A 100. 0 Ml Sample Of 0. 20 M Hf Is Titrated With 0. 10 M Koh. Determine The Ph Of The Solution After

A 100. 0 Ml Sample Of 0. 20 M Hf Is Titrated With 0. 10 M Koh. Determine The Ph Of The Solution After the titration process involves understanding the chemical properties of hafnium (Hf), its reaction with potassium hydroxide (KOH), and the resulting pH changes during titration. This comprehensive guide explains the step-by-step process to determine the pH after titration, covering concepts from acid-base chemistry, titration calculations, to the interpretation of titration curves. Whether you’re a student preparing for exams or a chemistry enthusiast seeking to deepen your understanding, this article provides detailed insights into this titration scenario.

Understanding the Components of the Titration

Before diving into calculations, it's essential to understand the nature of the substances involved and their roles in the titration process.

What is Hafnium (Hf)?

Hafnium is a transition metal with atomic number 72, and it typically exists in the +4 oxidation state. Hafnium compounds, such as hafnium fluoride (HfF_4), are often used in advanced materials and nuclear reactors. When considering Hf in aqueous solutions, it often forms hafnium ions (Hf^{4+}), which can hydrolyze in water, affecting the solution's acidity.

The Nature of Hafnium in Aqueous Solution

Hafnium(IV) ions tend to hydrolyze in water, forming various hydroxo complexes. The hydrolysis can be represented as:

$Hf^{4+} + xH_2O \rightleftharpoons [Hf(OH)_x]^{4-x} + xH^+$

This hydrolysis releases protons (H^+), making the solution slightly acidic, especially at higher concentrations.

Potassium Hydroxide (KOH)

KOH is a strong base that dissociates completely in water:

$KOH \rightarrow K^+ + OH^-$

During titration, KOH neutralizes the acid (or hydrolyzed species) from hafnium, and the titration aims to determine the point at which neutralization occurs, known as the equivalence point.

Titration Process Overview

The titration involves gradually adding KOH to the hafnium solution until the neutralization of the hydrolyzed hafnium species is complete. The key points of interest are:

- The initial concentration and volume of the hafnium solution.
- The concentration and volume of KOH added.
- The pH at various points, especially after reaching the equivalence point.

Understanding how to calculate the pH after titration requires analyzing the reactions and the resulting solution chemistry at various stages.

Step 1: Calculating the Moles of Hafnium and Potassium Hydroxide

Moles of Hafnium in the Sample

Given:
- Volume of Hf solution = 100.0 mL = 0.100 L
- Concentration of Hf = 0.20 M

Number of moles of Hf:

$$ n_{Hf} = M \times V = 0.20\, \text{mol/L} \times 0.100\, \text{L} = 0.020\, \text{mol} $$

Moles of KOH Needed for Neutralization

Given:
- Concentration of KOH = 0.10 M

The number of moles of KOH required to reach the equivalence point depends on the total amount of acid (including hydrolyzed Hf species) in the solution. Since hafnium hydrolyzes, it doesn't behave as a simple strong acid, but for simplification, the hydrolysis releases H^+ ions proportional to the amount of hafnium present, which will be neutralized by KOH.

Step 2: Understanding Hydrolysis and Acidic Nature of Hf(IV)

Hafnium(IV) ions hydrolyze in water, producing a slightly acidic solution even before titration begins. The hydrolysis constant (K_h) for Hf(IV) is relatively high, indicating that the solution is acidic.

The hydrolysis equilibrium:

$Hf^{4+} + H_2O \rightleftharpoons Hf(OH)^{3+} + H^+$

This process results in a certain concentration of free H^+ ions, which influences the initial pH.

Step 3: Determining the Initial pH of the Hafnium Solution

Since Hf hydrolyzes, the initial pH can be approximated based on the hydrolysis equilibrium constants. However, precise calculation involves complex equilibria. For simplicity, if we assume the solution is weakly acidic due to hydrolysis, the initial pH might be around 3-4.

Alternatively, if initial pH data is known or measured, it can be used directly. For demonstration, assume an initial pH of 3.5.

Step 4: Titration and the Equivalence Point

The Nature of the Equivalence Point

Because hafnium hydrolyzes and forms complexes, the equivalence point may not coincide exactly with neutralization of free H^+ ions but will occur at a pH influenced by the hydrolyzed species and their residual acidity.

In many cases with metal hydrolyzing cations, the equivalence point occurs in the slightly basic pH range (around 7-9).

Calculating the Volume of KOH at Equivalence

Total moles of Hf:

\[0.020\, \text{mol} \]

Since KOH neutralizes the acid from hydrolyzed hafnium species:

\[\text{Moles of KOH needed} \approx 0.020\, \text{mol} \]

Volume of KOH required:

\[V_{KOH} = \frac{n_{KOH}}{M_{KOH}} = \frac{0.020\, \text{mol}}{0.10\, \text{mol/L}} = 0.200\, \text{L} = 200\, \text{mL} \]

This is the volume needed to reach the equivalence point.

Step 5: pH After Titration – What Happens?

The pH after titration depends on how much KOH has been added relative to the equivalence point:

- Before the equivalence point: the solution is still acidic, and pH increases gradually with KOH addition.
- At the equivalence point: the solution is at a neutral or slightly basic pH, depending on hydrolysis.
- Beyond the equivalence point: excess KOH makes the solution basic, and pH increases significantly.

Calculating pH After Partial Titration

Suppose you've added a specific volume of KOH, say 150 mL, which is less than the 200 mL required to reach the equivalence point.

Moles of KOH added:

$$ n_{KOH} = 0.10\, \text{mol/L} \times 0.150\, \text{L} = 0.015\, \text{mol} $$

Remaining acid (from hydrolyzed Hf):

$$ n_{Hf,remaining} = 0.020\, \text{mol} - 0.015\, \text{mol} = 0.005\, \text{mol} $$

The solution still contains some acidic species, and the pH can be estimated based on the residual acid and the amount of base added.

Step 6: Calculating pH at the Half-Equivalence Point

At the half-equivalence point (adding half the volume needed to reach equivalence), the pH can be approximated by the pKa of the hydrolyzed hafnium species, since the concentrations of acid and its conjugate base are equal.

$$ pH \approx pK_a $$

If data on the hydrolysis constants are available, the pKa of Hf hydrolysis can be used to estimate the pH.

Step 7: Estimating pH at the Equivalence Point

Given the hydrolysis of Hf(IV), the equivalence point typically occurs at a pH slightly above 7, often around 7.5 to 8.0, due to residual hydroxo complexes.

Assuming the equivalence point pH is approximately 8.0, the solution contains a mixture of hafnium hydroxo complexes and water.

Step 8: Calculating pH After the Titration

Assuming you've added more than 200 mL of KOH, excess KOH determines the pH. The concentration of excess OH^- ions:

$$ \text{Excess KOH} = \text{Total moles added} - \text{Moles at equivalence} $$

Suppose total KOH added:

$$ V_{total} = 250\, \text{mL} = 0.250\, \text{L} $$

Moles of KOH:

$$ 0.10\, \text{mol/L} \times 0.250\, \text{L} = 0.025\, \text{mol} $$

Excess moles:

$$ 0.025\, \text{mol} - 0.020\, \text{mol} = 0.005\, \text{mol} $$

Concentration of excess OH^-:

$$ [OH^-] = \frac{0.005\, \text{mol}}{0.250\, \text{L}} = 0.020\, \text{M} $$

pOH:

\[pOH

Frequently Asked Questions

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

Hf is a weak acid, and its initial pH can be calculated using its concentration and acid dissociation constant. Since the dissociation is partial, pH ≈ 2.4 to 2.5, indicating an acidic solution.

At what point during titration with KOH does the solution reach the equivalence point?

The equivalence point is reached when moles of KOH added equal moles of Hf initially present. Given 100.0 mL of 0.20 M Hf, the moles of Hf are 0.020 mol. Since KOH is 0.10 M, the volume needed is 0.020 mol / 0.10 mol/L = 0.200 L or 200 mL.

How does the pH change during the titration of a weak acid like Hf with a strong base like KOH?

Initially, the pH is low; as KOH is added, the pH gradually increases. Near the equivalence point, the pH rises rapidly. After surpassing the equivalence point, the pH levels off in a basic range due to excess KOH.

What is the approximate pH of the solution after adding 100 mL of 0.10 M KOH during the titration?

At 100 mL of KOH added, half of the Hf has been neutralized, creating a buffer solution of weak acid and its conjugate base. The pH at this point is approximately equal to the pKa of Hf, which is around 2.5, but slightly higher due to partial hydrolysis.

How can the pH be calculated after the titration reaches the equivalence point?

After the equivalence point, excess KOH determines the pH. The moles of excess KOH are total KOH added minus initial moles of Hf, and pH can be calculated using the concentration of excess OH^-: $pOH = -\log[OH^-]$, then $pH = 14 - pOH$.

Additional Resources

A 100.0 mL sample of 0.20 M Hf is titrated with 0.10 M KOH. Determine the pH of the solution after is a classic titration problem that combines principles of acid-base chemistry, stoichiometry, and equilibrium calculations. Understanding how to approach such a problem not only deepens your grasp of titration concepts but also enhances your ability to analyze real-world laboratory scenarios. In this guide, we will walk through a comprehensive, step-by-step process to determine the pH of the solution after a specific volume of KOH has been added during the titration of hafnium (Hf) with potassium hydroxide (KOH).

Understanding the Titration Setup

The Basics of the Titration

- Sample Size: 100.0 mL of 0.20 M Hf
- Titrant: 0.10 M KOH
- Objective: Find the pH after a certain amount of KOH has been added

This titration involves a weak acid (or in some cases, a weak polyprotic acid or hydrolyzed species) reacting with a strong base. Hafnium (Hf) typically forms complex ions and can act as a weak acid under certain conditions, especially in aqueous solutions. For simplicity, we'll assume Hf exists predominantly as a hydrolyzed cation, such as Hf^{4+}, which can undergo hydrolysis, or as a neutral species that can be neutralized with KOH.

Note: The problem's wording does not specify the volume of KOH added, so to provide a complete guide, we'll analyze different stages of titration:

- Before any KOH is added (initial solution)
- After partial addition of KOH
- At the equivalence point
- After excess KOH is added

For this guide, we will focus primarily on the scenario after a specific volume of KOH has been added — say, after adding V mL of 0.10 M KOH. You can adapt this method for any volume.

Step 1: Calculating Moles of Hf and KOH

Moles of Hf in the Sample

Given:

- Volume of Hf solution = 100.0 mL = 0.100 L
- Concentration of Hf = 0.20 M

Calculation:

$$ \text{moles of Hf} = 0.20\, \text{mol/L} \times 0.100\, \text{L} = 0.020\, \text{mol} $$

Moles of KOH Added

Suppose we add V mL of 0.10 M KOH:

- Volume of KOH added = V mL = V/1000 L
- Concentration of KOH = 0.10 mol/L

Calculation:

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

Step 2: Determine the Reaction Progress

The Reaction Between Hf and KOH

Hafnium ions can undergo hydrolysis, and in the presence of excess base, they can form complexes or be neutralized.

A simplified neutralization reaction:

$$ \text{Hf}^{4+} + 4\, \text{OH}^{-} \rightarrow \text{Hf(OH)}_4 $$

Assumption: For this problem, assume Hf^{4+} is the dominant species and reacts with hydroxide ions in a 1:4 molar ratio.

Calculating the Moles of Hf Neutralized

- Moles of KOH added: $0.0001 \times V$
- Moles of Hf neutralized:

$$ \text{Hf neutralized} = \frac{\text{moles of OH}^{-}}{4} = \frac{0.0001\, V}{4} = 0.000025\, V $$

Remaining Moles of Hf

- Initial moles of Hf: 0.020 mol
- Moles neutralized: $0.000025\, V$

Remaining moles of Hf:

$$ 0.020 - 0.000025\, V $$

Step 3: Determine the Nature of the Solution After Partial Titration

Key Scenarios:

1. Before the equivalence point: Moles of KOH added are less than needed to neutralize all Hf.
2. At the equivalence point: Moles of KOH added exactly neutralize all Hf.
3. Beyond the equivalence point: Excess KOH remains in solution, making it strongly basic.

Calculating the Equivalence Point

- Total moles of Hf: 0.020 mol
- Moles of KOH needed for neutralization:

$$ 0.020\, \text{mol} \times 4 = 0.080\, \text{mol} $$

- Volume of KOH required:

$$ \frac{0.080\, \text{mol}}{0.10\, \text{mol/L}} = 0.80\, \text{L} = 800\, \text{mL} $$

Note: This indicates that to reach the equivalence point, 800 mL of 0.10 M

KOH must be added.

Step 4: Calculating pH at Different Stages

A. Before the Equivalence Point

Suppose V mL of KOH has been added, with V < 800 mL.

- Moles of KOH added: $0.0001\, V$
- Moles of Hf neutralized:

$$\frac{0.0001\, V}{4} = 0.000025\, V$$

- Remaining moles of Hf:

$$0.020 - 0.000025\, V$$

- Moles of hydroxide ions added:

$$0.0001\, V$$

- Moles of hydroxide ions available:

- From added KOH: $0.0001\, V$
- From hydrolysis or other equilibria (if applicable)—depends on the chemistry specifics.

Assuming the Hf solution behaves as a weak acid or complex that partially hydrolyzes, the pH calculation involves:

1. Calculating the concentration of free OH− ions:

$$\text{[OH}^{-}] = \frac{\text{excess moles of OH}^{-}}{\text{total volume}}$$

Total volume after addition:

$$V_{\text{total}} = 100\, \text{mL} + V\, \text{mL} = (100 + V)\, \text{mL} = 0.100 + \frac{V}{1000}\, \text{L}$$

Example Calculation:

If V=100 mL:

- Moles of KOH added: 0.01 mol
- Moles neutralized: 0.0025 mol
- Remaining moles of Hf: 0.020 - 0.0025 = 0.0175 mol
- Excess OH−: 0.01 mol
- Total volume: 0.100 + 0.100 = 0.200 L

Concentration of excess OH^-:

$$\text{[OH}^{-}] = \frac{0.01\, \text{mol}}{0.200\, \text{L}} = 0.05\, \text{M}$$

pOH:

$$\text{pOH} = -\log(0.05) \approx 1.30$$

pH:

$$\text{pH} = 14 - 1.30 = 12.70$$

This indicates a strongly basic solution, typical before reaching the equivalence point.

B. At the Equivalence Point

- Moles of KOH added: 0.080 mol
- Total volume:

$$100\, \text{mL} + 800\, \text{mL} = 900\, \text{mL} = 0.900\, \text{L}$$

- Remaining moles of Hf: zero
- All Hf has been neutralized to form neutral or hydrolyzed species.

Key Consideration:

- The solution now contains the hydrolyzed hafnium species, which can affect the pH through hydrolysis reactions.
- Hafnium's hydrolysis constants are known to be small, and the resulting solution is often acidic or slightly basic depending on hydrolysis equilibria.

Estimating pH at the Equivalence Point

- If the hydrolyzed species are acidic, the pH will be below 7.
- If they are basic, above 7.
- For Hf, which tends to hydrolyze producing acidic solutions, pH is often slightly below 7.

Simplified Approach:

- Use hydrolysis constants or data from literature to determine the pH.
- Alternatively, assume the hydrolyzed species produce an acidic solution, estimate pH around 4–6.

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C. After the Equivalence Point

Suppose V > 800 mL:

- Excess moles of KOH
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