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Understanding chemical reactions is fundamental in chemistry, especially when it comes to calculating the quantities of reactants involved. In this article, we will explore a specific reaction involving a weak base, Chn, and a strong acid, HCl, to determine the amount of reactants that react completely or partially. The problem involves calculating the extent of reaction, the remaining reactants, or the amount of products formed based on given molarities and volumes.

Introduction to Acid-Base Reactions

Acid-base reactions are among the most common reactions in chemistry. They involve the transfer of protons (H^+) from acids to bases. Typically, strong acids like HCl dissociate completely in solution, while weak bases like Chn (a generic weak base) do not. The key to solving such problems lies in understanding molarity, molar ratios, and the concept of equilibrium for weak bases.

Understanding the Reactants

- Chn (a weak base):
- Volume: 75.0 mL
- Molarity: 0.350 M
- Base dissociation constant (Kb): 1.7×10^{-7}

- Hydrochloric acid (HCl):
- Volume: 100.0 mL
- Molarity: 0.425 M

Calculating the Moles of Reactants

Before analyzing the reaction, we need to determine the number of moles of each reactant based on their given molarity and volume.

Calculating Moles of Chn

$$
\text{Moles of Chn} = Molarity \times Volume (L)
$$

Convert volume from mL to L:

$$
75.0\, \text{mL} = 0.075\, \text{L}
$$

Calculate:

$$
\text{Moles of Chn} = 0.350\, \text{mol/L} \times 0.075\, \text{L} = 0.02625\, \text{mol}
$$

Calculating Moles of HCl

Convert volume from mL to L:

$$
100.0\, \text{mL} = 0.100\, \text{L}
$$

Calculate:

$$
\text{Moles of HCl} = 0.425\, \text{mol/L} \times 0.100\, \text{L} = 0.0425\, \text{mol}
$$

Reaction Overview

The reaction between a weak base (Chn) and a strong acid (HCl) can be represented as:

$$
\text{Chn} + \text{HCl} \rightarrow \text{ChnH}^+ + \text{Cl}^-
$$

Since HCl is a strong acid, it dissociates completely, and Chn reacts with HCl to form its conjugate acid, $ChnH^+$.

Reaction Stoichiometry

- 1 mole of Chn reacts with 1 mole of HCl.
- The limiting reactant determines the extent of the reaction.

Determining the Limiting Reactant

Compare the number of moles:

- Chn: 0.02625 mol
- HCl: 0.0425 mol

Since 0.02625 mol of Chn requires 0.02625 mol of HCl to react completely, and we have 0.0425 mol HCl, HCl is in excess.

Conclusion:
Chn is the limiting reactant; it will be fully consumed, and some HCl will remain unreacted.

Calculating the Extent of Reaction

- Moles of Chn reacted: Equal to initial moles, 0.02625 mol.
- Moles of HCl reacted: Same as Chn, 0.02625 mol.

Remaining HCl:

$$
0.0425\, \text{mol} - 0.02625\, \text{mol} = 0.01625\, \text{mol}
$$

Remaining HCl volume:

$$
\text{Volume of remaining HCl} = \frac{0.01625\, \text{mol}}{0.425\, \text{mol/L}} \approx 0.0382\, \text{L} = 38.2\, \text{mL}
$$

Remaining HCl in solution:

- Volume: 38.2 mL
- Molarity: still approximately 0.425 M (assuming no volume change)

Calculating the Concentration of Conjugate Acid

The conjugate acid, $ChnH^+$, produced:

$$
\text{Moles of ChnH}^+ = 0.02625\, \text{mol}
$$

Total volume after mixing:

$$
75.0\, \text{mL} + 100.0\, \text{mL} = 175.0\, \text{mL} = 0.175\, \text{L}
$$

Concentration of $ChnH^+$:

$$
\text{Concentration} = \frac{0.02625\, \text{mol}}{0.175\, \text{L}} \approx 0.15\, \text{M}
$$

pH and pOH Calculations

Understanding the pH of the solution requires considering the weak base equilibrium and the conjugate acid formation.

Calculating the pH of the Resulting Solution

Since the limiting reactant is Chn, and a conjugate acid is formed, the solution contains a mixture of unreacted HCl and $ChnH^+$. To determine the pH:

- Consider the concentration of free HCl remaining.
- Use the Henderson-Hasselbalch equation if needed.

Remaining HCl:

$$
\text{Concentration} = \frac{0.01625\, \text{mol}}{0.175\, \text{L}} \approx 0.093\, \text{M}
$$

This is a strong acid; thus, the pH will be dominated by this residual HCl.

Calculate pH:

$$
\text{pH} = -\log [\text{H}^+]
$$

$$
[\text{H}^+] \approx 0.093\, \text{M}
$$

$$
\text{pH} \approx -\log(0.093) \approx 1.03
$$

The solution is acidic with a pH around 1.03.

Summary of Key Findings

- Limiting Reactant: Chn (0.02625 mol)
- Reacted HCl: 0.02625 mol
- Remaining HCl: approximately 0.01625 mol (~38.2 mL at 0.425 M)
- Conjugate acid ($ChnH^+$) formed: 0.02625 mol
- Final pH: approximately 1.03 due to residual HCl

Practical Applications and Significance

Understanding such calculations has real-world applications:

- Pharmaceuticals: Precise titrations for formulation.

- Environmental Chemistry: Assessing acid neutralization.
- Industrial Processes: Controlling reactant quantities for desired outcomes.

Common Challenges in Acid-Base Calculations

- Correctly identifying limiting reactants.
- Converting units accurately.
- Understanding weak acid/base equilibria.
- Calculating pH in mixed solutions with multiple species.

Conclusion

Calculating the quantitative aspects of acid-base reactions requires careful analysis of molarity, volume, and reaction stoichiometry. In the scenario of 75.0 mL of 0.350 M Chn reacting with 100.0 mL of 0.425 M HCl, we determined that Chn is the limiting reactant, and some HCl remains unreacted post-reaction. The residual acidity influences the final pH, which is approximately 1.03, emphasizing the importance of precise measurements in chemical reactions. Mastery of these calculations is essential for chemists working in research, industry, and environmental sciences to ensure accurate and efficient processes.

Keywords: acid-base reaction, molarity, limiting reactant, weak base, strong acid, titration, conjugate acid, pH calculation, stoichiometry, chemical reaction analysis

Frequently Asked Questions

What is the balanced chemical equation for the reaction between chloramine (Chn) and hydrochloric acid (HCl)?

The reaction involves chloramine (Chn) reacting with hydrochloric acid to produce nitrogenous compounds and other products. A simplified form is: Chn + HCl → products. The exact balanced equation depends on the specific form of chloramine, but typically, chloramine reacts with HCl to release nitrogen and form chlorides.

How do you calculate the moles of chloramine (Chn) and HCl used in the reaction?

Calculate moles using molarity and volume: for Chn, M = 0.350 M and V = 75.0 mL (0.075 L); moles = 0.350 mol/L × 0.075 L = 0.02625 mol. For HCl, M = 0.425 M and V = 100.0 mL (0.100 L); moles = 0.425 mol/L × 0.100 L = 0.0425 mol.

What is the limiting reagent in this reaction, and how can you determine it?

Compare the molar amounts of reactants: Chn (0.02625 mol) and HCl (0.0425 mol). Since Chn has fewer moles, it is the limiting reagent. The limiting reagent determines the maximum amount of product formed.

How do you calculate the quantity of nitrogen gas (N_2) produced from this reaction?

Assuming a reaction where each mole of Chn produces a certain amount of N_2, use the molar ratio from the balanced equation. For example, if 1 mol of Chn produces 1 mol of N_2, then moles of N_2 = 0.02625 mol. Convert to volume at STP: volume = moles × 22.4 L/mol = 0.02625 × 22.4 ≈ 0.588 L.

What are the key factors affecting the reaction's extent and the final quantities of products?

Factors include reactant concentrations, temperature, reaction time, and the presence of catalysts. The limiting reagent also influences the maximum amount of products formed. Additionally, reaction conditions such as pH and pressure can impact reaction efficiency.

How can the reaction be optimized to maximize nitrogen gas production in a laboratory setting?

Optimize by ensuring complete reaction of limiting reagent, maintaining appropriate temperature and pH conditions, and allowing sufficient reaction time. Using excess of one reactant carefully, controlling reaction environment, and stirring can also improve yield and efficiency.

Additional Resources

Chemical Reaction Analysis: An In-Depth Examination of Chn and HCl Interaction

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Introduction: Understanding the Reaction Dynamics

In the realm of chemistry, reactions between acids and bases are fundamental, often serving as foundational examples of chemical processes. The scenario under consideration involves the reaction of 75.0 mL of 0.350 M Chn (a hypothetical or unspecified compound with a base dissociation constant, \(K_b = 1.7 \times 10^{-5}\)) with 100.0 mL of 0.425 M HCl. This setup is a classic acid-base neutralization, but the specifics—volumes, concentrations, and equilibrium constants—add layers of complexity that require careful analysis.

This article aims to dissect this reaction comprehensively, exploring the stoichiometry, the extent of reaction, and the resulting chemical quantities.

Whether you're a student, educator, or enthusiast, this detailed review will illuminate the steps and considerations involved.

Understanding the Components: Chn and HCl

1. The Nature of Chn

Although "Chn" appears to be a placeholder, for our analysis, we will treat it as a weak base with a known base dissociation constant:

$$
K_b = 1.7 \times 10^{-5}
$$

This value indicates a weak base, which does not completely dissociate in water. Its behavior with acids like HCl depends on its base strength, which influences the extent of neutralization.

Properties of Chn:

- Volume: 75.0 mL (0.075 L)
- Concentration: 0.350 M
- Moles of Chn:

$$
n_{Chn} = M \times V = 0.350\, \text{mol/L} \times 0.075\, \text{L} = 0.02625\, \text{mol}
$$

Significance: The number of moles provides the initial quantity of base available for reaction.

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2. The Nature of HCl

Hydrochloric acid (HCl) is a strong acid, meaning it dissociates completely in aqueous solution:

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

- Volume: 100.0 mL (0.10 L)
- Concentration: 0.425 M
- Moles of HCl:

$$
n_{HCl} = 0.425\, \text{mol/L} \times 0.10\, \text{L} = 0.0425\, \text{mol}
$$

Implication: Since HCl is a strong acid, it will readily react with the weak base Chn, potentially leading to complete or partial neutralization depending on the quantities involved.

Stoichiometric Considerations: Reaction Equation & Molar Ratios

1. Establishing the Reaction Equation

The neutralization between a weak base (Chn) and a strong acid (HCl) can be represented generally as:

$$
\text{Chn} + \text{HCl} \rightarrow \text{ChnH}^+ + \text{Cl}^-
$$

Since Chn is a weak base, it accepts a proton:

$$
\text{Chn} + \text{H}^+ \rightarrow \text{ChnH}^+
$$

Note: The exact chemical nature of Chn is unspecified, but the key is that it reacts in a 1:1 molar ratio with HCl.

Stoichiometric ratio:

$$
\boxed{
\text{1 mol Chn} : \text{1 mol HCl}
}
$$

---

2. Calculating the Limiting Reactant

Compare the moles:

Reactant	Moles
Chn	0.02625 mol
HCl	0.0425 mol

Since 0.02625 mol of Chn requires an equal amount of HCl for complete neutralization, and there are 0.0425 mol of HCl available, Chn is the limiting reactant.

Conclusion: The reaction will consume all the Chn, leaving some HCl unreacted.

Reaction Progress and Extent

1. Moles of HCl Reacted

- Moles of HCl that react:

$$
n_{HCl,reacted} = n_{Chn} = 0.02625\, \text{mol}
$$

- Moles of HCl remaining:

$$
n_{HCl,remaining} = 0.0425 - 0.02625 = 0.01625\, \text{mol}
$$

Implication: After the reaction, all Chn has been neutralized, forming an equivalent amount of conjugate acid ($ChnH^+$).

2. Quantities of Products and Unreacted Reagents

- Conjugate acid ($ChnH^+$) formed: 0.02625 mol
- Unreacted HCl: 0.01625 mol

Calculating the Final Solution Composition

1. Total Volume of the Mixture

Since volumes are additive:

$$
V_{total} = 75.0\, \text{mL} + 100.0\, \text{mL} = 175.0\, \text{mL} = 0.175\, \text{L}
$$

2. Concentration of Reactants and Products After Reaction

- Conjugate acid ($ChnH^+$):

$$

$$
[M]_{ChnH^+} = \frac{0.02625\, \text{mol}}{0.175\, \text{L}} \approx 0.15\, \text{M}
$$

- Remaining HCl:

$$
[M]_{HCl,remaining} = \frac{0.01625\, \text{mol}}{0.175\, \text{L}} \approx 0.093\, \text{M}
$$

- Unreacted Chn:

$$
n_{Chn,remaining} = 0\, \text{mol} \quad \text{(completely reacted)}
$$

---

pH and Acid-Base Equilibrium Considerations

1. pH of the Final Solution

The pH depends on the remaining concentrations of unreacted HCl and the conjugate acid.

a. Contribution from unreacted HCl:

$$
\text{H}^+ \text{ from residual HCl} = 0.093\, \text{M}
$$
$$
\Rightarrow pH_{HCl} = -\log(0.093) \approx 1.03
$$

b. Contribution from conjugate acid (ChnH$^+$):

- Since ChnH$^+$ is a weak acid (being a conjugate acid of a weak base), its K_a can be estimated:

$$
K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-5}} \approx 5.88 \times 10^{-10}
$$

- The concentration of ChnH$^+$ is 0.15 M, which is significant enough to influence pH.

- Assuming complete dissociation is unlikely for such a weak acid, but for an approximation, the pH contributed by ChnH$^+$:

$$
pH_{ChnH^+} \approx -\frac{1}{2} \log(K_a \times \text{initial concentration}) \quad \text{(approximation)}
$$

However, since the conjugate acid is weak and present at 0.15 M, and the residual HCl is also strong, the solution will be acidic, likely dominated by the unreacted HCl.

Final pH estimate: Approximately 1.0–1.2, primarily influenced by the residual HCl.

Summary of Key Quantities

- Initial moles of Chn: 0.02625 mol
- Initial moles of HCl: 0.0425 mol
- Moles of HCl reacted: 0.02625 mol
- Unreacted HCl remaining: 0.01625 mol
- Moles of conjugate acid formed: 0.02625 mol
- Final solution volume: 0.175 L
- Concentrations:

Species	Concentration (M)
$ChnH^+$	0.15
HCl (remaining)	0.093

Implications and Practical Considerations

This detailed analysis reveals that in the reaction between the specified amounts of Chn and HCl:

- The weak base

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