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Understanding how to determine the pH of a buffer solution is fundamental in chemistry, especially when dealing with solutions containing weak acids and their conjugate bases or vice versa. In this article, we will analyze a specific scenario involving the preparation of a buffer solution by combining ammonia (NH_3) and ammonium nitrate (NH_4NO_3). We will guide you through the process of calculating the pH step-by-step, covering all relevant concepts, equations, and calculations necessary to arrive at an accurate answer.

Introduction to Buffer Solutions

Buffer solutions are mixtures that resist changes in pH when small amounts of acids or bases are added. They are typically composed of a weak acid and its conjugate base or a weak base and its conjugate acid. The common principle behind buffer solutions involves the equilibrium between these species, which can absorb or release hydrogen ions (H^+), maintaining a relatively stable pH.

Key points about buffer solutions:

- They contain significant amounts of both a weak acid and its conjugate base, or vice versa.
- The pH of a buffer can be calculated using the Henderson-Hasselbalch equation.
- The effectiveness of a buffer depends on the concentrations and the pKa of the acid-base pair involved.

Understanding the Components of the Buffer Solution

In this scenario, the buffer is prepared by mixing:

- 115 mL of 0.30 M NH_3 (ammonia), a weak base.
- 135 mL of 0.15 M NH_4NO_3 (ammonium nitrate), which provides NH_4^+ ions, the conjugate acid of NH_3.

What does this imply?

- NH_3 acts as the weak base in the buffer.
- NH_4^+ (from NH_4NO_3) acts as the conjugate acid.

This mixture creates a buffer system based on the NH_3/NH_4^+ equilibrium.

Calculating Moles of Each Species

Before applying the Henderson-Hasselbalch equation, we need to determine the number of moles of each component in the final mixture.

Step 1: Calculate moles of NH3

$$
\text{Moles of NH3} = \text{Concentration} \times \text{Volume (in liters)}
$$

$$
= 0.30\, \text{mol/L} \times 0.115\, \text{L} = 0.0345\, \text{mol}
$$

Step 2: Calculate moles of NH4+

$$
= 0.15\, \text{mol/L} \times 0.135\, \text{L} = 0.02025\, \text{mol}
$$

Note: The initial concentrations are given, but since the solutions are mixed, the total moles of each species determine the final concentrations after mixing.

Step 3: Determine the total volume of the mixture

$$
\text{Total volume} = 115\, \text{mL} + 135\, \text{mL} = 250\, \text{mL} = 0.250\, \text{L}
$$

Step 4: Calculate final concentrations

- $\text{Concentration of NH3} = \frac{0.0345\, \text{mol}}{0.250\, \text{L}} = 0.138\, \text{M}$
- $\text{Concentration of NH4+} = \frac{0.02025\, \text{mol}}{0.250\, \text{L}} = 0.081\, \text{M}$

Determining the pKa of the NH3/NH4+ System

The pKa value is crucial for calculating the pH of the buffer. For the ammonium ion ($NH4^+$), the conjugate acid of ammonia, the pKa is approximately 9.25 at 25°C.

Key point:

$$
\text{pKa of NH4+} \approx 9.25
$$

This value is derived from the acid dissociation constant (K_a) of NH4+:

$$
K_a = 10^{-\text{pKa}} = 10^{-9.25} \approx 5.62 \times 10^{-10}
$$

Using the Henderson-Hasselbalch Equation

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

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

In this context:

- The base is NH_3 (ammonia).
- The acid is NH_4^+ (ammonium ion).

Step 1: Plug in the known values:

$$
\text{pH} = 9.25 + \log \left(\frac{0.138}{0.081} \right)
$$

Step 2: Calculate the ratio:

$$
\frac{0.138}{0.081} \approx 1.70
$$

Step 3: Calculate the logarithm:

$$
\log(1.70) \approx 0.23
$$

Step 4: Final calculation:

$$
\text{pH} = 9.25 + 0.23 = 9.48
$$

Result:

The pH of the buffer solution is approximately 9.48.

Factors Affecting the pH Calculation

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

- Temperature variations can alter \(pKa\) values.
- Impurities or additional ions in the solution.
- Activity coefficients, especially at higher ionic strengths.
- Precise measurement of volumes and concentrations.

In typical laboratory settings, these factors are minimized, making the Henderson-Hasselbalch equation a reliable tool for pH estimation.

Summary of the Calculation Steps

Step	Description	Result
1	Calculate moles of NH_3	0.0345 mol
2	Calculate moles of NH_4^+	0.02025 mol
3	Find total volume	0.250 L
4	Final concentrations	NH_3: 0.138 M, NH_4^+: 0.081 M
5	Use pKa = 9.25	—
6	Apply Henderson-Hasselbalch	pH ≈ 9.48

Therefore, the pH of the prepared buffer solution is approximately 9.48.

Conclusion

Understanding how to calculate the pH of a buffer solution prepared from weak bases and their conjugate acids is essential in various fields, including chemistry, biology, and environmental science. By carefully determining the concentrations of the species involved and applying the Henderson-Hasselbalch equation, one can accurately estimate the pH of complex mixtures.

In this specific case, mixing 115 mL of 0.30 M NH_3 with 135 mL of 0.15 M NH_4NO_3 results in a buffer with a pH of approximately 9.48, highlighting the basic nature of this buffer system. Such calculations are invaluable for designing buffer solutions with desired pH values for experiments, industrial processes, or pharmaceutical formulations.

Remember:

- Always verify the pKa value at the specific temperature.
- Use precise measurements for accurate calculations.
- Consider the impact of solution volume and dilution when mixing solutions.

By mastering these concepts, you can confidently analyze and prepare buffer systems tailored to specific pH requirements.

Frequently Asked Questions

How do you determine the pH of a buffer solution prepared by mixing NH3 and NH4NO3?

You use the Henderson-Hasselbalch equation: pH = pKa + log([A-]/[HA]), where A- is the conjugate base (NH3) and HA is the conjugate acid (NH4+).

What is the significance of the concentrations of NH3 and NH4NO3 in calculating the buffer pH?

The concentrations determine the ratio of base to acid in the buffer, which directly influences the pH according to the Henderson-Hasselbalch equation.

How do you find the pKa of the NH4+ / NH3 buffer system?

The pKa of the NH4+ / NH3 system is approximately 9.25 at 25°C, based on the acid dissociation constant (Ka) for NH4+.

What steps are involved in calculating the pH of the buffer after mixing NH3 and NH4NO3?

First, calculate the moles of NH3 and NH4+ in the mixture, then determine their concentrations. Next, apply the Henderson-Hasselbalch equation using the known pKa to find the pH.

How do the volumes of the solutions affect the buffer pH calculation?

Volumes are used to convert concentrations to moles; the total volume after mixing affects the molarity of each species, which is essential for accurate pH calculation.

What is the final pH of the buffer solution prepared by adding 115 mL of 0.30 M NH3 to 135 mL of 0.15 M NH4NO3?

The pH is approximately 9.17, calculated by determining the moles of NH3 and NH4+, adjusting for total volume, and applying the Henderson-Hasselbalch equation with pKa ≈ 9.25.

Additional Resources

A Buffer Is Prepared By Adding 115 ML Of 0.30 M NH_3 To 135 ML Of 0.15 M NH_4NO_3: What Is The pH Of The Solution?

Introduction

Buffer solutions are integral to numerous scientific, industrial, and biological processes due to their ability to resist pH changes upon the addition of small amounts of acids or bases. Understanding and calculating the pH of a prepared buffer solution requires a comprehensive grasp of acid-base chemistry, particularly the principles of conjugate acid-base pairs, molarity, volume adjustments, and the Henderson-Hasselbalch equation.

In this analysis, we investigate a specific buffer preparation involving the addition of ammonia (NH_3) and ammonium nitrate (NH_4NO_3). The question at the core is: What is the pH of the solution when 115 mL of 0.30 M NH_3 is mixed with 135 mL of 0.15 M NH_4NO_3? This scenario exemplifies a classic buffer system involving a weak base (NH_3) and its conjugate acid (NH_4^+) provided by the ammonium ion (NH_4^+) from NH_4NO_3.

Background Concepts and Theoretical Foundations

Acid-Base Equilibria and Buffer Systems

A buffer system typically involves a weak acid and its conjugate base or a weak base and its conjugate acid. In aqueous solutions:

- Ammonia (NH_3) is a weak base capable of accepting protons:

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

- Ammonium ion (NH_4^+), derived from ammonium nitrate (NH_4NO_3), acts as a conjugate acid:

$$NH_4^+ \rightleftharpoons NH_3 + H^+$$

The equilibrium between these species determines the pH of the solution. The Henderson-Hasselbalch equation provides a straightforward calculation method when concentrations of the acid and base are known:

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

Where:

- pK_a is the negative base-10 logarithm of the acid dissociation constant K_a.
- $[\text{Base}]$ and $[\text{Acid}]$ are the molar concentrations of the weak base and conjugate acid, respectively.

The Acid Dissociation Constant of Ammonium Ion

The relevant pK_a value is derived from the base dissociation constant K_b for ammonia:

$$K_b (NH_3) \approx 1.8 \times 10^{-5}$$

Using the relation:

$$pK_a + pK_b = 14$$

we find:

$$ pK_a = 14 - \log K_b = 14 - \log(1.8 \times 10^{-5}) $$

Calculating:

$$ \log(1.8 \times 10^{-5}) \approx -4.74 $$

Thus:

$$ pK_a \approx 14 - (-4.74) = 14 + 4.74 = 18.74 $$

However, this is incorrect because the relation is:

$$ pK_a = 14 - pK_b $$

Given that:

$$ pK_b = -\log K_b = -\log(1.8 \times 10^{-5}) \approx 4.74 $$

So:

$$ pK_a = 14 - 4.74 = 9.26 $$

This is the correct pK_a for NH_4^+/NH_3 system.

Step-by-Step Calculation of the Buffer pH

1. Determining the Moles of NH_3 and NH_4NO_3

- NH_3 (weak base):

- Volume = 115 mL = 0.115 L
- Molarity = 0.30 M

$$
\text{Moles of NH}_3 = 0.30 \, \text{mol/L} \times 0.115 \, \text{L} = 0.0345 \, \text{mol}
$$

- NH_4NO_3 (source of NH_4^+):

- Volume = 135 mL = 0.135 L
- Molarity = 0.15 M

$$
\text{Moles of NH}_4\text{NO}_3 = 0.15 \, \text{mol/L} \times 0.135 \, \text{L} = 0.02025 \, \text{mol}
$$

Since NH_4NO_3 dissociates fully, the moles of NH_4^+ are:

$$
\text{Moles of NH}_4^+ = 0.02025 \, \text{mol}
$$

2. Calculating the Concentrations Post-Mix

Total volume after mixing:

$$
V_{total} = 115\, \text{mL} + 135\, \text{mL} = 250\, \text{mL} = 0.250\, \text{L}
$$

- Concentration of NH_3:

$$
[\text{NH}_3] = \frac{0.0345\, \text{mol}}{0.250\, \text{L}} = 0.138\, \text{M}
$$

- Concentration of NH_4^+:

$$
[\text{NH}_4^+] = \frac{0.02025\, \text{mol}}{0.250\, \text{L}} = 0.081\, \text{M}
$$

3. Applying the Henderson-Hasselbalch Equation

Using:

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

which translates to:

$$
pH = 9.26 + \log \left(\frac{0.138}{0.081} \right)
$$

Calculating the ratio:

$$
\frac{0.138}{0.081} \approx 1.70
$$

and the log:

$$
\log(1.70) \approx 0.23
$$

Finally, the pH:

$$
pH \approx 9.26 + 0.23 = 9.49
$$

Interpretation and Significance

The calculated pH of approximately 9.49 indicates a mildly basic buffer solution, which aligns with expectations given the concentrations of NH_3 and NH_4^+. The presence of excess NH_3 shifts the equilibrium toward the base form, raising the pH above neutral.

This calculation demonstrates the importance of molarity, volume adjustments, and conjugate acid-base relationships in preparing buffers with desired pH levels. Such precision is critical in contexts ranging from biochemical assays to industrial processes where pH stability is paramount.

Factors Influencing Buffer Capacity and pH Stability

While the calculation provides an initial estimate, real-world conditions can influence buffer effectiveness:

- Temperature Dependence: pK_a values vary with temperature changes.
- Ionic Strength: Increased ionic strength can alter activity coefficients.
- Additional Components: Presence of other ions or compounds may shift equilibria.
- Measurement Errors: Precise volumetric and concentration measurements are vital.

Understanding these factors helps chemists design buffers tailored to specific applications.

Broader Implications and Practical Applications

Buffer systems involving ammonia and ammonium salts are common in:

- Biological Systems: Maintaining physiological pH in cell cultures.
- Industrial Processes: pH control in wastewater treatment.
- Laboratory Settings: Titration and analytical chemistry procedures.

The ability to accurately predict and control pH enables scientists to optimize conditions, prevent unwanted reactions, and ensure reproducibility.

Conclusion

The comprehensive analysis reveals that mixing 115 mL of 0.30 M NH_3 with 135 mL of 0.15 M NH_4NO_3 results in a buffer solution with an approximate pH of 9.49. This outcome emphasizes the importance of stoichiometry, equilibrium chemistry, and the Henderson-Hasselbalch equation in preparing and understanding buffer solutions. Accurate pH prediction is essential across various scientific disciplines, underpinning the stability and efficacy of chemical and biological processes.

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