

A 4.70 Ml Sample Of An H_3PO_4 Solution Of Unknown Concentration Is Titrated With A 1.050102 M NaOH Solution.

Introduction

A 4.70 mL sample of an H_3PO_4 solution of unknown concentration is titrated with a 1.050102 M NaOH solution. This fundamental laboratory procedure exemplifies how titrations are employed to determine the concentration of unknown solutions—a critical skill in analytical chemistry. Phosphoric acid (H_3PO_4), a triprotic acid, is widely used in various industrial applications, including fertilizers, food additives, and cleaning agents. Accurate determination of its concentration is essential for quality control and process optimization.

In this article, we will explore the detailed process of titrating an unknown concentration of phosphoric acid with a sodium hydroxide (NaOH) solution of known molarity. We will delve into the chemical principles behind titration, step-by-step calculation methods, and the significance of precise molarity measurements. This comprehensive overview aims to provide clarity for students, chemists, and industry professionals seeking a thorough understanding of titration techniques involving polyprotic acids like H_3PO_4 .

Understanding the Chemistry of Phosphoric Acid and Sodium Hydroxide

Properties of Phosphoric Acid (H_3PO_4)

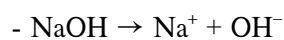
Phosphoric acid is a triprotic acid, meaning it has three ionizable hydrogen atoms:

- First dissociation: $\text{H}_3\text{PO}_4 \rightleftharpoons \text{H}^+ + \text{H}_2\text{PO}_4^-$
- Second dissociation: $\text{H}_2\text{PO}_4^- \rightleftharpoons \text{H}^+ + \text{HPO}_4^{2-}$
- Third dissociation: $\text{HPO}_4^{2-} \rightleftharpoons \text{H}^+ + \text{PO}_4^{3-}$

The acid dissociation constants (K_a values) for these steps indicate the extent of ionization at each stage. The first dissociation is strong compared to the subsequent ones, but all are significant in titration calculations.

Properties of Sodium Hydroxide (NaOH)

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



During titration, the hydroxide ions (OH^-) react with the hydrogen ions (H^+) from the acid to form water:



The complete dissociation of NaOH simplifies the calculation process since molarity directly relates to the number of moles of OH^- available for reaction.

The Titration Process: Step-by-Step

Preparation of Reagents

- Standard NaOH Solution: The molarity (1.050102 M) is precisely known, which is crucial for accurate calculations.
- Unknown H_3PO_4 Solution: The sample volume is 4.70 mL, and its concentration is initially unknown.

Conducting the Titration

1. Sample Measurement: Using a pipette, transfer 4.70 mL of the phosphoric acid solution into a clean titration flask.
2. Indicator Addition: Add a suitable pH indicator, such as phenolphthalein, which changes color around pH 8.2–10.0. For phosphoric acid, phenolphthalein is commonly used, but other indicators like methyl orange can also be employed depending on the titration endpoint.
3. Titrant Addition: Slowly add the NaOH solution from a burette to the acid sample while swirling continuously.
4. Endpoint Detection: The titration continues until a persistent color change indicates the completion of the reaction. For phenolphthalein, this occurs when the solution turns faint pink.

Data Recording and Calculation

- Record the volume of NaOH used to reach the endpoint (V_{NaOH}).
- Use the known molarity (M_{NaOH}) and volume (V_{NaOH}) to calculate the moles of NaOH reacted:

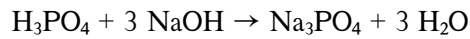
$$\text{Moles NaOH} = M_{\text{NaOH}} \times V_{\text{NaOH}} \text{ (in liters)}$$

- Since the reaction involves multiple proton transfers, consider the stoichiometry of the reaction to determine the moles of H_3PO_4 .

Stoichiometry of the Titration Reaction

Balanced Reaction Equation

The titration involves the neutralization of phosphoric acid's three protons with sodium hydroxide:



This indicates a 1:3 molar ratio between H_3PO_4 and NaOH .

Calculating the Concentration of the Unknown H_3PO_4 Solution

1. Determine Moles of NaOH used:

$$\text{Moles NaOH} = 1.050102 \text{ mol/L} \times V_{\text{NaOH}} \text{ (in liters)}$$

2. Determine Moles of H_3PO_4 :

Since the molar ratio is 1:3:

$$\text{Moles H}_3\text{PO}_4 = (\text{Moles NaOH}) / 3$$

3. Calculate Concentration of H_3PO_4 :

$$\text{Concentration (M}_{\text{H}_3\text{PO}_4}) = \text{Moles H}_3\text{PO}_4 / \text{volume of sample in liters}$$

$$\text{- Volume of sample} = 4.70 \text{ mL} = 0.00470 \text{ L}$$

Final formula:

$$\boxed{\text{Concentration of H}_3\text{PO}_4 = \frac{(\text{M}_{\text{NaOH}} \times V_{\text{NaOH}}) / 3}{V_{\text{sample}}}}$$

Where:

- $(\text{M}_{\text{NaOH}} = 1.050102 \text{ mol/L})$
- (V_{NaOH}) = volume of NaOH used in liters
- $(V_{\text{sample}} = 0.00470 \text{ L})$

Practical Considerations and Error Analysis

Precision of Measurements

- Use calibrated pipettes and burettes to minimize volumetric errors.
- Perform multiple titrations to obtain consistent endpoint volumes and calculate an average.

Indicators and Endpoint Detection

- Choose an indicator suitable for the titration's pH range.
- Be vigilant in observing color change to avoid overshooting the endpoint.

Dealing with Polyprotic Acid Titration Complexity

Since H_3PO_4 is triprotic, partial neutralizations can occur at different pH levels, complicating endpoint detection. Typically, titrations are performed with a suitable endpoint indicator or using potentiometric methods for higher accuracy.

Example Calculation

Suppose during a titration, 25.00 mL of NaOH solution was required to reach the endpoint.

Calculations:

- Convert volume to liters: $25.00 \text{ mL} = 0.02500 \text{ L}$
- Moles of NaOH: $1.050102 \text{ mol/L} \times 0.02500 \text{ L} = 0.02625255 \text{ mol}$
- Moles of H_3PO_4 : $0.02625255 \text{ mol} / 3 \approx 0.00875085 \text{ mol}$
- Concentration of H_3PO_4 :

$$\left[\frac{0.00875085 \text{ mol}}{0.00470 \text{ L}} \right] \approx 1.863 \text{ M}$$

Thus, the unknown phosphoric acid solution has an approximate concentration of 1.86 M.

Note: The actual concentration will vary depending on the titration data, but this example illustrates the calculation process.

Significance of Accurate Titration in Industry and Research

Accurate titration methods are vital across various sectors:

- Pharmaceutical Industry: Ensuring correct dosages of acids and bases.
- Agriculture: Determining fertilizer compositions, such as phosphates.
- Food Industry: Analyzing acidity in products.
- Environmental Monitoring: Detecting pollutants and assessing water quality.

The precision of molarity measurements, especially in solutions like NaOH, is crucial for reliable results. Advanced techniques, such as potentiometric titrations, can further enhance accuracy, especially for polyprotic acids.

Conclusion

Titration of an unknown concentration of phosphoric acid with a standardized NaOH solution exemplifies core principles of analytical chemistry. The process involves careful preparation, precise measurement, and understanding of chemical stoichiometry. By knowing the molarity of NaOH and accurately recording the titration volume, chemists can determine the concentration of the phosphoric acid sample with high confidence.

This method underscores the importance of meticulous laboratory techniques and reinforces fundamental concepts such as molarity, stoichiometry, and acid-base reactions. Whether in academic settings or industrial applications, such titration procedures remain indispensable tools for chemical analysis and quality assurance.

Key Takeaways:

- Phosphoric acid is triprotic; titration requires considering all ionizable protons.
- Accurate molarity of titrant (NaOH) is essential for reliable results.
- The 1:3 molar ratio guides calculations for unknown acid concentration.
- Proper endpoint detection minimizes titration errors.
- Repetition and averaging improve the accuracy of the determined concentration.

By mastering these concepts, practitioners can confidently analyze acid solutions and ensure the integrity of their chemical processes and research findings.

Frequently Asked Questions

What is the purpose of titrating an H_3PO_4 solution with NaOH in this experiment?

The purpose is to determine the unknown concentration of the H_3PO_4 solution by reacting it completely

with a known concentration of NaOH and measuring the volume used.

How do you calculate the molarity of the H₃PO₄ solution after titration?

Using the titration data, you calculate the moles of NaOH used (molarity × volume), then use the balanced chemical equation to find the moles of H₃PO₄, and divide by the volume of the H₃PO₄ sample to find its molarity.

What is the balanced chemical equation for the titration of H₃PO₄ with NaOH?

The balanced equation is $\text{H}_3\text{PO}_4 + 3 \text{NaOH} \rightarrow \text{Na}_3\text{PO}_4 + 3 \text{H}_2\text{O}$, indicating that 1 mole of phosphoric acid reacts with 3 moles of NaOH.

How does the triprotic nature of H₃PO₄ affect the titration process?

Since H₃PO₄ can donate three protons, the titration involves three equivalence points, but typically, the first or combined points are used to determine concentration, requiring careful endpoint detection.

Why is it important to use a precise molarity of NaOH in this titration?

A precise molarity ensures accurate calculation of the unknown H₃PO₄ concentration, as any errors in NaOH concentration directly affect the calculation results.

What indicators are suitable for titrating phosphoric acid with NaOH?

Indicators such as phenolphthalein or methyl orange can be used, with phenolphthalein often preferred for detecting the neutralization endpoint in this titration.

How can temperature affect the titration results in this experiment?

Temperature can influence the volume measurements and the reaction rate; maintaining a constant temperature minimizes errors and improves the accuracy of the titration.

Additional Resources

A 4.70 mL sample of an H₃PO₄ solution of unknown concentration is titrated with a 1.050102 M NaOH solution—a classic titration scenario that offers insight into acid-base chemistry, stoichiometry, and analytical techniques. This process not only helps determine the unknown concentration of phosphoric acid (H₃PO₄) but also reinforces foundational principles such as molarity calculations, balanced chemical equations, and titration methodology.

In this comprehensive guide, we'll explore every facet of this titration, from understanding the chemical reactions involved to performing calculations that reveal the concentration of the unknown acid solution. Whether you're a student seeking clarity or a professional brushing up on titration procedures, this article offers a detailed walkthrough of the process.

Understanding the Chemical Context

What is Titration?

Titration is an analytical laboratory technique used to determine the concentration of an unknown solution by reacting it with a solution of known concentration (the titrant). The process involves adding the titrant gradually until the reaction reaches its equivalence point, where molar quantities of acid and base have reacted completely.

The Components in Our Scenario

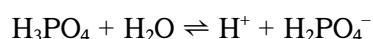
- Sample of unknown concentration of H_3PO_4 (phosphoric acid): A triprotic acid often used in fertilizers, food additives, and industrial applications.
- Titrant: Sodium hydroxide (NaOH) with a known molarity of 1.050102 M.
- Sample volume: 4.70 mL.

Step 1: Understanding the Chemistry of the Titration

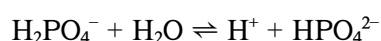
The Acid-Base Reaction

Phosphoric acid (H_3PO_4) is a triprotic acid, meaning it can donate three protons (H^+) per molecule. Its dissociation can be represented stepwise:

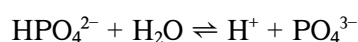
1. First dissociation:



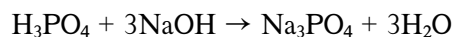
2. Second dissociation:



3. Third dissociation:



However, in titration with NaOH , the titrant reacts primarily with the free H^+ ions released from each dissociation step. For simplicity, the overall neutralization can be considered as:



This balanced equation indicates that one mole of phosphoric acid reacts with three moles of sodium hydroxide.

Step 2: Calculating the Moles of NaOH Used

Suppose during titration, a certain volume of NaOH solution was required to reach the endpoint. Since the molarity (concentration) of NaOH is known, calculating moles of NaOH is straightforward:

Moles of NaOH (n_{NaOH}):

$$n_{\text{NaOH}} = M_{\text{NaOH}} \times V_{\text{NaOH}}$$

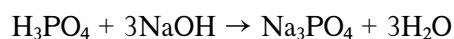
Where:

- $M_{\text{NaOH}} = 1.050102 \text{ mol/L}$
- $V_{\text{NaOH}} = \text{volume of NaOH used in liters}$

Note: The actual volume of NaOH used must be measured during titration to perform this calculation.

Step 3: Determining the Moles of H_3PO_4

Given the stoichiometry:



The molar ratio of H_3PO_4 to NaOH is 1:3.

Therefore:

Moles of H_3PO_4 ($n_{\text{H}_3\text{PO}_4}$):

$$n_{\text{H}_3\text{PO}_4} = n_{\text{NaOH}} / 3$$

This calculation allows us to determine the amount of phosphoric acid present in the original 4.70 mL sample.

Step 4: Calculating the Concentration of the Unknown H_3PO_4 Solution

Once the moles of H_3PO_4 are known, its concentration (molarity) can be calculated:

Concentration (M) of H_3PO_4 :

$$M_{\text{H}_3\text{PO}_4} = n_{\text{H}_3\text{PO}_4} / V_{\text{sample}}$$

Where:

$$- V_{\text{sample}} = 4.70 \text{ mL} = 0.00470 \text{ L}$$

This yields the molarity of the unknown phosphoric acid solution.

Practical Example: Hypothetical Data and Calculations

Let's walk through a hypothetical example with assumed titration data to illustrate the process.

Assumed Data:

- Volume of NaOH used to reach equivalence point: 12.00 mL (0.01200 L)
- Molarity of NaOH: 1.050102 mol/L
- Volume of H_3PO_4 sample: 4.70 mL (0.00470 L)

Step-by-step Calculation:

1. Moles of NaOH used:

$$\begin{aligned} n_{\text{NaOH}} &= M_{\text{NaOH}} \times V_{\text{NaOH}} \\ &= 1.050102 \text{ mol/L} \times 0.01200 \text{ L} \\ &\approx 0.0126012 \text{ mol} \end{aligned}$$

2. Moles of H_3PO_4 (from stoichiometry):

$$\begin{aligned} n_{\text{H}_3\text{PO}_4} &= n_{\text{NaOH}} / 3 \\ &\approx 0.0126012 \text{ mol} / 3 \\ &\approx 0.0042004 \text{ mol} \end{aligned}$$

3. Concentration of H_3PO_4 :

$$\begin{aligned} M_{\text{H}_3\text{PO}_4} &= n_{\text{H}_3\text{PO}_4} / V_{\text{sample}} \\ &= 0.0042004 \text{ mol} / 0.00470 \text{ L} \end{aligned}$$

$\approx 0.894 \text{ mol/L}$

Result: The unknown phosphoric acid solution has an approximate concentration of 0.894 M.

Important Considerations and Tips

Accuracy in Titration

- Use precise volumetric equipment such as burettes and pipettes.
- Ensure proper endpoint detection—either via pH indicator or pH meter.
- Conduct multiple trials to obtain consistent data.

Recognizing Acid-Base Behavior

- Phosphoric acid's multiple dissociation steps mean that titration curves can have multiple inflection points.
- For precise determination, especially in complex scenarios, a pH meter is preferable over indicators.

Handling the Data

- Always record the exact volume of titrant used.
- Convert all measurements to consistent units (liters for volume, mols for moles).

Additional Insights

Influence of pKa Values

Phosphoric acid has three pKa values (~ 2.15 , 7.20 , 12.35). These influence how the titration curve appears and the choice of indicators if used. For the first dissociation, an indicator that changes color around pH 3–4 is suitable.

Limitations and Error Sources

- Incomplete reactions or overshooting the endpoint can lead to inaccuracies.
- Impurities in reagents may affect results.
- Temperature variations influence molarity and dissociation equilibria.

Summary

- The titration of a 4.70 mL sample of an unknown H_3PO_4 solution with a 1.050102 M NaOH solution involves understanding the triprotic nature of phosphoric acid and its neutralization reaction.
- By measuring the volume of NaOH required to reach the endpoint, calculating moles of NaOH, and applying stoichiometry, the concentration of the unknown acid solution can be accurately determined.
- Proper technique, precise measurements, and understanding of acid-base reactions are essential for reliable results.

This process exemplifies fundamental principles in analytical chemistry, showcasing how titrations serve as powerful tools for solution characterization and quality control. Whether for academic purposes or industrial applications, mastering such techniques is vital for chemists and laboratory professionals alike.

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