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Understanding buffer solutions is essential in chemistry, especially when dealing with acid-base reactions and pH stability. A buffer solution is designed to resist changes in pH when small amounts of acid or base are added. In this article, we will analyze a specific buffer system comprising C6H5NH3Br (anilinium bromide) and C6H5NH2 (aniline), and determine the change in pH when a certain amount of acid or base is introduced.

Introduction to Buffer Solutions

Buffers are solutions that contain a weak acid and its conjugate base or a weak base and its conjugate acid. They maintain a relatively constant pH despite the addition of small quantities of acids or bases. The effectiveness of a buffer depends on the concentrations of its components and their acid-base dissociation constants.

Key Concepts:

- Weak Acid and Conjugate Base: A weak acid partially dissociates in water, and its conjugate base can neutralize added acids.
- Weak Base and Conjugate Acid: Conversely, a weak base reacts with added acids to minimize pH changes.
- Henderson-Hasselbalch Equation: Used to calculate the pH of buffer solutions.

Understanding the Components of the Buffer System

The given buffer contains:

- C6H5NH3Br (Anilinium Bromide): This is the salt form of aniline's conjugate acid, acting as a weak acid source.
- C6H5NH2 (Aniline): A weak base, serving as the conjugate base in the buffer system.

Chemical Equations:

1. Anilinium ion dissociation:

$$
\mathrm{C_6H_5NH_3^+ \leftrightarrow C_6H_5NH_2 + H^+}
$$

2. Aniline acting as a base:

$$
\mathrm{C_6H_5NH_2 + H_2O \leftrightarrow C_6H_5NH_3^+ + OH^-}
$$

Calculating the pKa of Aniline

The pKa of the conjugate acid (anilinium ion) is crucial to determine the pH of the buffer. It can be obtained from literature values.

Known Data:

- The K_a of anilinium ion is approximately 2.6×10^{-10}.
- Therefore, pK_a is calculated as:

$$
pK_a = -\log K_a = -\log(2.6 \times 10^{-10}) \approx 9.59
$$

Note: This value enables the use of the Henderson-Hasselbalch equation to find the pH.

Applying the Henderson-Hasselbalch Equation

The Henderson-Hasselbalch equation relates the pH of a buffer to the concentrations of the conjugate acid and base:

$$
pH = pK_a + \log \left(\frac{[\text{Base}]}{[\text{Acid}]} \right)
$$

In this case:

- Base: C6H5NH2 (aniline), with concentration $0.204\, \text{M}$.
- Acid: C6H5NH3Br (anilinium bromide), with concentration $0.349\, \text{M}$.

Calculating initial pH:

$$
pH_{initial} = 9.59 + \log \left(\frac{0.204}{0.349} \right)
$$

$$
pH_{initial} = 9.59 + \log(0.5845)
$$

$$
pH_{initial} = 9.59 + (-0.232)
$$

$$
pH_{initial} \approx 9.36
$$

This is the initial pH of the buffer before any addition of acid or base.

Determining pH Change When Acid or Base Is Added

To analyze how the pH changes, consider adding a specific amount of acid (HCl) or base (NaOH). We will explore both scenarios:

Scenario 1: Addition of a Strong Acid (HCl)

When a strong acid like HCl is added:

- It will react with the conjugate base (aniline), converting some of it into the conjugate acid.
- This shifts the ratio of acid to base, affecting pH.

Reaction:

$$
\mathrm{C_6H_5NH_2 + H^+ \rightarrow C_6H_5NH_3^+}
$$

Calculations:

Suppose V mL of HCl with molarity M_{HCl} is added. The moles of acid added:

$$
n_{HCl} = M_{HCl} \times V
$$

\]

The change in concentrations:

- The amount of base decreases by \(n_{HCl} \).
- The amount of acid increases by \(n_{HCl} \).

Assuming the total volume remains constant:

\[
[\text{Base}]_{new} = \frac{0.204\, \text{mol} - n_{HCl}}{V_{total}}
\]

\[
[\text{Acid}]_{new} = \frac{0.349\, \text{mol} + n_{HCl}}{V_{total}}
\]

The new pH:

\[
pH_{new} = pK_a + \log \left(\frac{[\text{Base}]_{new}}{[\text{Acid}]_{new}} \right)
\]

Scenario 2: Addition of a Strong Base (NaOH)

When a strong base like NaOH is added:

- It will react with the weak acid (anilinium ion), converting it into the conjugate base.
- The ratio of base to acid shifts accordingly.

Reaction:

\[
\mathrm{C_6H_5NH_3^+ + OH^- \rightarrow C_6H_5NH_2 + H_2O}
\]

Calculations:

Similarly, for \(V \) mL of NaOH with molarity \(M_{NaOH} \):

- Moles of base added: \(n_{NaOH} = M_{NaOH} \times V \).

- The acid decreases by \(n_{NaOH} \), and the base increases by \(n_{NaOH} \).

Updated concentrations:

\[
[\text{Acid}]_{new} = \frac{0.349\, \text{mol} - n_{NaOH}}{V_{total}}
\]

\[

$$
[\text{Base}]_{new} = \frac{0.204\, \text{mol} + n_{NaOH}}{V_{total}}
$$

The pH:

$$
pH_{new} = pK_a + \log \left(\frac{[\text{Base}]_{new}}{[\text{Acid}]_{new}} \right)
$$

Practical Example: Quantitative pH Change Calculation

Suppose 10 mL of 0.1 M HCl is added to the buffer:

- Moles of HCl:

$$
n_{HCl} = 0.1\, \text{mol/L} \times 0.010\, \text{L} = 0.001\, \text{mol}
$$

- Volume of total solution (assuming initial volume is 1 L):

$$
V_{total} \approx 1\, \text{L} + 0.010\, \text{L} \approx 1.01\, \text{L}
$$

- Updated concentrations:

$$
[\text{Base}]_{new} = \frac{0.204 - 0.001}{1.01} \approx 0.202\, \text{M}
$$

$$
[\text{Acid}]_{new} = \frac{0.349 + 0.001}{1.01} \approx 0.347\, \text{M}
$$

- New pH:

$$
pH_{new} = 9.59 + \log \left(\frac{0.202}{0.347} \right) = 9.59 + \log(0.582) \approx 9.59 - 0.235 = 9.355
$$

- pH change:

$$

\Delta pH = 9.36 - 9.355 \approx 0.005
\]

This small change illustrates the buffer's capacity to resist pH changes upon addition of small amounts of acid.

Factors Affecting the Buffer’s pH and Capacity

Several factors influence the effectiveness and pH stability of a buffer:

- Concentrations of Buffer Components: Higher concentrations generally increase buffer capacity.
- pKa of the Acid/Base Pair: The closer the pKa to the target pH, the better the buffer

Frequently Asked Questions

What is the primary purpose of using a buffer solution containing C6H5NH3Br and C6H5NH2?

The buffer solution is used to maintain a relatively stable pH when small amounts of acids or bases are added, leveraging the weak base aniline (C6H5NH2) and its conjugate acid (C6H5NH3Br).

How does adding a strong acid affect the pH of this buffer solution?

Adding a strong acid will react with the weak base C6H5NH2, converting it to its conjugate acid C6H5NH3+, thus causing a slight decrease in pH, but the buffer resists significant pH change.

What is the role of C6H5NH3Br in this buffer system?

C6H5NH3Br provides the conjugate acid form of aniline, which can neutralize added bases and help maintain the pH stability of the buffer solution.

How can we calculate the pH change when a known amount of acid or base is added to this buffer?

The pH change can be calculated using the Henderson-Hasselbalch equation, considering the initial concentrations of the weak base and its conjugate acid, along with the amount of acid or base added.

What factors influence the buffer capacity of this solution?

The buffer capacity depends on the concentrations of C6H5NH3Br and C6H5NH2; higher concentrations increase the solution's ability to resist pH changes upon addition of acids or bases.

If 0.01 mol of HCl is added to this buffer, how would you determine the resulting pH change?

You would use the initial concentrations, the moles of HCl added, and the Henderson-Hasselbalch equation to find the new concentrations of the weak base and its conjugate acid, then calculate the new pH.

Why is it important to know the initial concentrations of C6H5NH3Br and C6H5NH2 when analyzing pH changes?

Knowing these concentrations allows for accurate calculation of the buffer's capacity and the extent of pH change upon addition of acids or bases, ensuring precise pH predictions.

Additional Resources

A Buffer Solution Contains 0.349 M C6H5NH3Br and 0.204 M C6H5NH2 (Aniline): Determine the pH Change When

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Introduction

Buffer solutions are fundamental in chemistry due to their ability to resist significant changes in pH when small amounts of acids or bases are added. They play a vital role in biological systems, industrial processes, and analytical chemistry. Understanding how buffer systems respond to external disturbances—such as the addition of acids or bases—is crucial for optimizing their application.

This article provides a comprehensive investigation into a specific buffer solution composed of 0.349 M phenylammonium bromide ($C_6H_5NH_3Br$) and 0.204 M aniline ($C_6H_5NH_2$). The primary focus is to determine how the pH of this buffer changes upon the addition of small quantities of acid or base, analyzing the underlying chemical equilibria, and calculating the resulting pH variations.

Composition and Nature of the Buffer Components

Phenylammonium Bromide ($C_6H_5NH_3Br$)

Phenylammonium bromide is a salt derived from phenylamine (aniline) and hydrobromic acid (HBr). In aqueous solution, it dissociates into phenylammonium cations ($C_6H_5NH_3^+$) and bromide ions (Br^-):

$$ \mathrm{C_6H_5NH_3Br} \rightarrow \mathrm{C_6H_5NH_3^+} + \mathrm{Br^-} $$

The phenylammonium ion acts as a weak acid:

$$ \mathrm{C_6H_5NH_3^+} \leftrightarrow \mathrm{C_6H_5NH_2} + \mathrm{H^+} $$

Aniline ($C_6H_5NH_2$)

Aniline is a weak base, characterized by the following equilibrium:

$$ \mathrm{C_6H_5NH_2} + \mathrm{H_2O} \leftrightarrow \mathrm{C_6H_5NH_3^+} + \mathrm{OH^-} $$

The base dissociation constant (K_b) for aniline is approximately 4.3×10^{-10}.

Understanding the Buffer System

The buffer system comprises:

- Conjugate acid (phenylammonium ion): $\mathrm{C_6H_5NH_3^+}$
- Conjugate base (aniline): $\mathrm{C_6H_5NH_2}$

The equilibrium between these two species governs the pH stability.

The relevant acid dissociation equilibrium:

$$ \mathrm{C_6H_5NH_3^+} \leftrightarrow \mathrm{C_6H_5NH_2} + \mathrm{H^+} $$

The equilibrium constant (Ka) for phenylammonium is derived from the known Kb of aniline and the ion product of water:

$$ K_a = \frac{K_w}{K_b} $$

where:

- $K_w = 1.0 \times 10^{-14}$,
- $K_b (\mathrm{aniline}) = 4.3 \times 10^{-10}$.

Calculating K_a:

$$ K_a = \frac{1.0 \times 10^{-14}}{4.3 \times 10^{-10}} \approx 2.33 \times 10^{-5} $$

Corresponding to a pKa:

$$\mathrm{p}K_a = -\log K_a \approx 4.63$$

Initial pH of the Buffer Solution

Applying the Henderson-Hasselbalch Equation

The initial pH can be estimated using the Henderson-Hasselbalch equation:

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

Where:

- $[\mathrm{A^-}] = [\mathrm{C_6H_5NH_2}] = 0.204\, \mathrm{M}$,
- $[\mathrm{HA}] = [\mathrm{C_6H_5NH_3^+}] = 0.349\, \mathrm{M}$,
- $\mathrm{p}K_a \approx 4.63$.

Calculating:

$$\mathrm{pH}_0 = 4.63 + \log \left(\frac{0.204}{0.349} \right)$$

$$\mathrm{pH}_0 = 4.63 + \log (0.584)$$

$$\mathrm{pH}_0 = 4.63 - 0.233$$

$$\boxed{\mathrm{pH}_0 \approx 4.40}$$

This initial pH suggests a slightly acidic buffer, consistent with the presence of a weak acid and its conjugate base.

Effect of Acid and Base Addition on the Buffer

The buffer's ability to resist pH change hinges on the relative concentrations of $\mathrm{C_6H_5NH_3^+}$ and $\mathrm{C_6H_5NH_2}$. When a small amount of acid (e.g., HCl) or base (e.g., NaOH) is added, these species either consume or produce $\mathrm{H^+}$ or $\mathrm{OH^-}$, shifting the equilibrium minimally due to the buffer capacity, which can be quantified.

Quantitative Analysis of pH Change

Addition of Acid (e.g., HCl)

Suppose an infinitesimal amount δ molarity of HCl is added:

$$\mathrm{HCl} \rightarrow \mathrm{H^+} + \mathrm{Cl^-}$$

The $\mathrm{H^+}$ will react with the conjugate base (phenylamine):

$$\mathrm{C_6H_5NH_2} + \mathrm{H^+} \rightarrow \mathrm{C_6H_5NH_3^+}$$

This shifts the equilibrium towards the conjugate acid, increasing $[\mathrm{C_6H_5NH_3^+}]$ and decreasing $[\mathrm{C_6H_5NH_2}]$, thus lowering the pH.

The change in pH can be approximated using the Henderson-Hasselbalch equation after the addition:

$$\mathrm{pH}_\text{new} = \mathrm{p}K_a + \log \left(\frac{[\mathrm{A^-}] - \delta}{[\mathrm{HA}] + \delta} \right)$$

assuming the volume change is negligible and the added acid reacts completely with available base.

Addition of Base (e.g., NaOH)

Similarly, adding a small amount δ molarity of NaOH:

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

The $\mathrm{OH^-}$ will react with phenonium ion (weak acid):

$$\mathrm{C_6H_5NH_3^+} + \mathrm{OH^-} \rightarrow \mathrm{C_6H_5NH_2} + \mathrm{H_2O}$$

This shifts the equilibrium toward the conjugate base, increasing $[\mathrm{C_6H_5NH_2}]$ and decreasing $[\mathrm{C_6H_5NH_3^+}]$, raising the pH.

Practical Calculation: Estimating pH Change for a Specific Addition

Assuming the addition of 0.001 mol of HCl to 1 L of buffer:

- Initial concentrations:

- $[\mathrm{C_6H_5NH_3^+}] = 0.349\, \mathrm{M}$,

- $[\mathrm{C_6H_5NH_2}] = 0.204\, \mathrm{M}$.

- Moles of acid added: 0.001 mol.

- Moles of $\mathrm{C_6H_5NH_3^+}$ increase by 0.001 mol:

- New $[\mathrm{HA}] = 0.349 + 0.001 = 0.350\, \mathrm{M}$,

- Moles of $\mathrm{C_6H_5NH_2}$ decrease by 0.001 mol:

- New $[\mathrm{A^-}] = 0.204 - 0.001 = 0.203\, \mathrm{M}$.

Applying the Henderson-Hasselbalch equation:

$$\mathrm{pH}_\text{new} = 4.63 + \log \left(\frac{0.203}{0.350} \right)$$

$$\mathrm{pH}_\text{new} = 4.63 + \log (0.58)$$

\[

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