A 300.0 ML Sample Of 0.20 M HF Is Titrated With 0.10 M NaOH. Determine The PH Of The Solution After The

A 300.0 ML Sample Of 0.20 M HF Is Titrated With 0.10 M NaOH. Determine The PH Of The Solution After The titration process begins, it's essential to understand the fundamentals of acid-base titrations, especially when dealing with a weak acid like hydrofluoric acid (HF) and a strong base like sodium hydroxide (NaOH). This type of titration provides valuable insights into the concentration of unknown solutions and the pH changes that occur during the process. In this article, we will explore how to determine the pH after a specific volume of NaOH has been added to the HF solution, covering key concepts such as molarity, titration calculations, and pH determination at various stages of the titration.

Understanding the Titration Components and Initial Conditions

Initial Concentrations and Volumes

- Hydrofluoric Acid (HF):
- Volume: 300.0 mL (or 0.300 L)
- Molarity: 0.20 M
- Sodium Hydroxide (NaOH):
- Molarity: 0.10 M
- Volume added varies during titration, but for the calculation, we consider the volume added at specific points.

Key Concepts in Acid-Base Titration

- Strong Base and Weak Acid:
NaOH is a strong base that fully dissociates in solution, while HF is a weak acid that partially dissociates.
- Neutralization Reaction:
HF reacts with NaOH according to the equation:
$HF + NaOH \rightarrow NaF + H_2O$
- Equivalence Point:
The point where moles of NaOH added equal moles of HF initially present, leading to a neutralization and a significant change in pH.

Calculating the Moles of HF and NaOH

Initial Moles of HF

$$
\text{Moles of HF} = \text{Molarity} \times \text{Volume} = 0.20\, \text{mol/L} \times 0.300\, \text{L} = 0.060\, \text{mol}
$$

Determining the Moles of NaOH Added

Suppose we add a certain volume of NaOH, say V mL. The moles of NaOH added are:

$$
\text{Moles of NaOH} = 0.10\, \text{mol/L} \times \frac{V\, \text{mL}}{1000} = 0.0001\, \text{mol/mL} \times V\, \text{mL}
$$

or

$$
\text{Moles of NaOH} = 0.10 \times \frac{V}{1000}
$$

Example: NaOH volume added is 10 mL

$$
\text{Moles of NaOH} = 0.10 \times \frac{10}{1000} = 0.001\, \text{mol}
$$

This information allows us to determine the extent of neutralization and the resulting pH.

Stages of Titration and pH Calculations

The pH during titration depends on the amount of NaOH added relative to the initial moles of HF. The titration process can be divided into three main stages:

1. Before the equivalence point (more HF than NaOH added)
2. At the equivalence point (moles of NaOH added equal initial moles of HF)
3. After the equivalence point (more NaOH than HF)

Each stage requires different calculations and considerations.

Calculating pH Before the Equivalence Point

When less NaOH has been added than needed to neutralize all HF, the solution contains unreacted HF and some partially neutralized HF (forming F^- ions). Since HF is a weak acid, its dissociation equilibrium influences the pH.

Determining Remaining HF

- Moles of HF initially: 0.060 mol
- Moles of NaOH added: V mL × 0.0001 mol/mL

Remaining HF after addition:

$$
\text{Remaining HF} = 0.060\, \text{mol} - \text{moles of NaOH added}
$$

For example, if 10 mL of NaOH is added:

$$
\text{Remaining HF} = 0.060 - 0.001 = 0.059\, \text{mol}
$$

Remaining volume of the solution:

$$
V_{\text{total}} = 300\, \text{mL} + V\, \text{mL} = (300 + V)\, \text{mL}
$$

Concentration of remaining HF:

$$
[\text{HF}] = \frac{0.059\, \text{mol}}{(0.300 + V)/1000\, \text{L}}
$$

Calculating pH Using Acid Dissociation Equilibrium

- HF dissociation:

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

- Acid dissociation constant (K_a):

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

- Set up an ICE table (Initial, Change, Equilibrium) to find $[\text{H}^+]$:

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

- Approximate $[\text{H}^+]$ using the quadratic formula or simplified assumptions when concentrations are known.

The pH is then:

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

At the Equivalence Point

The equivalence point occurs when the moles of NaOH added equal the initial moles of HF:

$$
0.060\, \text{mol} = 0.10\, \text{mol/L} \times V_{\text{NaOH}}
$$

$$
V_{\text{NaOH}} = \frac{0.060}{0.10} = 0.60\, \text{L} = 60\, \text{mL}
$$

In this case, after adding 60 mL of NaOH:
- All HF has been neutralized to form F^- ions.
- The solution contains the conjugate base (F^-), which can hydrolyze to produce OH^- ions, creating a basic solution.

Calculating pH at the Equivalence Point

- Moles of F^-:

$$
0.060\, \text{mol}
$$

- Volume of solution:

$$
300\, \text{mL} + 60\, \text{mL} = 360\, \text{mL} = 0.36\, \text{L}
$$

- Concentration of F^-:

$$
[\text{F}^-] = \frac{0.060\, \text{mol}}{0.36\, \text{L}} \approx 0.167\, \text{M}
$$

- Hydrolysis of F^-:

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

- Equilibrium constant for hydrolysis (K_b):

$$
K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{6.6 \times 10^{-4}} \approx 1.5 \times 10^{-11}
$$

- Using K_b, set up the hydrolysis equilibrium to find $[\text{OH}^-]$:

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

$$
[\text{OH}^-] = \sqrt{K_b \times [\text{F}^-]} \approx \sqrt{1.5 \times 10^{-11} \times 0.167} \approx 1.58 \times 10^{-6}\, \text{M}
$$

- Calculate pOH:

$$
\text{pOH} = -\log [\text{OH}^-] \approx 5.80
$$

- Therefore, pH:

$$
\text{pH} = 14 - \text{pOH} \approx 8.20
$$

This indicates the solution is slightly basic at the equivalence point.

After the Equivalence Point: Excess NaOH

Once more NaOH has been added than the amount needed to neutralize HF, the solution contains excess OH^- ions, and the pH increases significantly.

Calculating pH with Excess NaOH

- For example, if an additional 10 mL of

Frequently Asked Questions

What is the initial pH of the 0.20 M HF solution before titration?

HF is a weak acid; using its concentration and Ka value (approx. 6.6×10^{-4}), the initial pH can be calculated as approximately 2.87.

How do you determine the pH after adding some NaOH during the titration of HF?

Calculate the moles of HF and NaOH at that point; use the Henderson-Hasselbalch equation or account for the partial neutralization to find the pH.

What is the pH at the equivalence point in the titration of HF with NaOH?

Since HF is a weak acid and NaOH is a strong base, the equivalence point will be basic, with a pH greater than 7, approximately around 8.2.

How do you calculate the pH after adding a specific volume of

NaOH before reaching the equivalence point?

Determine moles of HF and NaOH present, find the remaining unneutralized HF, and then calculate pH using the acid dissociation expression or Henderson-Hasselbalch equation.

What is the significance of the titration curve in understanding the pH changes during titration?

The titration curve illustrates how pH varies with added titrant, showing the buffer region, the equivalence point, and the steep change in pH, essential for understanding acid-base neutralization.

Additional Resources

A 300.0 mL sample of 0.20 M HF is titrated with 0.10 M NaOH. Determine the pH of the solution after the addition of a specific volume of NaOH. This problem is a classic example of acid-base titration involving a weak acid (HF) and a strong base (NaOH). Understanding the nuances of this titration requires a grasp of concepts related to molarity, titration calculations, acid dissociation, and buffer solutions. In this comprehensive guide, we will walk through each step to determine the pH after different volumes of NaOH are added, providing insights into the chemistry behind the process and detailed calculations to elucidate the concepts involved.

Introduction to Acid-Base Titration and HF

Titration is a laboratory procedure used to determine the concentration of an unknown solution by reacting it with a solution of known concentration. When titrating a weak acid like hydrofluoric acid (HF) with a strong base such as sodium hydroxide (NaOH), the pH changes systematically as the titration progresses.

Why Focus on HF and NaOH?

- Hydrofluoric acid (HF) is a weak acid with a dissociation constant (Ka) of approximately 6.6 × 10^(-4). It partially dissociates in water, producing H+ and F- ions.
- Sodium hydroxide (NaOH) is a strong base that dissociates completely in water, providing OH- ions.
- The titration involves neutralizing HF's acidity with NaOH, leading to different pH regions depending on the amount of base added.

Fundamental Concepts Needed

Before diving into calculations, let's review some key concepts:

Molarity (M)

- Molarity measures concentration as moles of solute per liter of solution.
- Given:
- HF: 0.20 M in 300.0 mL

- NaOH: 0.10 M, variable volume

Moles of Substance

- Moles = Molarity × Volume (in liters)

pH and pOH

- pH = -log[H+]
- pOH = -log[OH-]
- pH + pOH = 14 (at 25°C)

Acid-Base Neutralization

- Strong acid + strong base: pH at equivalence is 7.
- Weak acid + strong base: pH at equivalence > 7 due to the conjugate base's hydrolysis.

Step 1: Calculating Initial Moles of HF

Let's start with the initial amount of HF in moles:

- Volume of HF solution = 300.0 mL = 0.300 L
- Molarity of HF = 0.20 M

Moles of HF initially:

$$ \text{moles HF} = 0.20\, \text{mol/L} \times 0.300\, \text{L} = 0.060\, \text{mol} $$

This is the total amount of HF available to react during titration.

Step 2: Calculating Moles of NaOH Added

Suppose we add a volume V mL of NaOH solution. The amount of NaOH in moles is:

- Convert volume to liters: $V\, \text{mL} = V/1000\, \text{L}$
- Molarity of NaOH = 0.10 M

Moles of NaOH added:

$$ \text{moles NaOH} = 0.10\, \text{mol/L} \times \frac{V}{1000}\, \text{L} = 0.0001\, V\, \text{mol} $$

Step 3: Determining the Reaction Progress

The reaction between HF and NaOH is:

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

- Complete reaction occurs until either:
- All HF is neutralized (at the equivalence point).
- NaOH is exhausted (before equivalence point).
- HF remains (after equivalence point).

Cases to consider:

1. Before the equivalence point: moles of NaOH < moles of HF
2. At the equivalence point: moles of NaOH = moles of HF
3. After the equivalence point: moles of NaOH > moles of HF

Step 4: Calculating pH at Different Stages

Let's analyze three typical stages: before equivalence, at equivalence, and after equivalence.

Stage 1: Before the Equivalence Point

Example: Suppose 10 mL of NaOH has been added.

- Moles NaOH added:

$$ 0.10\, \text{mol/L} \times 0.010\, \text{L} = 0.001\, \text{mol} $$

- Moles of HF remaining:

$$ 0.060\, \text{mol} - 0.001\, \text{mol} = 0.059\, \text{mol} $$

- Moles of F-

$$ \text{formed} = \text{moles NaOH added} = 0.001\, \text{mol} $$

- Total volume of solution:

$$ 300\, \text{mL} + 10\, \text{mL} = 310\, \text{mL} = 0.310\, \text{L} $$

- Concentrations:

- HF: $\frac{0.059\, \text{mol}}{0.310\, \text{L}} \approx 0.190\, \text{M}$

- F-: $\frac{0.001\, \text{mol}}{0.310\, \text{L}} \approx 0.00323\, \text{M}$

pH Calculation:

Since HF is a weak acid, and F- is its conjugate base, the solution now is a buffer. To find the pH, use the Henderson-Hasselbalch equation:

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

- pK_a of HF:

$$ pK_a = -\log(6.6 \times 10^{-4}) \approx 3.18 $$

- Plugging in the concentrations:

$$ \text{pH} = 3.18 + \log \left(\frac{0.00323}{0.190} \right) $$

$$ \text{pH} = 3.18 + \log(0.017) $$

$$ \text{pH} = 3.18 + (-1.77) = 1.41 $$

Result: The pH after adding 10 mL of NaOH is approximately 1.41.

Stage 2: At the Equivalence Point

Determine the volume of NaOH needed to reach equivalence:

$$ \text{Moles of HF} = 0.060\, \text{mol} $$

$$ \text{Volume of NaOH} = \frac{0.060\, \text{mol}}{0.10\, \text{mol/L}} = 0.600\, \text{L} = 600\, \text{mL} $$

At this point:

- All HF has been neutralized:

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

- Moles of F- formed = initial moles of HF = 0.060 mol

- Total volume:

$$ 300\, \text{mL} + 600\, \text{mL} = 900\, \text{mL} = 0.900\, \text{L} $$

- Concentration of F-:

$$ \frac{0.060\, \text{mol}}{0.900\, \text{L}} \approx 0.067\, \text{M} $$

pH at equivalence:

Since F- is the conjugate base of HF, its hydrolysis determines the pH:

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

- The base hydrolysis constant K_b:

\[K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{6.6 \times 10^{-4}} \approx 1.52 \times 10^{-11} \]

- Using the concentration of F- and \(K_b\) to find \([\text{OH}^-]\):

\[\text{F}^- + \text{H}_2\text{O} \leftrightarrows \text{HF} + \text{OH}^- \]

\[\text{Initial F}^- = 0.067\, \text{M

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