A Buffer Solution Is Prepared By Mixing 250 ML Of 1.00 M Nitrous Acid With 50 ML Of 1.00 M Sodium Hydroxide.

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Understanding Buffer Solutions and Their Importance

A buffer solution is a specialized mixture that resists changes in pH when small amounts of acid or base are added. These solutions are vital in various fields such as chemistry, biology, medicine, and environmental science because many biological and chemical processes require a stable pH environment. The ability of a buffer to maintain a consistent pH depends on the presence of a weak acid and its conjugate base or a weak base and its conjugate acid.

In this article, we will explore the preparation of a buffer solution by mixing nitrous acid (a weak acid) with sodium hydroxide (a strong base), analyze the chemical reactions involved, and understand how such a mixture exhibits buffering capacity.

Details of the Buffer Preparation

Initial Components and Their Concentrations

The preparation involves two main components:

- Nitrous Acid (HNO_2): 250 mL of a 1.00 M solution.
- Sodium Hydroxide (NaOH): 50 mL of a 1.00 M solution.

This mixture aims to produce a buffer solution with a particular pH, primarily dictated by the weak acid (nitrous acid) and the base (NaOH).

Understanding the Volumes and Moles

Calculating the initial moles of each component:

- Moles of Nitrous Acid:

$$
n_{HNO_2} = M \times V = 1.00\, \text{mol/L} \times 0.250\, \text{L} = 0.250\, \text{mol}
$$

- Moles of Sodium Hydroxide:

$$
n_{NaOH} = 1.00\, \text{mol/L} \times 0.050\, \text{L} = 0.050\, \text{mol}
$$

Since NaOH is a strong base, it will react with the weak acid, nitrous acid, altering the concentrations of both species in the solution.

Chemical Reactions Involved in Buffer Formation

Reaction Between Nitrous Acid and Sodium Hydroxide

The primary reaction is:

$$
HNO_2 + OH^- \rightarrow NO_2^- + H_2O
$$

This is a typical acid-base neutralization where the weak acid (HNO_2) reacts with the hydroxide ion (OH^-) from NaOH to produce the conjugate base, nitrite ion (NO_2^-), and water.

Extent of the Reaction

- The amount of NaOH added (0.050 mol) will react with an equivalent amount of nitrous acid.
- Remaining nitrous acid after reaction:

$$
n_{HNO_2,\,remaining} = 0.250\, \text{mol} - 0.050\, \text{mol} = 0.200\, \text{mol}
$$

- Moles of nitrite ion formed:

$$
n_{NO_2^-} = 0.050\, \text{mol}
$$

This process results in a solution containing both unreacted nitrous acid and its conjugate base, nitrite ion, creating a buffer system.

Calculating Final Concentrations of Buffer Components

Final Volumes

Total volume after mixing:

$$
V_{total} = 250\, \text{mL} + 50\, \text{mL} = 300\, \text{mL} = 0.300\, \text{L}
$$

Concentrations of Acid and Conjugate Base

- Concentration of remaining nitrous acid (HNO_2):

$$
[HNO_2] = \frac{0.200\, \text{mol}}{0.300\, \text{L}} \approx 0.667\, \text{M}
$$

- Concentration of nitrite ion (NO_2^-):

$$
[NO_2^-] = \frac{0.050\, \text{mol}}{0.300\, \text{L}} \approx 0.167\, \text{M}
$$

These concentrations define the buffer system: a mixture of a weak acid and its conjugate base.

---

pH Calculation of the Buffer Solution

Using the Henderson-Hasselbalch Equation

The pH of a buffer solution can be estimated by:

$$
\text{pH} = pK_a + \log \left(\frac{[\text{conjugate base}]}{[\text{weak acid}]} \right)
$$

\]

Where:

- pK_a of nitrous acid is approximately 3.37.

Plugging in the concentrations:

$$
\text{pH} = 3.37 + \log \left(\frac{0.167}{0.667} \right)
$$

Calculating:

$$
\log \left(\frac{0.167}{0.667} \right) = \log(0.25) \approx -0.60
$$

Thus:

$$
\text{pH} \approx 3.37 - 0.60 = 2.77
$$

The resulting pH indicates an acidic buffer, suitable for various applications such as biochemical assays.

Applications of the Buffer Solution

Buffers like the one prepared from nitrous acid and sodium hydroxide are essential in numerous contexts:

- Biochemical experiments: Maintaining enzyme activity and stability.
- Industrial processes: pH control in manufacturing.
- Environmental monitoring: Stabilizing pH in water treatment.
- Analytical chemistry: Preparing standards and controlling pH during titrations.

Factors Affecting Buffer Capacity and Stability

The buffer’s effectiveness depends on several factors:

- Concentration of weak acid and conjugate base: Higher concentrations generally increase buffer

capacity.
- pKa of the acid: The closest the pKa is to the desired pH, the better the buffering capacity.
- Temperature: Changes in temperature can alter pKa and buffer capacity.
- Addition of external acids or bases: Excessive addition can overwhelm the buffer system.

Advantages of Using Nitrous Acid and Sodium Hydroxide in Buffer Preparation

- Simple preparation: Mixing known concentrations of acid and base allows precise control over pH.
- Flexibility: Adjusting volumes and concentrations can tailor the buffer to specific pH requirements.
- Relevance in laboratory settings: Common reagents that are readily available and easy to handle.

Practical Tips for Preparing Buffer Solutions

- Always use calibrated instruments for measuring liquids.
- Mix solutions gradually and stir continuously.
- Confirm the pH with a calibrated pH meter after preparation.
- Adjust pH if necessary by adding small amounts of acid or base.

Conclusion

Preparing a buffer solution by mixing 250 mL of 1.00 M nitrous acid with 50 mL of 1.00 M sodium hydroxide is a straightforward process rooted in fundamental acid-base chemistry. This mixture results in a solution containing both unreacted nitrous acid and its conjugate base, nitrite ion, forming an effective buffer system. Through calculations involving reaction stoichiometry and the Henderson-Hasselbalch equation, we can estimate the pH and understand the buffer’s capacity to resist pH changes. Such buffer solutions are indispensable in scientific research, industrial processes, and environmental management, highlighting the importance of mastering their preparation and application.

Keywords: buffer solution, nitrous acid, sodium hydroxide, pH, Henderson-Hasselbalch, conjugate base, weak acid, chemical equilibrium, buffer capacity

Frequently Asked Questions

What is the purpose of preparing a buffer solution using nitrous acid and sodium hydroxide?

The purpose is to create a solution that can resist pH changes upon addition of small amounts of acids or bases, due to the weak acid (nitrous acid) and its conjugate base formed by reacting with NaOH.

How do you determine the initial moles of nitrous acid and sodium hydroxide used in the solution?

The initial moles are calculated by multiplying volume (in liters) by molarity: for nitrous acid, 0.250 L × 1.00 M = 0.25 mol; for sodium hydroxide, 0.050 L × 1.00 M = 0.05 mol.

What is the expected pH of the buffer solution after mixing nitrous acid with sodium hydroxide?

The pH can be estimated using the Henderson-Hasselbalch equation, considering the remaining nitrous acid and the formed sodium nitrite; it will be slightly above the pKa of nitrous acid (~3.37), indicating a mildly acidic buffer.

How does the addition of sodium hydroxide affect the composition of the buffer solution?

NaOH reacts with nitrous acid to produce its conjugate base, nitrite ions, thus increasing the conjugate base concentration and maintaining the pH within the buffer range.

Why is the total volume of the buffer solution important in calculating the final pH?

The total volume determines the concentrations of the acid and conjugate base; after mixing, the combined volume affects molarity calculations and the resulting pH.

Can this buffer solution effectively neutralize added acids or bases? Why?

Yes, because it contains both a weak acid and its conjugate base, allowing it to neutralize small amounts of added acids or bases without significant pH change.

What role does the molarity of the initial solutions play in the strength of the resulting buffer?

Higher molarity of the acid and conjugate base results in a more robust buffer capacity, allowing it to resist larger pH changes upon addition of external acids or bases.

Additional Resources

A Buffer Solution Is Prepared By Mixing 250 ML Of 1.00 M Nitrous Acid With 50 ML Of 1.00 M Sodium Hydroxide

Creating buffer solutions is a fundamental aspect of chemistry, especially in applications that require maintaining a stable pH environment. In the specific scenario where 250 mL of 1.00 M nitrous acid (HNO_2) is mixed with 50 mL of 1.00 M sodium hydroxide (NaOH), an interesting chemical process unfolds. This process exemplifies the principles of acid-base reactions, buffer capacity, and pH calculation, offering a valuable insight into how buffers are formed and their characteristics. This review delves into the chemistry behind this particular mixture, exploring the reaction mechanisms, resulting pH, and practical implications, while providing a comprehensive understanding suitable for students, educators, and professionals alike.

Understanding the Components

Nitrous Acid (HNO_2)

Nitrous acid is a weak, monobasic acid with the chemical formula HNO_2. It exists primarily in the form of undissociated acid in aqueous solutions but can dissociate partially according to the following equilibrium:

$$\mathrm{HNO_2} \leftrightarrow \mathrm{H^+} + \mathrm{NO_2^-}$$

The acid dissociation constant (Ka) for nitrous acid is approximately 4.5×10^{-4}, which indicates moderate acidity. Its conjugate base, nitrite ion (NO_2^-), plays a significant role in buffer action when present in sufficient quantities.

Sodium Hydroxide (NaOH)

Sodium hydroxide is a strong base that dissociates completely in aqueous solution:

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

When added to an acidic solution, it neutralizes some of the acid, generating water and the conjugate base, which then acts as part of the buffer system.

Reaction Mechanism and Buffer Formation

When 50 mL of 1.00 M NaOH is added to 250 mL of 1.00 M nitrous acid, the hydroxide ions (OH^-) react with the acid (HNO_2) as follows:

$$ \mathrm{HNO_2} + \mathrm{OH^-} \rightarrow \mathrm{NO_2^-} + \mathrm{H_2O} $$

This reaction consumes some of the nitrous acid, producing nitrite ions (NO_2^-). The mixture then contains both undissociated HNO_2 and its conjugate base NO_2^-, which is characteristic of a buffer solution.

Key Points:

- The amount of NaOH added is less than the amount of nitrous acid present, so not all acid is neutralized.
- The resulting solution contains a mixture of HNO_2 and NO_2^-, which can resist changes in pH upon addition of small amounts of acid or base.
- The pH of the mixture can be calculated using the Henderson-Hasselbalch equation, considering the concentrations of the acid and conjugate base.

Calculating the Resulting pH

Step 1: Determine Moles of Reactants

- Moles of nitrous acid initially:

$$ 250\, \mathrm{mL} \times 1.00\, \mathrm{mol/L} = 0.250\, \mathrm{L} \times 1.00\, \mathrm{mol/L} = 0.250\, \mathrm{mol} $$

- Moles of NaOH added:

$$ 50\, \mathrm{mL} \times 1.00\, \mathrm{mol/L} = 0.050\, \mathrm{L} \times 1.00\, \mathrm{mol/L} = 0.050\, \mathrm{mol} $$

Step 2: Reaction and Remaining Moles

- NaOH will neutralize an equivalent amount of HNO_2:

$$ \mathrm{HNO_2} + \mathrm{OH^-} \rightarrow \mathrm{NO_2^-} + \mathrm{H_2O} $$

- Moles of HNO_2 after reaction:

$$ 0.250\, \mathrm{mol} - 0.050\, \mathrm{mol} = 0.200\, \mathrm{mol} $$

- Moles of NO_2^- formed:

$$ 0.050\, \mathrm{mol} $$

Step 3: Concentrations in the Final Mixture

- Total volume of the mixture:

$$ 250\, \mathrm{mL} + 50\, \mathrm{mL} = 300\, \mathrm{mL} = 0.300\, \mathrm{L} $$

- Concentration of remaining HNO_2:

$$ \frac{0.200\, \mathrm{mol}}{0.300\, \mathrm{L}} \approx 0.667\, \mathrm{M} $$

- Concentration of NO_2^-:

$$ \frac{0.050\, \mathrm{mol}}{0.300\, \mathrm{L}} \approx 0.167\, \mathrm{M} $$

Step 4: Applying the Henderson-Hasselbalch Equation

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

Where:

- $pK_a = -\log(Ka) \approx -\log(4.5 \times 10^{-4}) \approx 3.35$
- $[\mathrm{A^-}] = 0.167\, \mathrm{M}$
- $[\mathrm{HA}] = 0.667\, \mathrm{M}$

Calculating:

$$ \mathrm{pH} = 3.35 + \log \left(\frac{0.167}{0.667} \right) = 3.35 + \log(0.25) \approx 3.35 - 0.60 = 2.75 $$

Result: The pH of the resulting buffer solution is approximately 2.75.

Features and Practical Significance

Advantages of the Prepared Buffer

- Stability: The buffer resists pH changes upon addition of small amounts of acids or bases, making it useful in various chemical and biological experiments.
- Relevance in Biological Systems: Such buffers mimic natural systems where weak acids and their conjugate bases maintain pH homeostasis.
- Adjustability: By altering the amounts of acid and base, the pH can be fine-tuned for specific applications.

Limitations and Considerations

- Limited Buffer Capacity: The amount of acid and conjugate base determines the amount of acid or base the buffer can neutralize before pH shifts significantly.
- pH Range: The effective pH range is typically within ±1 unit of the pKa; in this case, around pH 2.7, which is quite acidic.
- Incomplete Neutralization: Excess reactants can shift the pH outside the buffer range or cause precipitation or other side reactions.

Features Summary:

Feature	Description
pH Range	Approximately 2.7, suitable for acidic buffer applications
Buffer Capacity	Moderate; dependent on initial concentrations
Ease of Preparation	Simple mixing of known concentrations and volumes
Chemical Stability	Stable under proper storage conditions

Applications and Implications

This buffer system has practical applications in various fields:

- Analytical Chemistry: Maintaining consistent pH during titrations or spectroscopic measurements.
- Biochemistry: Mimicking biological pH environments, especially in studies involving nitrous acid derivatives or nitrite chemistry.
- Industrial Processes: In manufacturing where controlled acidity is essential.
- Environmental Chemistry: Understanding natural buffers in soil and water systems containing nitrous compounds.

Furthermore, understanding the chemistry behind such buffer solutions aids in designing systems for controlled pH environments, essential in pharmaceuticals, food preservation, and chemical manufacturing.

Conclusion

Mixing 250 mL of 1.00 M nitrous acid with 50 mL of 1.00 M sodium hydroxide exemplifies a classic acid-base neutralization that results in a buffer solution. The key takeaway is how partial neutralization leads to a mixture of weak acid and its conjugate base, conferring the solution with buffering capacity. The calculated pH of approximately 2.75 highlights the acidic nature of the resulting buffer, which can be further tailored by modifying initial concentrations or volumes. Understanding these principles is fundamental for chemists working in laboratory, industrial, or environmental contexts, as it underscores the importance of precise calculations and knowledge of chemical equilibria in creating effective buffer systems. The process also underscores the broader significance of weak acid-conjugate base buffers in maintaining pH stability across various scientific disciplines.

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