What Mass Of NH4Cl Must Be Added To 0.750 L Of A 0.1M Solution Of NH3, To Give A Buffer Solution With

What Mass Of NH4Cl Must Be Added To 0.750 L Of A 0.1M Solution Of NH3, To Give A Buffer Solution With the desired pH? This question involves understanding the principles of buffer solutions, the relationship between weak acids and their conjugate salts, and applying stoichiometry to determine the required mass of ammonium chloride (NH4Cl). In this article, we will explore the detailed steps to calculate the mass of NH4Cl needed to convert a given ammonia (NH3) solution into a buffer with a specific pH value. We will cover the fundamental concepts, relevant equations, and practical calculations, providing a comprehensive guide for students and chemists alike.

Understanding Buffer Solutions

What Is a Buffer Solution?

A buffer solution is a mixture of a weak acid and its conjugate base (or vice versa) that resists changes in pH when small amounts of acids or bases are added. Buffers are essential in many biological and chemical systems where maintaining a stable pH is critical.

Components of a Buffer Solution

- Weak Acid and Its Conjugate Base: For example, NH4+ (from NH4Cl) and NH3.
- pH Range: The effective pH range of a buffer depends on the pKa of the weak acid.

Why Use NH3 and NH4Cl as a Buffer?

Ammonia (NH_3) is a weak base, and ammonium chloride (NH_4Cl) provides the conjugate acid. Together, they form a buffer system that can resist pH changes in a specific range, typically around the pKa of the NH_4^+/NH_3 system.

Fundamental Concepts and Equations

Henderson-Hasselbalch Equation

The primary equation used to determine the pH of a buffer solution is:

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

Where:

- $[\text{A}^-]$ is the concentration of the conjugate base (NH_3 here).
- $[\text{HA}]$ is the concentration of the weak acid (NH_4^+ here).
- $\text{p}K_a$ is the negative logarithm of the acid dissociation constant.

Relevant Constants

- For the ammonium ion (NH_4^+), $\text{p}K_a$ is approximately 9.25.
- Initial concentration of NH_3: 0.1 M.
- Volume of solution: 0.750 L.

Steps to Calculate the Required Mass of NH4Cl

1. Determine the desired pH of the buffer.

2. Use the Henderson-Hasselbalch equation to find the ratio $[\text{A}^-]/[\text{HA}]$.

3. Calculate the molar amounts of NH3 and NH4+ needed.

4. Determine the additional amount of NH4+ (from NH4Cl) required.

5. Convert moles of NH4+ to mass of NH4Cl.

Step-by-Step Calculation Process

Step 1: Determine the Target pH

Suppose the desired pH of the buffer is 9.0, a common pH for ammonia buffers.

Step 2: Apply Henderson-Hasselbalch Equation

Given:

- $\text{p}K_a$ = 9.25

- Desired pH = 9.0

Calculate the ratio:

$$ 9.0 = 9.25 + \log \left(\frac{[\text{A}^-]}{[\text{HA}]} \right) $$

$$ \log \left(\frac{[\text{A}^-]}{[\text{HA}]} \right) = 9.0 - 9.25 = -0.25 $$

$$ \frac{[\text{A}^-]}{[\text{HA}]} = 10^{-0.25} \approx 0.562 $$

This means:

$$ \frac{[\text{NH}_3]}{[\text{NH}_4^+]} \approx 0.562 $$

Step 3: Calculate Moles of NH_3 in the Original Solution

Initial concentration of NH_3: 0.1 M

Volume: 0.750 L

Moles of NH_3:

$$ n_{NH_3} = 0.1\, \text{mol/L} \times 0.750\, \text{L} = 0.075\, \text{mol} $$

Since the NH_3 is initially at 0.1 M, the total moles are 0.075 mol.

Step 4: Determine Moles of NH_4^+ Needed

Using the ratio:

$$ \frac{[\text{NH}_3]}{[\text{NH}_4^+]} \approx 0.562 $$

Rearranged:

$$ [\text{NH}_4^+] = \frac{[\text{NH}_3]}{0.562} $$

Total moles of NH_3 in the solution:

$$ 0.075\, \text{mol} $$

Thus:

$$ n_{NH_4^+} = \frac{0.075\, \text{mol}}{0.562} \approx 0.133\, \text{mol} $$

Note: The initial NH_3 contributes 0.075 mol, but to achieve the desired pH, an additional amount of NH_4^+ must be added to reach the calculated value of approximately 0.133 mol.

Step 5: Calculate the Additional Moles of NH4Cl Required

Additional moles of NH4+ needed:

$$ n_{NH_4^+\, \text{additional}} = 0.133\, \text{mol} - 0.075\, \text{mol} = 0.058\, \text{mol} $$

Step 6: Convert Moles of NH4Cl to Mass

Molecular weight of NH4Cl:

- N: 14.01 g/mol
- H: 1.008 g/mol (4 H atoms)
- Cl: 35.45 g/mol

Calculate:

$$ \text{Molecular weight of NH4Cl} = 14.01 + (4 \times 1.008) + 35.45 \approx 53.49\, \text{g/mol} $$

Mass of NH4Cl required:

$$ m_{NH_4Cl} = 0.058\, \text{mol} \times 53.49\, \text{g/mol} \approx 3.1\, \text{g} $$

Final Answer and Practical Considerations

To convert 0.750 L of a 0.1 M NH3 solution into a buffer with a pH of approximately 9.0, you need to add about 3.1 grams of NH4Cl.

Practical Tips:

- Always verify the target pH before calculating the required salt.
- Ensure thorough mixing of the solution after adding NH4Cl.

- Adjust the pH slightly if necessary by adding small amounts of acid or base.
- Use analytical balances for precise measurement of NH_4Cl.

Additional Factors to Consider

Effect of Temperature

The $\text{p}K_a$ value for NH_4^+ can vary slightly with temperature, affecting the calculations.

Buffer Capacity

The buffer capacity depends on the total concentrations of NH_3 and NH_4^+. Higher concentrations provide greater resistance to pH changes.

Limitations of the Calculation

- Assumes ideal behavior.
- Does not account for activity coefficients.
- Real-world systems may require adjustments.

Summary

- Identify the desired pH and relevant $\text{p}K_a$.
- Use the Henderson-Hasselbalch equation to find the ratio of base to acid.

- Calculate the moles of NH3 and NH4+ needed.
- Determine the additional moles of NH4+ from NH4Cl.
- Convert moles to mass considering molar mass.

This systematic approach ensures accurate preparation of buffer solutions tailored to specific pH requirements, essential in laboratory experiments, pharmaceutical formulations, and chemical processes.

References

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By understanding the principles outlined above, you can confidently determine the mass of NH4Cl required for various buffer preparation scenarios

Frequently Asked Questions

What is the purpose of adding NH4Cl to an NH3 solution?

Adding NH4Cl introduces ammonium ions (NH4+), which react with NH3 to create a buffer solution that resists pH changes.

How do you determine the mass of NH_4Cl needed to prepare a buffer with a specific pH?

Using the Henderson-Hasselbalch equation, you calculate the required molar ratio of NH_4^+ to NH_3 and then find the corresponding mass of NH_4Cl based on its molar mass.

What is the molar mass of NH_4Cl used in calculations?

The molar mass of NH_4Cl is approximately 53.5 g/mol.

How many moles of NH_3 are present in 0.750 L of a 0.1 M solution?

Number of moles = molarity × volume = 0.1 mol/L × 0.750 L = 0.075 mol.

What is the target pH of the buffer solution in this problem?

The question is incomplete, but typically, buffer solutions involving NH_3/NH_4^+ aim for a pH around 9.25, which is the pKa of the NH_4^+/NH_3 system.

How do you calculate the amount of NH_4Cl needed to reach the desired buffer pH?

Use the Henderson-Hasselbalch equation: $pH = pKa + \log([A^-]/[HA])$, determine the required molar ratio, and then find the moles of NH_4Cl needed to provide that ratio relative to NH_3.

What assumptions are made in calculating the required mass of NH_4Cl?

Assumptions include complete dissociation of NH_4Cl into NH_4^+ ions, no volume change upon addition, and that the initial NH_3 concentration remains unchanged initially.

Can you provide a step-by-step calculation for the mass of NH4Cl needed?

Yes. First, determine the desired [NH4+]/[NH3] ratio using pH and pKa, then calculate moles of NH4+ required, and finally convert moles to grams using molar mass (53.5 g/mol).

What factors can affect the accuracy of preparing this buffer solution?

Factors include measurement precision, purity of chemicals, temperature fluctuations affecting pKa, and assuming ideal behavior in calculations.

Why is it important to prepare buffer solutions with precise amounts of NH4Cl and NH3?

Precise amounts ensure the buffer maintains the desired pH stability, which is critical in many chemical and biological applications.

Additional Resources

What Mass Of NH_4Cl Must Be Added To 0.750 L Of A 0.1M Solution Of NH_3 To Give A Buffer Solution With Desired pH

Creating a buffer solution involves a delicate balance between a weak base and its conjugate acid. In this context, the weak base is ammonia (NH_3), and its conjugate acid is ammonium chloride (NH_4Cl). To determine the precise amount of NH_4Cl needed to convert a given amount of NH_3 into a buffer with a specific pH, one must understand the underlying principles of buffer chemistry, the relevant equilibrium expressions, and perform careful calculations.

This comprehensive review explores the theoretical foundations, detailed calculations, practical considerations, and implications involved in preparing such a buffer solution.

Understanding Buffer Solutions and Their Significance

Buffer solutions are aqueous mixtures that resist changes in pH upon the addition of small amounts of acids or bases. They are essential in many biological, industrial, and laboratory processes where maintaining a stable pH is crucial.

Key features of buffer solutions include:

- Comprising a weak acid and its conjugate base or a weak base and its conjugate acid.
- Capable of neutralizing added acids or bases within certain limits.
- Characterized by their buffer capacity, which depends on the concentrations of the buffering components.

In this specific scenario, the buffer is formed from ammonia (NH_3) and ammonium chloride (NH_4Cl). This combination leverages the equilibrium between ammonia and ammonium ions.

Fundamental Chemistry Principles

Ammonia – The Weak Base

NH_3 is a weak base that undergoes hydrolysis in water:

$$\mathrm{NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-}$$

The equilibrium constant associated with this process is the base dissociation constant, K_b.

Ammonium Ion – The Conjugate Acid

NH_4^+, the conjugate acid of ammonia, also participates in hydrolysis:

$$\mathrm{NH_4^+ + H_2O \leftrightarrow NH_3 + H_3O^+}$$

Its dissociation is characterized by the acid dissociation constant, K_a.

Relationship between K_a and K_b

Because NH_3 and NH_4^+ are conjugates:

$$K_a \times K_b = K_w = 1.0 \times 10^{-14}$$

Given that:

- K_b for NH_3 is approximately 1.8×10^{-5},
- Therefore, K_a for NH_4^+ is:

$$K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} \approx 5.56 \times 10^{-10}$$

Calculating the Required Amount of NH_4Cl

The goal is to add a specific mass of NH_4Cl to a 0.750 L of 0.1 M NH_3 solution to produce a buffer with a desired pH. To approach this, we need to:

1. Understand the initial conditions.
2. Determine the relationship between the buffer's pH and the ratio of $[\mathrm{NH_4^+}]$ to $[\mathrm{NH_3}]$.
3. Use the Henderson-Hasselbalch equation to relate pH to concentrations.
4. Calculate the moles of NH_4Cl required.
5. Convert moles to mass.

Step 1: Initial Conditions

- Volume of NH_3 solution: $V_{NH_3} = 0.750\, \text{L}$
- Concentration of NH_3: $C_{NH_3} = 0.1\, \text{M}$

Number of moles of NH_3 initially present:

$$
n_{NH_3} = C_{NH_3} \times V_{NH_3} = 0.1\, \text{mol/L} \times 0.750\, \text{L} = 0.075\, \text{mol}
$$

Step 2: Buffer pH and the Henderson-Hasselbalch Equation

The pH of the buffer will depend on the ratio of conjugate base to conjugate acid:

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

In this case:

- $[\mathrm{A^-}]$ corresponds to $[\mathrm{NH_3}]$,
- $[\mathrm{HA}]$ corresponds to $[\mathrm{NH_4^+}]$.

Rearranged for the ratio:

$$
\frac{[\mathrm{NH_4^+}]}{[\mathrm{NH_3}]} = 10^{pH - pK_a}
$$

Suppose the target pH is specified; if not, an example pH (e.g., 9.25) can be used based on typical buffer ranges.

For illustration, assume a desired pH of 9.25, which is common for ammonia buffers.

Since:

$$
pK_a \approx 9.25
$$

then:

$$
\frac{[\mathrm{NH_4^+}]}{[\mathrm{NH_3}]} = 10^{9.25 - 9.25} = 10^0 = 1
$$

This implies that:

$$
[\mathrm{NH_4^+}] = [\mathrm{NH_3}] \Rightarrow n_{NH_4^+} / V_{total} = n_{NH_3} / V_{initial}
$$

Because the initial volume is 0.750 L, and the total volume after adding NH_4Cl remains approximately the same (assuming negligible volume change), the moles of NH_4Cl added should produce an equivalent molar concentration of NH_4^+:

$$
[\mathrm{NH_4^+}] = \frac{n_{NH_4^+}}{V_{total}}
$$

Given the initial moles of NH_3 are 0.075 mol, the number of moles of NH_4Cl (which fully dissociates into NH_4^+) needed:

$$
n_{NH_4Cl} = n_{NH_4^+} = 0.075\, \text{mol}
$$

Step 3: Calculating the Mass of NH_4Cl

Molar mass of NH_4Cl:

$$
M_{NH_4Cl} = 14.01\, (\mathrm{N}) + 4 \times 1.008\, (\mathrm{H}) + 35.45\, (\mathrm{Cl}) \approx 53.49\, \text{g/mol}
$$

Mass required:

$$
m_{NH_4Cl} = n_{NH_4Cl} \times M_{NH_4Cl} = 0.075\, \text{mol} \times 53.49\, \text{g/mol} \approx 4.01\, \text{g}
$$

Therefore, approximately 4.01 grams of NH_4Cl need to be added to the solution to achieve the buffer with pH 9.25.

Adjustments for Different pH Values

The previous calculation hinges on the assumption that the pH is 9.25, which aligns with pK_a. If the target pH differs, the ratio of $[\mathrm{NH_4^+}]$ to $[\mathrm{NH_3}]$ changes accordingly.

For example:

- Target pH = 9.00:

$$
\frac{[\mathrm{NH_4^+}]}{[\mathrm{NH_3}]} = 10^{9.00 - 9.25} \approx 10^{-0.25} \approx 0.56
$$

- Target pH = 9.50:

$$
\frac{[\mathrm{NH_4^+}]}{[\mathrm{NH_3}]} = 10^{0.25} \approx 1.78
$$

Correspondingly, the moles of NH_4Cl needed would be:

$$
n_{NH_4Cl} = \text{ratio} \times n_{NH_3}
$$

And mass calculated similarly.

Key takeaway: The amount of NH_4Cl is directly proportional to the desired pH and the initial moles of NH_3.

Practical Considerations and Limitations

1. Volume Changes:

Adding solid NH_4Cl will slightly increase the total volume, affecting concentration calculations. For most practical purposes, this volume change is negligible, but for precise work, it should be considered.

2. Solubility Limitations:

NH_4Cl is highly soluble in water (~37 g/100 mL at room temperature). The

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