A 25. 0 Ml Sample Of 0. 150 M Hydrofluoric Acid Is Titrated With A 0. 150 M Naoh Solution. What Is The

Introduction

A 25.0 mL sample of 0.150 M hydrofluoric acid is titrated with a 0.150 M NaOH solution. What is the resulting pH at various stages of the titration process, and how can we determine the equivalence point and the concentration of unreacted species? Understanding this titration involves exploring acid-base chemistry principles, calculating moles, and analyzing pH changes during the titration process. Hydrofluoric acid (HF) is a weak acid, while sodium hydroxide (NaOH) is a strong base, making this a classic weak acid-strong base titration scenario. This article will delve into the detailed steps to analyze this titration, including calculations of initial concentrations, volume of titrant needed to reach equivalence, pH at various points, and the interpretation of titration curves.

Fundamentals of Hydrofluoric Acid and Sodium Hydroxide

Properties of Hydrofluoric Acid (HF)

- HF is a weak acid with a dissociation constant (Ka) of approximately 6.6×10^{-4} at 25°C.
- It partially dissociates in aqueous solution:

$HF \rightleftharpoons H^+ + F^-$

- Despite being weak, HF is highly dangerous due to its ability to penetrate tissues and cause severe burns.

Properties of Sodium Hydroxide (NaOH)

- NaOH is a strong base that dissociates completely in water:

$NaOH \rightarrow Na^+ + OH^-$

- The complete dissociation simplifies calculations of titration.

Initial Conditions and Calculations

Moles of Hydrofluoric Acid in the Sample

Given data:

- Volume of HF sample: 25.0 mL = 0.0250 L
- Concentration of HF: 0.150 M

Calculating moles of HF:

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Moles of HF = concentration × volume
= 0.150 mol/L × 0.0250 L
= 0.00375 mol
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Moles of NaOH Needed to Reach Equivalence Point

- Since the titration involves equal molar amounts of HF and NaOH at equivalence, the moles of NaOH required:

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Moles of NaOH = moles of HF = 0.00375 mol
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- Volume of NaOH solution needed:

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Volume of NaOH = moles / concentration
= 0.00375 mol / 0.150 mol/L
= 0.0250 L = 25.0 mL
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Thus, 25.0 mL of 0.150 M NaOH will exactly neutralize the HF sample.

Understanding the Titration Curve

Ph at the Start of Titration

- The initial pH is determined by the weak acid's dissociation.

Calculation of initial pH:

- Use the expression for weak acid dissociation:

$HF \rightleftharpoons H^+ + F^-$

- Set up the equilibrium:

Let $x = [H^+]$ at equilibrium,

$Ka = [H^+][F^-] / [HF]$

- Initial concentration of HF: 0.150 M

- At the start, assuming x is small:

$Ka = x^2 / (0.150 - x) \approx x^2 / 0.150$

$x^2 = Ka \times 0.150$

$x^2 = 6.6 \times 10^{-4} \times 0.150 \approx 9.9 \times 10^{-5}$

$x \approx \sqrt{(9.9 \times 10^{-5})} \approx 0.00995$ M

- $pH = -\log[H^+] \approx -\log(0.00995) \approx 2.00$

Initial pH ≈ 2.00

pH at the Equivalence Point

- At equivalence, all HF has been neutralized to F^- ions.

- The solution contains only the conjugate base, fluoride ions (F^-).

- Since fluoride is a weak base, the solution's pH is determined by hydrolysis:

$F^- + H_2O \rightleftharpoons HF + OH^-$

- To find the pOH and then pH:

- Calculate the concentration of F^-:

Moles of F^- = moles of HF initially = 0.00375 mol

Total volume after titration:

25.0 mL HF + 25.0 mL NaOH = 50.0 mL = 0.0500 L

$[F^-]$ = 0.00375 mol / 0.0500 L = 0.075 M

- Equilibrium expression:

$K_b = K_w / K_a = 1.0 \times 10^{-14} / 6.6 \times 10^{-4} \approx 1.52 \times 10^{-11}$

For F^- hydrolysis:

$F^- + H_2O \rightleftharpoons HF + OH^-$

Let $y = [OH^-]$:

$K_b = y^2 / (\text{initial } F^- \text{ concentration} - y) \approx y^2 / 0.075$

$y^2 = K_b \times 0.075 \approx 1.52 \times 10^{-11} \times 0.075 \approx 1.14 \times 10^{-12}$

$y \approx \sqrt{(1.14 \times 10^{-12})} \approx 1.07 \times 10^{-6}$ M

- $pOH = -\log(1.07 \times 10^{-6}) \approx 5.97$

- $pH = 14 - pOH \approx 8.03$

Therefore, at the equivalence point, $pH \approx 8.03$

pH After Partial Titration (Before Equivalence Point)

- When titrating with NaOH, the pH increases gradually.

- For small additions of NaOH (<25 mL), the solution contains mostly unreacted HF and some formed HF molecules.

- To calculate pH at any point before equivalence, use the remaining HF and the amount of F^- formed.

Example: pH after adding 10 mL of NaOH

- Volume of NaOH added: 10 mL = 0.010 L

- Moles of NaOH added:

0.150 mol/L × 0.010 L = 0.0015 mol

- Moles of HF remaining:

Moles initially: 0.00375 mol

Moles reacted: 0.0015 mol

Remaining HF: 0.00375 - 0.0015 = 0.00225 mol

- Moles of F^- formed:

= moles of NaOH added = 0.0015 mol

- Total volume:

25 mL HF + 10 mL NaOH = 35 mL = 0.035 L

- Concentrations:

[HF] = 0.00225 mol / 0.035 L ≈ 0.0643 M

$[F^-]$ = 0.0015 mol / 0.035 L ≈ 0.0429 M

- The solution is a mixture of weak acid and its conjugate base, requiring a buffer pH calculation using the Henderson-Hasselbalch equation:

$pH = pKa + \log([F^-]/[HF])$

- pKa = -log(Ka) ≈ 3.18

- pH calculation:

pH ≈ 3.18 + log(0.0429 / 0.0643) ≈ 3.18 + log(0.666) ≈ 3.18 - 0.177 ≈ 3.00

Thus, after adding 10 mL of NaOH, pH ≈ 3.00

Graphing and Interpreting the Titration Curve

Key Features of the Titration Curve

- Initial pH: approximately 2.00
- Buffer region: where pH gradually increases, near the half-equivalence point
- Equivalence point: pH ≈ 8.03
- Post-equivalence: pH rises sharply, characteristic of a strong base in solution

Constructing the Titration Curve

- Plot volume of NaOH added on the x-axis
- Plot pH on the y-axis
- Identify the steep rise near the equivalence point
- Mark initial pH and pH at various points for analysis

Significance and Applications

Industrial and Laboratory Relevance

- Precise titration techniques are essential for quality control
- Hydrofluoric acid's use in etching and cleaning makes understanding its titration critical
- Accurate pH measurements inform safety precautions and process controls

Educational Importance

- Demonstrates principles of weak acid-

Frequently Asked Questions

What is the balanced chemical equation for the titration of hydrofluoric acid with sodium hydroxide?

$HF + NaOH \rightarrow NaF + H_2O$

How do you calculate the moles of hydrofluoric acid in the 25.0 mL sample?

Moles of HF = concentration × volume = 0.150 M × 0.025 L = 3.75×10^{-3} mol

What is the volume of NaOH solution required to completely neutralize the hydrofluoric acid?

Since the molar ratio is 1:1, moles of NaOH needed = 3.75×10^{-3} mol. Volume = moles / concentration = 3.75×10^{-3} mol / 0.150 M = 0.025 L or 25.0 mL

How do we determine the endpoint in this titration?

The endpoint is typically detected using a suitable indicator that changes color at the equivalence point, such as phenolphthalein, which turns pink in basic solution.

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

Since hydrofluoric acid is a weak acid and NaOH is a strong base, the pH at the equivalence point will be greater than 7, approximately around 8-9 due to hydrofluoride ion hydrolysis.

Why is hydrofluoric acid considered hazardous despite being a weak acid?

Hydrofluoric acid is highly toxic and can penetrate skin, cause severe burns, and interfere with calcium metabolism, making it extremely dangerous even at low concentrations.

What is the significance of molarity in this titration problem?

Molarity allows us to relate the volume of solutions to the number of moles of solute, enabling calculations of the titrant volume needed for neutralization.

How can this titration be used to determine the concentration of an unknown hydrofluoric acid sample?

By titrating the unknown sample with a known concentration of NaOH and measuring the volume used to reach the endpoint, you can calculate the concentration of HF in the sample using stoichiometry.

Additional Resources

A 25.0 mL sample of 0.150 M hydrofluoric acid is titrated with a 0.150 M NaOH solution. What is the significance of this titration, and how do we determine the resulting pH at various stages? This scenario exemplifies fundamental concepts in acid-base chemistry, emphasizing the importance of understanding titration principles, acid strengths, and calculating pH changes during titration curves. In this article, we explore the detailed analytical process behind this titration, including the chemical reactions involved, the calculations necessary to interpret the titration data, and the broader implications for laboratory practice and chemical understanding.

Introduction to Acid-Base Titrations

Titration is a classical laboratory technique used to determine the concentration of an unknown solution by reacting it with a solution of known concentration. It is a cornerstone in analytical chemistry, providing insights into chemical composition, purity, and reactivity. The process involves gradually adding a titrant (here, sodium hydroxide, NaOH) to a analyte (hydrofluoric acid, HF) until the reaction reaches its equivalence point, where the amount of titrant added is stoichiometrically equivalent to the analyte present.

In this particular scenario, a 25.0 mL sample of hydrofluoric acid, with a known molarity of 0.150 M, is titrated with a 0.150 M NaOH solution. Both solutions have identical molarity, which simplifies some calculations but also raises interesting questions about the titration process, especially considering the weak acid nature of HF and its partial dissociation in aqueous solutions.

Understanding the Chemistry of Hydrofluoric Acid and Sodium Hydroxide

Properties of Hydrofluoric Acid (HF)

Hydrofluoric acid is a weak, monoprotic acid characterized by its ability to partially dissociate in water:

$$\text{HF} \leftrightarrow \text{H}^+ + \text{F}^-$$

Unlike strong acids such as HCl or H_2SO_4, HF does not fully dissociate in solution. Its acid dissociation constant (Ka) is approximately 6.6×10^{-4}, indicating a significant degree of dissociation but still incomplete. This partial dissociation influences how the pH changes during titration and necessitates specific calculations for accurate pH determination.

Sodium Hydroxide (NaOH) as a Strong Base

NaOH is a strong base that dissociates completely in aqueous solution:

$$\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-$$

The complete dissociation simplifies calculations related to the titrant, as the molarity of hydroxide ions directly equals the molarity of NaOH.

Setting Up the Titration: Initial Conditions and Calculations

Initial Moles of Hydrofluoric Acid

Given:
- Volume of HF solution: 25.0 mL = 0.0250 L

- Concentration of HF: 0.150 M

Calculating moles of HF:

$$ \text{Moles of HF} = \text{Concentration} \times \text{Volume} $$
$$ = 0.150\, \text{mol/L} \times 0.0250\, \text{L} = 3.75 \times 10^{-3}\, \text{mol} $$

This is the initial amount of hydrofluoric acid in the sample prior to titration.

Initial pH of the HF Solution

Since HF is a weak acid, its initial pH can be estimated using the acid dissociation constant:

$$ \text{HF} \leftrightarrow \text{H}^+ + \text{F}^- $$

The equilibrium expression:

$$ K_a = \frac{[\text{H}^+][\text{F}^-]}{[\text{HF}]} $$

Assuming x is the concentration of H^+ at equilibrium:

$$ K_a = \frac{x^2}{[\text{HF}] - x} $$

Given:

$$ [\text{HF}] \approx 0.150\, \text{M} $$

and

$$ K_a = 6.6 \times 10^{-4} $$

we approximate:

$$ x^2 \approx K_a \times [\text{HF}] = 6.6 \times 10^{-4} \times 0.150 \approx 9.9 \times 10^{-5} $$

$$ x \approx \sqrt{9.9 \times 10^{-5}} \approx 0.00995\, \text{M} $$

Thus, initial pH:

$$ pH = -\log [\text{H}^+] = -\log 0.00995 \approx 2.00 $$

This initial pH indicates a moderately acidic solution, consistent with weak acid behavior.

The Titration Process: Step-by-Step Analysis

Adding NaOH: The Neutralization Reaction

The core chemical reaction during titration is:

$$ \text{HF} + \text{NaOH} \rightarrow \text{NaF} + \text{H}_2\text{O} $$

Because both solutions are 0.150 M, the stoichiometry is 1:1. When equal molar amounts of HF and NaOH are combined, the acid is neutralized to form fluoride ions and water.

Calculating the Volume of NaOH at the Equivalence Point

To find the equivalence point volume:

$$ \text{Moles of HF} = 3.75 \times 10^{-3}\, \text{mol} $$

Since NaOH molarity is the same:

$$ \text{Volume of NaOH} = \frac{\text{Moles of NaOH}}{\text{Concentration}} = \frac{3.75 \times 10^{-3}\, \text{mol}}{0.150\, \text{mol/L}} = 0.025\, \text{L} = 25.0\, \text{mL} $$

Therefore, adding 25.0 mL of NaOH will bring the titration to the equivalence point, where the amount of base exactly neutralizes the initial acid.

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pH at the Equivalence Point: Analyzing the Results

Since HF is a weak acid and its conjugate base, fluoride (F^-), is relatively weak, the solution at the equivalence point will not be neutral (pH 7). Instead, it will be slightly basic due to hydrolysis of fluoride ions:

$$ \text{F}^- + \text{H}_2\text{O} \leftrightarrow \text{HF} + \text{OH}^- $$

This hydrolysis will generate hydroxide ions, increasing the pH beyond 7.

Calculating the pH at the Equivalence Point

First, determine the concentration of F^- ions:

- Moles of F^- formed:

$$ 3.75 \times 10^{-3}\, \text{mol} $$

- Total volume after titration:

$$ 25.0\, \text{mL} + 25.0\, \text{mL} = 50.0\, \text{mL} = 0.050\, \text{L} $$

- Concentration of F^-:

$$ \frac{3.75 \times 10^{-3}\, \text{mol}}{0.050\, \text{L}} = 0.075\, \text{M} $$

Next, examine the hydrolysis reaction:

$$ \text{F}^- + \text{H}_2\text{O} \leftrightarrow \text{HF} + \text{OH}^- $$

The equilibrium expression:

$$ K_b = \frac{K_w}{K_a} $$

where:

- $K_w = 1.0 \times 10^{-14}$
- K_a for HF is 6.6×10^{-4}

Calculating K_b:

$$ K_b = \frac{1.0 \times 10^{-14}}{6.6 \times 10^{-4}} \approx 1.52 \times 10^{-11} $$

Applying the hydrolysis to F^-:

$$ \text{F}^- + \text{H}_2\text{O} \leftrightarrow \text{HF} + \text{OH}^- $$

Let x be the concentration of OH^- produced:

$$ K_b = \frac{[\text{OH}^-]^2}{[\text{F}^-]} $$

$$ 1.52 \times 10^{-11} = \frac{x^2}{0.075} $$

$$ x^2 = 1.52 \times 10^{-11} \times 0.075 \approx 1.14 \times 10^{-12} $$

$$ x = \sqrt{1.14 \times 10^{-12}} \approx 1.07 \times 10^{-6}\, \text{M} $$

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