

What Is The Ph Of A Buffer Made By Combining 45.0 Ml Of 0.150 M Nitrous Acid And 20.0 Ml Of 0.175 M Sodium

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Understanding the pH of buffer solutions is fundamental in chemistry, particularly in fields like biochemistry, environmental science, and industrial applications. Buffer solutions help maintain a stable pH in various biological and chemical processes, making them crucial for ensuring optimal conditions for reactions and processes. This article delves into the detailed calculation of the pH of a specific buffer solution formed by mixing nitrous acid (HNO_2) and sodium nitrite (NaNO_2), explaining each step comprehensively to enhance understanding for students, educators, and professionals alike.

Introduction to Buffer Solutions and Their Importance

Buffers are solutions that resist significant changes in pH upon the addition of small amounts of acids or bases. They are vital in maintaining the pH of biological fluids, industrial processes, and environmental systems within narrow limits. The typical buffer system involves a weak acid and its conjugate base or a weak base and its conjugate acid.

Why Buffers Matter:

- Biological significance: Human blood, for example, maintains a pH around 7.4 thanks to buffer systems involving carbonic acid and bicarbonate.
- Industrial applications: Manufacturing processes often require specific pH levels for optimal product quality.
- Environmental relevance: Natural waters and soils contain buffer systems that influence nutrient availability and pollutant behavior.

Understanding how to calculate the pH of such buffer solutions is essential for designing and controlling chemical processes effectively.

Components of the Buffer Solution

The specific buffer solution discussed here is formed by mixing:

- 45.0 mL of 0.150 M nitrous acid (HNO_2)
- 20.0 mL of 0.175 M sodium nitrite (NaNO_2)

Let's clarify the roles of these components:

- Nitrous acid (HNO_2): A weak acid that partially dissociates in water.
- Sodium nitrite (NaNO_2): The salt of nitrous acid, providing the conjugate base, nitrite ion (NO_2^-).

The mixture of these components results in a buffer system capable of resisting pH changes, primarily governed by the principles of weak acid and conjugate base equilibria.

Calculating the pH of the Buffer Solution

The process involves several steps:

1. Determine the moles of each component before mixing.
2. Calculate the total moles of each component after mixing.
3. Use the Henderson-Hasselbalch equation to find the pH.

Let's walk through each step in detail.

Step 1: Calculate the Moles of Nitrous Acid (HNO_2)

The molarity (M) is given as 0.150 M, and the volume is 45.0 mL.

Convert volume to liters:

$$45.0 \text{ mL} = 0.0450 \text{ L}$$

Calculate moles:

$$\text{moles of HNO}_2 = M \times V = 0.150 \text{ mol/L} \times 0.0450 \text{ L} = 0.00675 \text{ mol}$$

Step 2: Calculate the Moles of Sodium Nitrite (NaNO_2)

Similarly, for NaNO_2 :

$$\text{Volume: } 20.0 \text{ mL} = 0.0200 \text{ L}$$

$$\text{Molarity: } 0.175 \text{ mol/L}$$

Calculate moles:

$$\text{moles of NaNO}_2 = 0.175 \text{ mol/L} \times 0.0200 \text{ L} = 0.00350 \text{ mol}$$

Step 3: Determine the Concentrations in the Final Mixture

When these solutions are combined, the total volume is:

$$V_{\text{total}} = 45.0\text{ mL} + 20.0\text{ mL} = 65.0\text{ mL} = 0.0650\text{ L}$$

Calculate the concentrations of each component in the mixture:

- Nitrous acid (HNO_2):

$$[\text{HNO}_2] = \frac{0.00675\text{ mol}}{0.0650\text{ L}} \approx 0.1038\text{ M}$$

- Nitrite ion (NO_2^-):

$$[\text{NO}_2^-] = \frac{0.00350\text{ mol}}{0.0650\text{ L}} \approx 0.0538\text{ M}$$

The Acid-Base Equilibrium and Henderson-Hasselbalch Equation

The pH of a buffer solution containing a weak acid and its conjugate base can be effectively calculated using the Henderson-Hasselbalch equation:

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

Where:

- pK_a is the negative logarithm of the acid dissociation constant (K_a)
- $[\text{A}^-]$ is the concentration of the conjugate base
- $[\text{HA}]$ is the concentration of the weak acid

Identifying the acid dissociation constant:

Nitrous acid (HNO_2) has a pK_a value approximately 3.37 at 25°C.

Applying the values:

- $[\text{HA}] = 0.1038\text{ M}$ (nitrous acid)
- $[\text{A}^-] = 0.0538\text{ M}$ (nitrite ion)

Calculate:

$$\text{pH} = 3.37 + \log \left(\frac{0.0538}{0.1038} \right)$$

$$\text{pH} = 3.37 + \log(0.518)$$

$$\text{pH} = 3.37 + (-0.285)$$

$$\text{pH} \approx 3.085$$

Thus, the pH of the buffer solution is approximately 3.09.

Understanding the Significance of the Calculated pH

The pH value around 3.09 indicates an acidic buffer solution, which is consistent with the presence of a weak acid (nitrous acid) and its conjugate base (nitrite). This pH is ideal for specific applications, such as certain biochemical assays or industrial processes requiring mildly acidic conditions.

Implications of this pH:

- The buffer effectively resists pH changes around this value.
- It can be used in contexts where a stable pH near 3.1 is necessary.
- Understanding the relationship between component concentrations and pH helps in designing buffers for targeted applications.

Factors Affecting Buffer pH and Their Control

While the calculations provide a theoretical pH, real-world buffers may experience variations due to several factors:

- Temperature: Changes in temperature affect pK_a values.
- Additional acids or bases: Any extraneous acids or bases can shift the pH.
- Dilution effects: Further dilution can influence the buffer capacity and pH.
- Impurities: Contaminants can alter the equilibrium.

Controlling these factors is essential for maintaining the desired pH in practical applications.

Applications of Nitrous Acid/Nitrite Buffer Systems

Understanding the pH of nitrous acid-nitrite buffers is vital across various fields:

- Analytical Chemistry: Used in titrations and pH calibration standards.
- Environmental Science: Monitoring nitrogen compounds in water bodies.
- Food Industry: Preserving certain food products with specific pH requirements.
- Pharmaceuticals: Formulating medications requiring specific pH environments.

By precisely calculating buffer pH, scientists can optimize processes and ensure consistent results.

Conclusion

The detailed calculation of the pH of a buffer formed by mixing nitrous acid and sodium nitrite demonstrates the importance of stoichiometry, equilibrium chemistry, and the Henderson-Hasselbalch equation in practical applications. In this case, mixing 45.0 mL of 0.150 M nitrous acid with 20.0 mL of 0.175 M sodium nitrite results in a buffer with an approximate pH of 3.09. Mastery of such calculations enables chemists to design effective buffer systems tailored to specific pH requirements, ensuring optimal conditions across a broad spectrum of scientific and industrial endeavors.

Frequently Asked Questions

What is the pH of a buffer made by combining 45.0 mL of 0.150 M nitrous acid and 20.0 mL of 0.175 M sodium nitrite?

To find the pH, first calculate the moles of each component: Moles of nitrous acid = $0.150 \text{ mol/L} \times 0.045 \text{ L} = 0.00675 \text{ mol}$; moles of sodium nitrite = $0.175 \text{ mol/L} \times 0.020 \text{ L} = 0.0035 \text{ mol}$. Since sodium nitrite is the conjugate base, use the Henderson-Hasselbalch equation: $\text{pH} = \text{pK}_a + \log([A^-]/[HA])$. The pK_a of nitrous acid is approximately 3.37. Adjust concentration for the total volume ($45.0 \text{ mL} + 20.0 \text{ mL} = 65.0 \text{ mL}$): $[A^-] = 0.0035 \text{ mol} / 0.065 \text{ L} \approx 0.0538 \text{ M}$, $[HA] = 0.00675 \text{ mol} / 0.065 \text{ L} \approx 0.1038 \text{ M}$. Therefore, $\text{pH} \approx 3.37 + \log(0.0538/0.1038) \approx 3.37 + \log(0.518) \approx 3.37 - 0.285 \approx 3.09$.

How do you determine the pH of a buffer solution created from nitrous acid and sodium nitrite?

You determine the pH by applying the Henderson-Hasselbalch equation, which requires knowing the concentrations of the weak acid (nitrous acid) and its conjugate base (sodium nitrite) after mixing, as well as the pK_a of nitrous acid. Calculate the moles of each component, convert to concentrations based on the total volume, and then substitute into the equation to find the pH.

Why is the pH of the buffer made from nitrous acid and

sodium nitrite important?

The pH of this buffer is crucial because it determines the solution's ability to resist pH changes upon addition of acids or bases. This is important in various chemical and biological applications where maintaining a stable pH is necessary for optimal function or stability.

What is the role of sodium nitrite in the buffer solution with nitrous acid?

Sodium nitrite provides the conjugate base (NO_2^-) that pairs with nitrous acid, allowing the solution to act as a buffer by neutralizing added acids or bases, thereby stabilizing the pH.

How does changing the volume of nitrous acid or sodium nitrite affect the pH of the buffer?

Changing the volume of either component alters the moles and therefore the concentrations of the acid and base in the buffer. Increasing the amount of nitrous acid lowers pH, while increasing sodium nitrite raises pH. The balance between these determines the buffer's pH according to the Henderson-Hasselbalch equation.

Can you estimate the pH of the buffer if you double the volume of sodium nitrite used?

Yes. Doubling the volume of sodium nitrite would double the moles of conjugate base, increasing its concentration relative to the acid. This would increase the ratio $[\text{A}^-]/[\text{HA}]$, resulting in a higher pH. Precise calculation would involve recomputing concentrations after the volume change and applying the Henderson-Hasselbalch equation.

Additional Resources

What Is The pH Of A Buffer Made By Combining 45.0 mL Of 0.150 M Nitrous Acid And 20.0 mL Of 0.175 M Sodium?

Understanding how to determine the pH of a buffer solution created by mixing a weak acid and a strong base is fundamental in chemistry, especially in fields like biochemistry, environmental science, and pharmaceutical development. In this guide, we will explore the step-by-step process to calculate the pH of a buffer formed by combining 45.0 mL of 0.150 M nitrous acid (HNO_2) with 20.0 mL of 0.175 M sodium hydroxide (NaOH). This example illustrates core concepts such as molarity, titration, acid-base reactions, and the application of the Henderson-Hasselbalch equation.

Introduction to Buffer Solutions and pH Calculation

Buffer solutions are vital in maintaining stable pH levels in biological systems and industrial processes. They typically consist of a weak acid and its conjugate base or a weak

base and its conjugate acid. When small amounts of acids or bases are added, these solutions resist significant changes in pH.

In our case, nitrous acid (HNO_2) is a weak acid, and sodium hydroxide (NaOH) is a strong base. When mixed, they react to produce the conjugate base (nitrite ion, NO_2^-), creating a buffer system. The pH of this buffer depends on the relative amounts of the acid and conjugate base present after the reaction, which can be calculated using the Henderson-Hasselbalch equation.

Step 1: Understanding the Components

Nitrous Acid (HNO_2)

- Concentration: 0.150 M
- Volume: 45.0 mL (0.045 L)
- Initial moles:

$$\text{moles of HNO}_2 = M \times V = 0.150 \text{ mol/L} \times 0.045 \text{ L} = 0.00675 \text{ mol}$$

Sodium Hydroxide (NaOH)

- Concentration: 0.175 M
- Volume: 20.0 mL (0.020 L)
- Initial moles:

$$\text{moles of NaOH} = 0.175 \text{ mol/L} \times 0.020 \text{ L} = 0.0035 \text{ mol}$$

Step 2: Reaction Between Nitrous Acid and NaOH

The reaction between the weak acid (HNO_2) and the strong base (NaOH) is a neutralization reaction:



Since NaOH is a strong base, it will react completely with an equivalent amount of HNO_2 .

Moles of reactants after reaction:

- Moles of HNO_2 remaining:

$$0.00675 \text{ mol} - 0.0035 \text{ mol} = 0.00325 \text{ mol}$$

- Moles of NO_2^- formed (conjugate base):

$$0.0035 \text{ mol}$$

Final amounts in the solution:

- Remaining HNO_2 : 0.00325 mol
- Formed NO_2^- : 0.0035 mol

Step 3: Calculating Final Concentrations

The total volume of the solution is the sum of the individual volumes:

$$V_{\text{total}} = 45.0 \text{ mL} + 20.0 \text{ mL} = 65.0 \text{ mL} = 0.065 \text{ L}$$

Concentration of remaining HNO_2 :

$$[\text{HNO}_2] = \frac{0.00325 \text{ mol}}{0.065 \text{ L}} \approx 0.050 \text{ M}$$

Concentration of NO_2^- :

$$[\text{NO}_2^-] = \frac{0.0035 \text{ mol}}{0.065 \text{ L}} \approx 0.054 \text{ M}$$

Step 4: Applying the Henderson-Hasselbalch Equation

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

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

- A^- : conjugate base (NO_2^-)
- HA : weak acid (HNO_2)

Step 4.1: Find pK_a of Nitrous Acid

The pK_a of nitrous acid is approximately 3.37 (from reference data).

Step 4.2: Calculate pH

$$\text{pH} = 3.37 + \log \left(\frac{0.054}{0.050} \right)$$

$$\text{pH} = 3.37 + \log (1.08)$$

$$\text{pH} = 3.37 + 0.033$$

$$\text{pH} \approx 3.40$$

This value indicates the pH of the buffer after the reaction completes.

Additional Considerations and Clarifications

Factors Affecting Buffer pH

- Initial concentrations: Higher concentrations of acid and conjugate base yield more effective buffers with larger capacity.
- Extent of reaction: Complete reaction of NaOH with HNO_2 ensures the formation of a proper buffer system.
- Temperature: pK_a values are temperature-dependent; typical calculations assume standard room temperature ($\sim 25^\circ\text{C}$).

Limitations and Assumptions

- The calculation assumes complete reaction of NaOH with HNO_2 .
- No additional acids or bases are introduced during the process.

- The solution is well-mixed, and activity coefficients are close to 1 (ideal solution).

Practical Applications of Buffer pH Calculation

Understanding how to calculate the pH of a buffer solution is essential in various scenarios:

- Biological systems: Maintaining stable pH in blood and cellular environments.
- Industrial processes: Designing buffers for manufacturing or analytical chemistry.
- Environmental science: Buffering capacity of natural waters against acidification.

Summary of Key Steps

1. Determine initial moles of weak acid and strong base.
2. Calculate moles reacted and remaining after neutralization.
3. Find the concentrations of the acid and conjugate base in the final volume.
4. Apply the Henderson-Hasselbalch equation using known pKa values.
5. Interpret the resulting pH to understand the buffer's characteristics.

Final Thoughts

Calculating the pH of a buffer made by combining nitrous acid and sodium hydroxide involves a clear understanding of acid-base reactions, molarity, and the Henderson-Hasselbalch equation. By carefully following each step—determining initial moles, reactions, final concentrations, and applying the equation—you can accurately predict the pH of complex buffer systems. This process enhances your grasp of fundamental chemistry principles and equips you with practical skills for laboratory and real-world applications.

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