A Buffer Is Prepared By Adding 21.0g Of Sodium Acetate (CH3COONa) To 500mL Of A 0.145M Acetic Acid (CH3COOH)

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Creating a buffer solution is a fundamental technique in chemistry that allows for the maintenance of a stable pH environment when acids or bases are introduced. In this context, preparing a buffer by mixing sodium acetate with acetic acid is a classic example illustrating the principles of acid-base chemistry and buffer solutions. Specifically, adding 21.0 grams of sodium acetate (CH_3COONa) to 500 mL of a 0.145 M acetic acid (CH_3COOH) results in a solution capable of resisting significant pH changes upon the addition of small amounts of acids or bases. This article explores the detailed process of preparing such a buffer, the underlying chemical principles, calculations involved, and practical considerations for laboratory implementation.

Understanding Buffer Solutions and Their Significance

What Is a Buffer Solution?

A buffer solution is a mixture of a weak acid and its conjugate base or a weak base and its conjugate acid. It functions to stabilize the pH of a solution despite the addition of acids or bases. This stability is crucial in many biological, chemical, and industrial processes where maintaining a specific pH range is essential.

Importance of Buffer Solutions

- Biological systems: Enzymatic reactions, blood pH regulation
- Industrial processes: Chemical manufacturing, food preservation
- Laboratory experiments: Ensuring consistent pH during titrations and reactions

Components of the Prepared Buffer

In the scenario involving acetic acid and sodium acetate:
- Weak acid: Acetic acid (CH_3COOH)
- Conjugate base: Sodium acetate (CH_3COONa)

This combination forms an acetate buffer capable of resisting pH fluctuations within a specific range, typically around the pKa of acetic acid (~4.76).

Chemical Principles Underpinning Buffer Preparation

Role of Acetic Acid and Sodium Acetate

Acetic acid is a weak acid with partial ionization in water:

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

Sodium acetate provides the conjugate base (CH_3COO^-), which can neutralize added acids:

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

Together, they form a buffer system that can mitigate pH changes.

The Henderson-Hasselbalch Equation

The pH of a buffer solution is calculated using the Henderson-Hasselbalch equation:

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

Where:
- pK_a of acetic acid ≈ 4.76
- $[\mathrm{A}^-]$ = concentration of conjugate base (sodium acetate)
- $[\mathrm{HA}]$ = concentration of weak acid (acetic acid)

This equation highlights the importance of the ratio of conjugate base to weak acid in determining the pH.

Calculations for Preparing the Buffer Solution

Step 1: Determine the Moles of Sodium Acetate Added

Given:
- Mass of sodium acetate = 21.0 g
- Molar mass of sodium acetate (CH_3COONa) ≈ 82.03 g/mol

Calculations:

$$
\text{Moles of } \mathrm{CH_3COONa} = \frac{21.0\, \text{g}}{82.03\, \text{g/mol}} \approx 0.256\, \text{mol}
$$

Step 2: Determine the Moles of Acetic Acid in Solution

Given:
- Volume of acetic acid solution = 500 mL = 0.5 L
- Concentration of acetic acid = 0.145 M

Calculations:

$$
\text{Moles of } \mathrm{CH_3COOH} = 0.145\, \text{mol/L} \times 0.5\, \text{L} = 0.0725\, \text{mol}
$$

Step 3: Concentrations After Addition

Assuming the sodium acetate dissolves completely and the total volume remains approximately 0.5 L (considering minimal volume change), the concentrations are:

- Acetic acid:

$$
[\mathrm{CH_3COOH}] \approx \frac{0.0725\, \text{mol}}{0.5\, \text{L}} = 0.145\, \text{M}
$$

- Sodium acetate:

$$
[\mathrm{CH_3COO^-}] \approx \frac{0.256\, \text{mol}}{0.5\, \text{L}} = 0.512\, \text{M}
$$

Step 4: pH Calculation Using Henderson-Hasselbalch

Applying the values:

$$
pH = 4.76 + \log \left(\frac{0.512}{0.145}\right) \approx 4.76 + \log(3.53) \approx 4.76 + 0.548 = 5.308
$$

The resulting pH of the buffer is approximately 5.31, suitable for applications requiring a slightly acidic environment.

Practical Considerations for Buffer Preparation

Materials Needed

- Sodium acetate (CH_3COONa)
- Acetic acid (glacial or dilute)
- Distilled water
- Volumetric flask (500 mL)
- Balance for weighing

- pH meter or pH indicator

Preparation Steps

1. Weigh 21.0 g of sodium acetate accurately using a digital balance.

2. Pour approximately 400 mL of distilled water into a 500 mL volumetric flask.

3. Dissolve the sodium acetate completely in the water.

4. In a separate container, measure 500 mL of acetic acid solution (0.145 M). If using glacial acetic acid, dilute accordingly.

5. Add the acetic acid to the sodium acetate solution gradually while stirring.

6. Mix thoroughly and adjust the final volume to 500 mL with distilled water.

7. Measure the pH of the solution with a calibrated pH meter. Adjust if necessary by adding small amounts of acetic acid or sodium acetate.

8. Store the buffer solution properly, labeling it with concentration and pH information.

Adjusting pH

If the pH needs to be fine-tuned:
- Add more sodium acetate to increase pH.
- Add more acetic acid to decrease pH.
- Re-measure after each addition until desired pH is achieved.

Applications of the Acetate Buffer System

Biological Significance

- Maintaining blood pH around 7.4 (though acetate buffer is slightly more acidic, similar principles apply)
- Stabilizing intracellular pH in cell cultures

Industrial and Laboratory Uses

- In enzymatic reactions where pH stability is crucial
- During titrations involving weak acids and bases
- In food industry for preservation and flavor stability

Research and Development

- Designing experiments that require consistent pH conditions
- Developing pharmaceutical formulations

Conclusion

Preparing a buffer solution by adding 21.0 g of sodium acetate to 500 mL of 0.145 M acetic acid exemplifies fundamental principles of chemistry related to acid-base equilibria and buffer capacity. Through precise calculations and careful laboratory techniques, one can produce a solution with a predictable and stable pH, tailored for specific applications. Understanding the interplay between weak acids and their conjugate bases via the Henderson-Hasselbalch equation allows chemists to design buffers with desired pH ranges, essential in biological research, industrial processes, and analytical chemistry. Proper preparation, pH adjustment, and storage ensure the buffer's effectiveness and longevity, demonstrating the importance of meticulous laboratory practice in achieving reliable results.

Keywords: buffer solution, acetic acid, sodium acetate, pH stability, Henderson-Hasselbalch, buffer preparation, chemical calculations, laboratory techniques

Frequently Asked Questions

What is the purpose of preparing a buffer solution using sodium acetate and acetic acid?

The purpose is to create a solution that can resist pH changes upon addition of small amounts of acids or bases, maintaining a relatively stable pH suitable for various chemical applications.

How do you determine the pH of the buffer solution prepared with 21.0 g of sodium acetate and 0.145 M acetic acid?

You use the Henderson-Hasselbalch equation: $pH = pKa + \log([A^-]/[HA])$. First, calculate moles of sodium acetate, convert to molarity, and then substitute the values along with the acetic acid concentration to find the pH.

Why is the amount of sodium acetate (21.0 g) important when preparing the buffer?

The amount determines the concentration of acetate ions in the solution, which influences the buffer’s capacity and its resulting pH when mixed with acetic acid.

What is the approximate pH of the buffer solution created by adding 21.0 g of sodium acetate to 500 mL of 0.145 M acetic acid?

Calculations indicate the pH will be around 4.76, slightly above the pKa of acetic acid (~4.76), due to the added acetate ions from sodium acetate.

How does increasing the amount of sodium acetate affect the pH of the buffer solution?

Increasing sodium acetate raises the concentration of acetate ions, which shifts the buffer's pH to a more basic (higher) value within the buffering range.

Can this buffer solution be used to stabilize pH in biological experiments?

Yes, because it maintains a stable pH around 4.76, making it suitable for experiments that require a mildly acidic environment similar to biological conditions.

Additional Resources

A Buffer Is Prepared By Adding 21.0g Of Sodium Acetate (CH_3COONa) To 500mL Of A 0.145M Acetic Acid (CH_3COOH)

Introduction

Buffer solutions are fundamental in chemistry, biological systems, and industrial processes due to their ability to resist pH changes upon the addition of small amounts of acids or bases. The preparation of an acetate buffer by combining acetic acid (CH_3COOH) with its conjugate base, sodium acetate (CH_3COONa), is a classical example illustrating the principles of chemical equilibrium and buffer capacity. This investigation explores the detailed process, calculations, and implications behind creating such a buffer, starting from the given quantities: 21.0 grams of sodium acetate added to 500 mL of 0.145 M acetic acid solution.

Background and Theoretical Foundations

The Nature of Buffer Solutions

Buffer solutions comprise a weak acid and its conjugate base, or vice versa, which equilibrate to maintain a relatively constant pH. The classic acetic acid–sodium acetate buffer system operates according to the equilibrium:

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

The addition of sodium acetate supplies the conjugate base, CH_3COO^-, which can neutralize added acids, while the acetic acid provides the weak acid component that can neutralize added bases.

The Henderson-Hasselbalch Equation

The pH of the resulting buffer can be accurately predicted using the Henderson-Hasselbalch equation:

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

Where:
- $ pK_a $ is the negative logarithm of the acid dissociation constant, $ K_a $,
- $ [\text{A}^-] $ is the molar concentration of the conjugate base,
- $ [\text{HA}] $ is the molar concentration of the weak acid.

For acetic acid, $ pK_a \approx 4.76 $.

Quantitative Analysis of Buffer Preparation

Step 1: Calculating Moles of Acetic Acid

Given:
- Volume of acetic acid solution $ V = 500\, \text{mL} = 0.5\, \text{L} $,
- Concentration $ C_{HA} = 0.145\, \text{mol/L} $.

Number of moles:

$$ n_{HA} = C_{HA} \times V = 0.145\, \text{mol/L} \times 0.5\, \text{L} = 0.0725\, \text{mol} $$

Step 2: Calculating Moles of Sodium Acetate Added

Given:
- Mass of sodium acetate $ m = 21.0\, \text{g} $,
- Molar mass of sodium acetate $ M_{CH_3COONa} \approx 82.03\, \text{g/mol} $.

Number of moles:

$$ n_{A^-} = \frac{m}{M} = \frac{21.0\, \text{g}}{82.03\, \text{g/mol}} \approx 0.256\, \text{mol} $$

This amount greatly exceeds the initial moles of acetic acid, indicating the buffer will be rich in conjugate base after mixing.

Preparation and Expected Composition

Step 3: Mixing and Resulting Concentrations

Assuming complete dissolution and mixing:

- Moles of acetic acid remain approximately $ 0.0725\, \text{mol} $,
- Moles of sodium acetate added: $ 0.256\, \text{mol} $.

Since the volume added is approximately 0.5 L (assuming negligible volume change), the concentrations after mixing are:

$$ [\text{HA}] = \frac{0.0725\, \text{mol}}{0.5\, \text{L}} = 0.145\, \text{M} $$
$$ [\text{A}^-] = \frac{0.256\, \text{mol}}{0.5\, \text{L}} = 0.512\, \text{M} $$

The conjugate base (sodium acetate) is in significant excess compared to the acetic acid, which will influence the resulting pH.

Calculating the pH of the Prepared Buffer

Using the Henderson-Hasselbalch equation:

$$ \text{pH} = 4.76 + \log \left(\frac{0.512}{0.145} \right) $$

Calculate the ratio:

$$ \frac{0.512}{0.145} \approx 3.53 $$

Taking the logarithm:

$$ \log 3.53 \approx 0.548 $$

Thus,

$$ \text{pH} \approx 4.76 + 0.548 = 5.308 $$

The resulting pH of the buffer is approximately 5.31, indicating a slightly basic environment, appropriate for many biological and chemical applications.

Practical Considerations and Limitations

Volume Changes and Ionic Strength

In laboratory practice, the slight increase in volume due to the addition of sodium acetate solution (if prepared as a solution rather than directly adding solid) can influence molarity calculations. Additionally, increased ionic strength can affect the activity coefficients, slightly modifying the actual pH.

Purity and Sources of Error

- Purity of sodium acetate and acetic acid solutions impacts the final pH.
- Temperature variations affect pK_a; typically, pK_a of acetic acid decreases slightly with increasing temperature.
- The assumption of complete mixing and no volume change simplifies calculations but may not precisely reflect experimental conditions.

Implications and Applications

Biological Significance

An acetate buffer with pH around 5.3 is useful in biological systems where a slightly acidic environment is necessary, such as in certain enzyme reactions or cell culture conditions.

Industrial and Laboratory Uses

- Stabilizing pH in chemical syntheses,
- Preparing media for microbiological growth,
- Calibration and standardization of pH meters.

Educational Value

This case study exemplifies fundamental principles of acid-base chemistry, buffer capacity, and the importance of stoichiometry in solution preparation.

Further Investigations

- Adjusting pH: By varying the amount of sodium acetate added, students can explore how buffer pH changes and optimize for desired pH.
- Buffer Capacity: Quantify how much acid or base the buffer can neutralize before a significant pH change occurs.
- Temperature Dependence: Study how temperature shifts influence pK_a and buffer pH.

Conclusion

The preparation of a buffer solution by adding 21.0 g of sodium acetate to 500 mL of 0.145 M acetic acid exemplifies the practical application of acid-base equilibrium principles. Through detailed stoichiometric calculations, the anticipated pH of approximately 5.31 highlights how adjusting the ratio of weak acid to conjugate base tailors solution properties. Such systems are central in chemical, biological, and industrial contexts, emphasizing the importance of precise measurement, understanding of equilibrium, and the significance of buffer solutions in maintaining stable environmental conditions.

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This comprehensive review underscores the meticulous process involved in buffer preparation and the foundational chemistry principles that govern solution behavior.

A Buffer Is Prepared By Adding 21 0g Of Sodium Acetate Ch3coona To 500ml Of A 0 145m Acetic Acid Ch3cooh

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