$$

Calculating the pH of the Mixture

Since the solution contains a weak acid (remaining acetic acid) and its conjugate base (acetate), it forms a classic buffer system. The pH can be calculated using the Henderson-Hasselbalch equation:

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

Where:
- pK_a for acetic acid ≈ 4.76
- $[\mathrm{A^-}] = [\mathrm{CH_3CO_2^-}] = 0.005\, \text{M}$
- $[\mathrm{HA}] = [\mathrm{CH_3CO_2H}] = 0.045\, \text{M}$

Plug in the values:

$$\text{pH} = 4.76 + \log \left(\frac{0.005}{0.045} \right)$$
$$\text{pH} = 4.76 + \log (0.1111)$$
$$\text{pH} = 4.76 + (-0.954)$$
$$\text{pH} \approx 3.81$$

Features and Implications of the Result

The calculated pH of approximately 3.81 indicates a slightly acidic solution, which makes sense given the partial neutralization of the weak acid by the small amount of strong base.

Features:
- The solution is a buffer, containing both acetic acid and acetate ions.
- The pH is lower than the pKa, reflecting the dominance of the unneutralized weak acid.
- The presence of acetate ions provides some buffering capacity, resisting drastic pH changes if small additional amounts of acid or base are added.

Pros:
- Buffer capacity: The mixture can resist pH changes upon addition of small quantities of acid or base.
- Predictability: Using Henderson-Hasselbalch provides an accurate pH estimate for such buffer systems.

Cons:
- Limited capacity: Since the base amount is small, the buffer can be overwhelmed if more base is added.
- Assumption of complete dissociation: The calculations assume ideal behavior; in reality, activity coefficients and temperature effects may slightly alter the pH.

Additional Considerations and Variations

While the above calculation provides a good estimate, several factors can influence the actual pH:

- Temperature: pKa values vary with temperature; at higher temperatures, pKa might decrease slightly.
- Activity coefficients: In real solutions, ions do not behave ideally; corrections may be necessary for precise calculations.
- Dilution effects: Further dilutions or additions can shift the pH significantly, especially near the equivalence point.

Conclusion

In summary, the pH of a solution prepared by mixing 25.00 mL of 0.10 M acetic acid with 25.00 mL of 0.010 M NaOH is approximately 3.81. This pH reflects a buffer system formed by the partial neutralization of acetic acid, producing acetate ions. The calculation demonstrates how stoichiometry, equilibrium principles, and buffer chemistry come together to determine the acidity of mixed solutions.

This understanding is not only crucial for academic pursuits but also has practical applications in laboratory titrations, industrial processes, and biological systems where pH control is vital. Recognizing the features and limitations of such calculations enables chemists to design better experiments, interpret results accurately, and develop solutions with desired pH levels.

What Is The Ph Of A Solution Prepared By Mixing 25 00 Ml Of 0 10 M Ch 3co 2h With 25 00 Ml Of 0 010

What Is The Ph Of A Solution Prepared By Mixing 25 00 Ml Of 0 10 M Ch 3co 2h With 25 00 Ml Of 0 010

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