A 100.0 ML Sample Of 0.10 M NH3 Is Titrated With 0.10 M HNO3. Determine The PH Of The Solution After

A 100.0 ML Sample Of 0.10 M NH3 Is Titrated With 0.10 M HNO3. Determine The PH Of The Solution After this titration process involves understanding acid-base chemistry, calculating the equivalence point, and applying principles of pH calculation for weak acids and bases. This comprehensive guide will walk you through the entire process, from initial concepts to detailed calculations, ensuring you grasp every step involved in determining the pH of the solution after titration.

Understanding the Titration of Ammonia (NH3) with Nitric Acid (HNO3)

Titration is a fundamental laboratory technique used to determine the concentration of an unknown solution by adding a titrant of known concentration until the reaction reaches its equivalence point. In this case, we are titrating ammonia, a weak base, with nitric acid, a strong acid.

Key Concepts in Acid-Base Titration

- **Weak Base:** Ammonia (NH3) is a weak base that partially reacts with water to produce hydroxide ions (OH^-).
- **Strong Acid:** Nitric acid (HNO3) is a strong acid that completely dissociates in water, providing H^+ ions.
- **Equivalence Point:** The point during titration where moles of acid equal moles of base, resulting in a neutralized solution.
- **pH at Different Stages:** Before equivalence, the solution contains excess base or acid; at equivalence, the solution is neutral or slightly basic/acidic depending on the species present; after equivalence, excess titrant determines the pH.

Initial Conditions and Data

Let's begin by summarizing the initial data:

- Volume of ammonia solution (NH_3): 100.0 mL (0.100 L)
- Concentration of ammonia: 0.10 M
- Concentration of nitric acid (HNO_3): 0.10 M

The total moles of ammonia initially present:

$$
\text{Moles of NH}_3 = \text{Concentration} \times \text{Volume} = 0.10\, \text{mol/L} \times 0.100\, \text{L} = 0.010\, \text{mol}
$$

The moles of HNO_3 added depend on the volume added during titration, but since the problem asks to determine the pH after a certain point, we'll analyze specific stages of titration.

Step-by-Step Calculation of pH During Titration

The titration process can be divided into three primary stages:

1. Before the equivalence point (less HNO_3 added)
2. At the equivalence point
3. After the equivalence point (more HNO_3 added)

Each stage requires different calculations based on the species present in solution.

Stage 1: Before the Equivalence Point

In this stage, less nitric acid has been added than needed to neutralize all ammonia. The solution contains:

- Unreacted NH_3 (weak base)
- Its conjugate acid, NH_4^+ (formed in the reaction)

Reaction:

$$
\text{NH}_3 + \text{H}^+ \rightarrow \text{NH}_4^+
$$

Calculating Moles of HNO3 added:

Suppose we add a volume V_{HNO_3} of 0.10 M HNO3. The moles of HNO3 added:

$$
\text{Moles HNO}_3 = 0.10\, \text{mol/L} \times V_{HNO_3} \text{(in liters)}
$$

The moles of ammonia remaining after reaction:

$$
\text{Moles of NH}_3\, \text{remaining} = 0.010\, \text{mol} - \text{moles HNO}_3
$$

Similarly, the moles of ammonium ions formed:

$$
\text{Moles of NH}_4^+ = \text{moles HNO}_3
$$

pH Calculation:

The solution's pH depends on the hydrolysis of ammonium ions:

$$
\text{NH}_4^+ + \text{H}_2O \rightleftharpoons \text{NH}_3 + \text{H}^+
$$

The K_a for NH4+ can be derived from the K_b of NH3:

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

Given:

$$
K_b\, \text{of NH}_3 = 1.8 \times 10^{-5}
$$

Calculate K_a:

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

Using the Henderson-Hasselbalch equation:

$$
\text{pH} = pK_a + \log \left(\frac{[\text{NH}_3]}{[\text{NH}_4^+]} \right)
$$

\]

Where:

- \([\text{NH}_3]\) and \([\text{NH}_4^+]\) are concentrations after titration at a given volume.

Stage 2: At the Equivalence Point

The equivalence point occurs when:

\[
\text{Moles of HNO}_3 = \text{Moles of NH}_3 = 0.010\, \text{mol}
\]

The total volume of the solution:

\[
V_{total} = 100.0\, \text{mL} + V_{HNO_3}
\]

Since the concentrations of the titrant and analyte are equal (both 0.10 M), the volume of HNO3 needed:

\[
V_{HNO_3} = \frac{0.010\, \text{mol}}{0.10\, \text{mol/L}} = 0.100\, \text{L} = 100\, \text{mL}
\]

Total volume:

\[
V_{total} = 100\, \text{mL} + 100\, \text{mL} = 200\, \text{mL} = 0.200\, \text{L}
\]

What species are present?

- Complete neutralization of ammonia to ammonium ions
- The solution contains 0.005 mol of NH4+ in 0.200 L:

\[
[\text{NH}_4^+] = \frac{0.005\, \text{mol}}{0.200\, \text{L}} = 0.025\, \text{M}
\]

pH at the equivalence point:

Since NH4+ hydrolyzes, it makes the solution slightly acidic:

$$
\text{NH}_4^+ + \text{H}_2O \rightleftharpoons \text{NH}_3 + \text{H}^+
$$

The K_a of NH4+ is 5.56×10^{-10}.

Using the hydrolysis expression:

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

Assuming $[\text{H}^+]$ is small and equal to x:

$$
K_a = \frac{x^2}{[\text{NH}_4^+]} \Rightarrow x = \sqrt{K_a \times [\text{NH}_4^+]} = \sqrt{5.56 \times 10^{-10} \times 0.025}
$$

Calculating:

$$
x \approx \sqrt{1.39 \times 10^{-11}} \approx 3.73 \times 10^{-6}\, \text{M}
$$

pH:

$$
\text{pH} = -\log [\text{H}^+] = -\log (3.73 \times 10^{-6}) \approx 5.43
$$

Thus, the pH at the equivalence point is approximately 5.43, indicating a slightly acidic solution.

Stage 3: After the Equivalence Point

When more HNO3 is added beyond the equivalence point, the solution contains excess H+ ions, making it acidic.

Suppose an additional volume V_{extra} of 0.10 M HNO3 is added, resulting in:

$$
\text{Extra moles of HNO}_3 = 0.10\, \text{mol/L} \times V_{extra}
$$

Total moles of HNO3:

$$
\text{Total HNO}_3 = 0.010\, \text{mol} + \text{extra}
$$

Remaining H+ ions:

$$
[\text{H
$$

Frequently Asked Questions

What is the initial pH of the 0.10 M NH3 solution before titration?

The initial pH can be calculated using the hydrolysis of NH3: pOH = -log[OH−] = -log(0.10), so pOH ≈ 1.00. Therefore, pH = 14 - 1.00 = 13.00.

How do you determine the pH after adding some volume of HNO3 during titration?

You first calculate the moles of NH3 and HNO3 before and after addition, identify the limiting reagent, and then determine the remaining species in solution to find the pH based on the resulting equilibrium.

What is the equivalence point in this titration between NH3 and HNO3?

The equivalence point occurs when moles of HNO3 added equal the initial moles of NH3, which is 0.010 mol (since 0.10 M 0.100 L).

How do you calculate the pH at the equivalence point of titrating NH3 with HNO3?

At equivalence, only the salt NH4+ is present, which hydrolyzes to produce H+ ions. The pH is determined from the hydrolysis constant of NH4+, resulting in a slightly acidic pH, typically around 5.5.

What is the pH after adding half the volume of HNO3 needed to reach equivalence?

At half equivalence, the solution contains equal moles of NH3 and NH4+, creating a buffer. The pH can be calculated using the Henderson-Hasselbalch equation: pH = pKa + log([base]/[acid]).

How do you find the pH after adding less than the equivalence volume of HNO3?

Remaining NH3 will dominate, and the pH can be calculated from the concentration of unreacted NH3, considering its hydrolysis. The solution remains basic until near equivalence.

What is the pH after adding more HNO3 than the amount needed to reach equivalence?

Beyond the equivalence point, excess HNO3 determines the pH. The pH can be calculated from the concentration of remaining H+ ions in excess of the titrant.

How does the titration curve of NH3 with HNO3 look like, and what are its key features?

The titration curve starts at high pH (~13), gradually decreases, has a steep drop near the equivalence point (~pH 5.5), and then levels off at low pH as excess acid dominates. Key features include the initial pH, buffering region, equivalence point, and post-equivalence pH.

Additional Resources

Understanding the pH Dynamics in the Titration of Ammonia with Nitric Acid

Titration experiments serve as fundamental tools in analytical chemistry, providing insights into the concentration of unknown solutions and the behavior of acids and bases. When dealing with weak bases such as ammonia (NH_3) and strong acids like nitric acid (HNO_3), the titration process offers a fascinating window into the interplay of chemical equilibria, pH changes, and stoichiometry. This article explores a specific scenario: a 100.0 mL sample of 0.10 M NH_3 being titrated with 0.10 M HNO_3, and aims to determine the pH of the solution at various stages of the titration process.

Introduction to Ammonia and Nitric Acid Titration

The Chemistry of Ammonia (NH_3)

Ammonia is a weak, monoprotic base characterized by its ability to accept a proton:

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

In aqueous solutions, ammonia establishes an equilibrium with water, producing ammonium ions (NH_4^+) and hydroxide ions (OH^-). Its base dissociation constant (Kb) is

approximately 1.8×10^{-5}, which indicates moderate basicity. The pH of a solution containing ammonia depends on its concentration and the extent of this equilibrium.

The Chemistry of Nitric Acid (HNO_3)

Nitric acid is a strong acid that completely dissociates in water:

$HNO_3 \rightarrow H^+ + NO_3^-$

Given its complete ionization, HNO_3's role in titration is straightforward, providing a source of protons (H^+) to neutralize the ammonia base.

The Titration Process

The titration involves gradually adding nitric acid to the ammonia solution. Initially, the pH is high due to ammonia's basic nature. As acid is added, the pH decreases, passing through various stages—buffer regions, equivalence point, and finally, the point where excess acid determines the pH.

Initial Conditions and Calculations

Initial Moles of Ammonia

Given:
- Volume of ammonia sample = 100.0 mL = 0.100 L
- Concentration of ammonia = 0.10 M

Calculating initial moles:
$$ \text{Moles of NH}_3 = 0.10\, \text{mol/L} \times 0.100\, \text{L} = 0.010\, \text{mol} $$

Initial pH of the Ammonia Solution

The initial pH can be calculated from the ammonia's equilibrium:

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

Using the Kb value:
$$ K_b = 1.8 \times 10^{-5} $$

The initial concentration of NH_3 is 0.10 M, and we assume x mol/L of NH_3 reacts to produce x mol/L of NH_4^+ and OH^-:

$$ K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} \approx \frac{x^2}{0.10 - x} $$

Since x is small relative to 0.10, approximate:

$$ K_b \approx \frac{x^2}{0.10} \Rightarrow x^2 = K_b \times 0.10 = 1.8 \times 10^{-5} \times 0.10 = 1.8 \times 10^{-6} $$

$$ x = \sqrt{1.8 \times 10^{-6}} \approx 1.34 \times 10^{-3} \, \text{M} $$

This is the concentration of OH^- ions initially present.

Calculating pOH:
$$ pOH = -\log [OH^-] = -\log (1.34 \times 10^{-3}) \approx 2.87 $$

And thus pH:
$$ pH = 14 - pOH \approx 11.13 $$

Initial pH ≈ 11.13

Progression of the Titration: Key Stages and Calculations

The titration process can be divided into distinct stages:

1. Initial Stage: Before any acid is added.
2. Buffer Region: When NH_3 and NH_4^+ coexist.
3. Equivalence Point: When all ammonia has reacted with acid.
4. Post-Equivalence (Excess Acid): When additional HNO_3 is added beyond the equivalence point.

Each stage involves specific pH calculations based on the amounts of reactants and the nature of the species present.

Stage 1: Before Titration (0 mL of HNO_3 added)

As previously calculated, the initial pH of the ammonia solution is approximately 11.13.

Stage 2: During Titration (Partial neutralization)

When some nitric acid has been added, the ammonia begins to convert into ammonium:

$$NH_3 + HNO_3 \rightarrow NH_4^+ + NO_3^-$$

Suppose we add a volume V_{acid} of 0.10 M HNO_3.

- Moles of HNO_3 added:
$$ n_{HNO_3} = 0.10\, \text{mol/L} \times V_{acid} $$

- Moles of NH_3 remaining after addition:
$$ n_{NH_3, remaining} = 0.010\, \text{mol} - n_{HNO_3} $$

- Moles of NH_4^+ produced:
$$ n_{NH_4^+} = n_{HNO_3} $$

The pH at this point depends on the buffer system formed by the weak base NH_3 and its conjugate acid NH_4^+.

Buffer Calculation:
Using the Henderson-Hasselbalch equation:
$$ pH = pK_a + \log \left(\frac{[base]}{[acid]} \right) $$

Where:
- $ pK_a $ of NH_4^+ can be derived from $ K_a $ of NH_4^+:

$$
K_a = \frac{K_w}{K_b} = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} \approx 5.56 \times 10^{-10}
$$

$$
pK_a = -\log K_a \approx 9.26
$$

- Concentrations:

$$
[base] = \frac{\text{remaining NH}_3}{\text{total volume}}, \quad [acid] = \frac{\text{NH}_4^+}{\text{total volume}}
$$

Total volume after adding acid:
$$ V_{total} = 0.100\, \text{L} + V_{acid} $$

Calculations for specific V_{acid} can be performed to find pH at different points.

Stage 3: At the Equivalence Point

The equivalence point occurs when all ammonia has been neutralized:

$$
n_{NH_3} = n_{HNO_3}
$$

\]

- Moles of acid added:
\[n_{HNO_3} = 0.010\, \text{mol} \]
- Volume of acid needed:
\[V_{eq} = \frac{0.010\, \text{mol}}{0.10\, \text{mol/L}} = 0.100\, \text{L} = 100\, \text{mL} \]

At this point, all NH_3 has converted into NH_4^+. Because NH_4^+ is a weak acid, it hydrolyzes in water:

\[
NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+
\]

The pH is determined by the hydrolysis of NH_4^+.

Calculating pH at the Equivalence Point:

- Moles of NH_4^+:
\[0.010\, \text{mol} \]
- Total volume:
\[0.100\, \text{L} + 0.100\, \text{L} = 0.200\, \text{L} \]
- Concentration of NH_4^+:
\[[NH_4^+] = \frac{0.010\, \text{mol}}{0.200\, \text{L}} = 0.05\, \text{M} \]

Hydrolysis constant:
\[
K_a^{NH_4^+} = \frac{K_w}{K_b} = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} \approx 5.56 \times 10^{-10}
\]

Set up equilibrium:
\[
NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3

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