A 118.8 Ml Sample Of 0.120 M Methylamine (ch3nh2;kb=3.7104) Is Titrated With 0.245 M Hno3. Calculate

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Introduction to Acid-Base Titration and Its Significance

Titration is a fundamental laboratory technique used extensively in analytical chemistry to determine the concentration of an unknown solution. It involves the gradual addition of a standard solution of known concentration (titrant) to an analyte until the reaction reaches its endpoint, often indicated by a color change or an electrical measurement. This precise method enables chemists to quantify concentrations with high accuracy, which is crucial for various applications in pharmaceuticals, environmental analysis, food chemistry, and industrial processes.

In the context of our problem, we are dealing with a titration involving methylamine, a weak base, and nitric acid (HNO_3), a strong acid. The goal is to determine the amount of acid required to neutralize a specified volume and concentration of methylamine. Such calculations are essential for understanding the properties of weak bases, their dissociation constants, and their behavior in aqueous solutions.

Understanding the Chemical Components

Methylamine (CH_3NH_2)

Methylamine is an organic compound classified as a primary amine. Its chemical formula is CH_3NH_2, and it exhibits basic properties due to the lone pair of electrons on the nitrogen atom. It reacts with acids to form ammonium salts. The base dissociation constant (Kb) for methylamine is provided as 3.7104, indicating its strength as a weak base.

Nitric Acid (HNO_3)

Nitric acid is a strong acid that fully dissociates in aqueous solutions. It readily reacts with bases like methylamine to produce corresponding ammonium salts and water.

Reaction Details and Stoichiometry

Balanced Chemical Equation

The titration involves a simple acid-base neutralization:

$$ \mathrm{CH_3NH_2 + HNO_3 \rightarrow CH_3NH_3^+ + NO_3^-} $$

This reaction occurs in a 1:1 molar ratio, meaning one mole of methylamine reacts with one mole of nitric acid.

Key Data Provided

- Volume of methylamine solution: 118.8 mL (or 0.1188 L)
- Concentration of methylamine: 0.120 M
- Concentration of nitric acid: 0.245 M
- Base dissociation constant (Kb): 3.7104 (for methylamine)

Step-by-Step Calculation of the Titration

The main objective is to determine either the volume of HNO_3 needed to neutralize the methylamine or to understand the moles involved during titration. The process involves several calculations:

Step 1: Calculate the moles of methylamine present

$$
\text{Moles of methylamine} = \text{Concentration} \times \text{Volume}
$$

$$
= 0.120\, \text{mol/L} \times 0.1188\, \text{L} = 0.014256\, \text{mol}
$$

Step 2: Determine the moles of HNO_3 required for complete neutralization

Since the reaction ratio is 1:1,

$$
\text{Moles of HNO}_3 = \text{Moles of methylamine} = 0.014256\, \text{mol}
$$

Step 3: Calculate the volume of HNO_3 needed

Using the molarity of HNO_3,

$$
\text{Volume of HNO}_3 = \frac{\text{Moles of HNO}_3}{\text{Concentration of HNO}_3}
$$

$$
= \frac{0.014256\, \text{mol}}{0.245\, \text{mol/L}} \approx 0.0582\, \text{L} = 58.2\, \text{mL}
$$

Therefore, approximately 58.2 mL of 0.245 M HNO_3 is required to completely titrate the methylamine sample.

Calculating the pH at Various Stages of Titration

Understanding pH changes during titration is vital for identifying the equivalence point and the buffer regions. Since methylamine is a weak base, its initial pH will be higher, and as acid is added, the pH will decrease until neutralization is complete.

Initial pH of Methylamine Solution

Given the concentration of methylamine and its Kb, we can calculate the initial pOH and pH.

Step 1: Find the concentration of OH^- ions from methylamine

$$
K_b = \frac{[\mathrm{CH_3NH_3^+}][\mathrm{OH}^-]}{[\mathrm{CH_3NH_2}]}
$$

Assuming initial dissociation:

$$
[\mathrm{OH}^-] = x
$$

$$
K_b = \frac{x^2}{[\mathrm{CH_3NH_2}] - x} \approx \frac{x^2}{0.120}
$$

Since Kb is small, $x \ll 0.120$, so:

$$
x^2 = K_b \times 0.120 = 3.7104 \times 0.120 = 0.445248
$$

$$
x = \sqrt{0.445248} \approx 0.6675\, \text{M}
$$

But this indicates a very high concentration of OH^-, which suggests the initial approximation may not be valid directly; instead, solving the quadratic equation precisely is better.

Exact calculation:

$$
K_b = \frac{x^2}{0.120 - x}
$$

$$
x^2 + K_b \times x - K_b \times 0.120 = 0
$$

Plugging in the numbers:

$$
x^2 + 3.7104 x - (3.7104 \times 0.120) = 0
$$

$$
x^2 + 3.7104 x - 0.445248 = 0
$$

Solve using quadratic formula:

$$
x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}
$$

where $a=1$, $b=3.7104$, $c=-0.445248$:

$$
x = \frac{-3.7104 \pm \sqrt{(3.7104)^2 - 4 \times 1 \times (-0.445248)}}{2}
$$

$$
x = \frac{-3.7104 \pm \sqrt{13.779 - (-1.780992)}}{2}
$$

$$
x = \frac{-3.7104 \pm \sqrt{13.779 + 1.780992}}{2} = \frac{-3.7104 \pm \sqrt{15.56}}{2}
$$

$$
x = \frac{-3.7104 \pm 3.944}{2}
$$

Positive root:

$$
x = \frac{-3.7104 + 3.944}{2} = \frac{0.2336}{2} \approx 0.1168\, \text{M}
$$

Thus, initial $[\mathrm{OH}^-] \approx 0.1168\, \text{M}$.

Step 2: Calculate initial pOH and pH

$$
\text{pOH} = -\log(0.1168) \approx 0.932
$$

\[

$$
\text{pH} = 14 - 0.932 \approx 13.07
$$

This high pH confirms methylamine's basic nature.

pH at the Equivalence Point

At the equivalence point, all methylamine has been converted into its ammonium salt, which is a weak acid. Therefore, the pH will be less than 7, determined by the hydrolysis of $\mathrm{CH_3NH_3^+}$.

Understanding the Role of Kb and Calculating pKa

Since the pH at various points depends on the equilibria involving methylamine and its conjugate acid, understanding the relationship between Kb and Ka is important.

$$
K_a \times K_b = K_w
$$

where $K_w = 1.0 \times 10^{-14}$ at 25°C.

Calculating K_a:

$$
K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{3.7104} \approx 2.695 \times 10^{-15}
$$

And the pKa:

$$
\text{pKa} = -\log K_a \approx -\log(2.695 \times 10^{-15}) \approx 14 - \log(2.695) \approx 14 - 0.431 = 13.569
$$

The conjugate acid is very weak, indicating the

Frequently Asked Questions

What is the molarity of methylamine in the sample?

The molarity of methylamine (CH_3NH_2) in the sample is 0.120 M.

How do you determine the moles of methylamine in the sample?

Multiply the volume in liters by the molarity: 0.1188 L × 0.120 mol/L = 0.014256 mol of methylamine.

What is the balanced chemical equation for the titration of methylamine with HNO3?

$CH_3NH_2 + HNO_3 \rightarrow CH_3NH_3^+ + NO_3^-$

How many moles of HNO3 are needed to completely neutralize the methylamine sample?

Since the molar ratio is 1:1, moles of HNO_3 needed = 0.014256 mol.

What volume of 0.245 M HNO3 is required to titrate the methylamine sample?

Volume = moles / molarity = 0.014256 mol / 0.245 mol/L ≈ 0.0582 L or 58.2 mL.

How can the pKa of methylamine be estimated using the given Kb value?

pKa can be calculated using pKa + pKb = 14; first find pKb = $-\log(3.7104 \times 10^{-4}) \approx 3.43$, then pKa = 14 - 3.43 ≈ 10.57.

What is the significance of the Kb value for methylamine in this titration?

Kb indicates the base strength of methylamine, affecting the pH during titration and the calculation of equivalence point.

How does the titration process help determine the pH of the solution at

the equivalence point?

At equivalence, the solution contains methylammonium ions, and the pH is determined by the hydrolysis of these ions, which can be calculated using Ka and Kb values.

What is the final pH of the solution at the equivalence point of this titration?

The pH at equivalence can be calculated from the hydrolysis of methylammonium ions, typically resulting in a slightly acidic pH, around 5.5 to 6.0, depending on concentrations.

Additional Resources

A 118.8 mL sample of 0.120 M methylamine (CH_3NH_2; Kb = 3.71×10^{-4}) titrated with 0.245 M HNO_3 is an intriguing chemical problem that combines principles of acid-base titration, stoichiometry, and equilibrium calculations. This scenario offers a comprehensive opportunity to analyze how weak bases interact with strong acids, interpret titration data, and perform precise calculations to determine unknown quantities, such as the concentration of residual species or the volume of titrant used at various points in the titration. Such problems are fundamental in analytical chemistry, providing insight into the behavior of chemical species in solution and the quantitative methods used to analyze them.

Understanding the Components of the Titration

1. Methylamine (CH_3NH_2)

Methylamine is a weak base characterized by its amino group, which can accept protons from acids. Its base dissociation constant (Kb) is given as 3.71×10^{-4}. This value indicates its relative strength compared to other weak bases—smaller Kb values correspond to weaker bases. The initial concentration (0.120 M) and volume (118.8 mL) define the initial amount of methylamine present, which is critical for calculations involving moles and titration progress.

Features of methylamine:
- It is a primary amine, with the formula CH_3NH_2.
- It acts as a weak base, reacting with H^+ ions to form $CH_3NH_3^+$.
- It has a Kb value of 3.71×10^{-4}, indicating moderate basicity.

Pros:

- Acts as a good buffer in certain pH ranges.
- Its behavior is well-understood, making it ideal for titration experiments.

Cons:
- Being a weak base, it doesn't completely react with strong acids, requiring careful calculation.
- It can be hazardous if not handled properly.

2. Nitric Acid (HNO_3)

Nitric acid is a strong acid, fully dissociating in aqueous solution. Its concentration (0.245 M) determines the amount of H^+ ions available to react with methylamine during titration.

Features of HNO_3:
- Fully dissociates into H^+ and NO_3^- ions.
- Provides a straightforward titration endpoint because of its complete dissociation.
- Used widely in analytical chemistry due to its strong acidity.

Pros:
- Reliable and easy to work with in titrations.
- Provides clear endpoint indications with suitable indicators.

Cons:
- Corrosive and hazardous; requires proper handling.
- Excess nitric acid can cause complications if not titrated properly.

Calculations and Titration Mechanics

1. Determining Initial Moles of Methylamine

First, calculate the initial moles of methylamine present before titration:

$$
\text{Moles of methylamine} = M \times V = 0.120\, \text{mol/L} \times 0.1188\, \text{L} = 0.014256\, \text{mol}
$$

This initial amount sets the baseline for subsequent titration calculations.

2. Moles of HNO_3 Needed for Complete Neutralization

Since nitric acid is a strong acid and methylamine is a weak base, their reaction is:

$$
\text{CH}_3\text{NH}_2 + \text{HNO}_3 \rightarrow \text{CH}_3\text{NH}_3^+ + \text{NO}_3^-
$$

The mole ratio is 1:1. Therefore, to fully neutralize methylamine:

$$
\text{Moles of HNO}_3 = \text{Moles of methylamine} = 0.014256\, \text{mol}
$$

The volume of HNO_3 required for complete neutralization:

$$
V_{HNO_3} = \frac{\text{moles}}{\text{concentration}} = \frac{0.014256\, \text{mol}}{0.245\, \text{mol/L}} \approx 0.0582\, \text{L} = 58.2\, \text{mL}
$$

Thus, approximately 58.2 mL of 0.245 M HNO_3 is needed for complete neutralization.

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pH Calculations at Different Titration Stages

1. Initial pH of the Methylamine Solution

The initial pH of the methylamine solution can be calculated using the Kb expression:

$$
K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]}
$$

Assuming x is the concentration of OH^- produced at equilibrium:

$$
K_b = \frac{x^2}{0.120 - x} \approx \frac{x^2}{0.120}
$$

Since K_b is small, x is small relative to initial concentration:

$$
x^2 = K_b \times 0.120 = 3.71 \times 10^{-4} \times 0.120 = 4.452 \times 10^{-5}
$$
$$
x = \sqrt{4.452 \times 10^{-5}} \approx 6.67 \times 10^{-3}\, \text{M}
$$

pOH:

$$
pOH = -\log(6.67 \times 10^{-3}) \approx 2.18
$$

pH:

$$
pH = 14 - pOH = 14 - 2.18 = 11.82
$$

The initial solution is basic with a pH around 11.82.

2. Equivalence Point pH

At the equivalence point, all methylamine has reacted to form methylammonium ions (CH_3NH_3^+). The solution now contains only the conjugate acid of methylamine, which is a weak acid.

The moles of methylammonium formed:

$$
\text{Moles} = 0.014256\, \text{mol}
$$

Total volume at equivalence:

$$
V_{total} = 118.8\, \text{mL} + 58.2\, \text{mL} = 177\, \text{mL} = 0.177\, \text{L}
$$

Concentration of methylammonium:

$$
[\text{CH}_3\text{NH}_3^+] = \frac{0.014256\, \text{mol}}{0.177\, \text{L}} \approx 0.0805\, \text{M}
$$

To find the pH, consider the hydrolysis of methylammonium:

$$
\text{CH}_3\text{NH}_3^+ + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_2 + \text{H}_3\text{O}^+
$$

The acid dissociation constant K_a relates to K_b:

$$
K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{3.71 \times 10^{-4}} \approx 2.69 \times 10^{-11}
$$

Using the expression:

$$
K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{NH}_2]}{[\text{CH}_3\text{NH}_3^+]} \approx \frac{x^2}{0.0805}
$$

$$
x^2 = K_a \times 0.0805 \approx 2.69 \times 10^{-11} \times 0.0805 \approx 2.17 \times 10^{-12}
$$

$$
x = \sqrt{2.17 \times 10^{-12}} \approx 1.47 \times 10^{-6}\, \text{M}
$$

pH:

$$
pH = -\log(1.47 \times 10^{-6}) \approx 5.83
$$

The solution at the equivalence point is slightly acidic (pH ~ 5.83) due to hydrolysis of methylammonium ions.

Additional Calculations and Considerations

1. Determining the pH at Various Volumes

Between the initial solution and the equivalence point, the pH can be calculated at any given volume of titrant added by accounting for the remaining methylamine and the formed methylammonium ions, using appropriate equilibrium expressions.

2. Beyond the Equivalence Point

If titration continues past the equivalence point, excess HNO_3 determines the pH:

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