

A Student Titrated A 25.00 ML Sample Of HCl With An Unknown Concentration With 0.150M NaOH. Phenolphthalein

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Performing titration experiments is a fundamental skill in analytical chemistry, allowing students and scientists to determine the concentration of unknown solutions accurately. In a typical laboratory setting, a student might titrate a 25.00 mL sample of hydrochloric acid (HCl) with a known concentration of sodium hydroxide (NaOH), using phenolphthalein as an indicator. This process not only helps in understanding acid-base chemistry but also reinforces concepts like molarity, stoichiometry, and data analysis. In this article, we will explore the detailed steps involved in such a titration, the principles behind it, and how to interpret the results effectively.

Understanding the Titration Process

What is Titration?

Titration is a laboratory technique used to determine the concentration of an unknown solution by reacting it with a solution of known concentration. The process involves adding the titrant (NaOH in this case) gradually to the analyte (HCl) until the reaction reaches its equivalence point, where the amount of acid equals the amount of base.

Significance of Using Phenolphthalein as an Indicator

Phenolphthalein is a common acid-base indicator that changes color at a pH range of approximately 8.2 to 10.0. It is colorless in acidic solutions and turns pink in basic solutions. During the titration of HCl with NaOH, phenolphthalein allows the student to visually identify the endpoint—the point at which all the acid has been neutralized, and the solution begins to turn faintly pink.

Materials and Methods

Materials Needed

- 25.00 mL sample of hydrochloric acid (HCl) with unknown concentration
- 0.150 M sodium hydroxide (NaOH) solution

- Phenolphthalein indicator
- Burette
- Conical flask (Erlenmeyer flask)
- Graduated cylinder
- Distilled water
- Clamp stand and burette holder
- Wash bottle

Procedure

1. Fill the burette with 0.150 M NaOH solution, ensuring no air bubbles are present.
2. Use a pipette to transfer exactly 25.00 mL of the HCl sample into the conical flask.
3. Add 2-3 drops of phenolphthalein indicator to the acid sample.
4. Gradually add the NaOH from the burette to the acid while swirling the flask continuously to mix.
5. Observe the color change. The solution will remain colorless in acidic conditions and turn faint pink as the endpoint approaches.
6. Stop adding NaOH as soon as a persistent light pink color appears, indicating the endpoint has been reached.
7. Record the volume of NaOH used to reach the endpoint.
8. Repeat the titration multiple times to obtain consistent readings, and calculate the average volume of NaOH used.

Calculations and Data Analysis

Determining the Molarity of HCl

Once the titration data is collected, the primary goal is to determine the unknown concentration of the HCl sample. The key steps involve stoichiometric calculations based on the titration data.

Step-by-Step Calculation

- Identify the average volume of NaOH used at the endpoint (e.g., 30.50 mL).
- Convert the volume of NaOH to liters:
 $\text{Volume (L)} = 30.50 \text{ mL} \times (1 \text{ L} / 1000 \text{ mL}) = 0.03050 \text{ L}$
- Calculate the moles of NaOH used:
 $\text{Moles NaOH} = \text{Molarity} \times \text{Volume} = 0.150 \text{ mol/L} \times 0.03050 \text{ L} = 0.004575 \text{ mol}$
- Use the molar ratio from the balanced chemical equation:
 $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
1 mol HCl reacts with 1 mol NaOH
- Calculate the moles of HCl in the 25.00 mL sample:
 $\text{Moles HCl} = \text{Moles NaOH} = 0.004575 \text{ mol}$
- Determine the molarity of HCl:
 $\text{Molarity (M)} = \text{Moles} / \text{Volume (L)} = 0.004575 \text{ mol} / 0.02500 \text{ L} = 0.183 \text{ mol/L}$

Final Result

The concentration of the unknown HCl sample is approximately 0.183 M, based on this particular titration. Repeating the experiment and averaging multiple titrations increase the accuracy and reliability of this value.

Importance of Accuracy and Precision

Ensuring Precise Measurements

Accurate titration results depend on careful measurement techniques:

- Proper calibration of the burette and pipette.
- Consistent swirling of the solution during titration.
- Adding the titrant slowly near the endpoint to avoid overshooting.

Reproducibility and Replication

Repeating the titration multiple times and calculating the average volume helps minimize errors and improve confidence in the results. Discrepancies can arise from:

- Incomplete mixing of solutions.
- Inconsistent endpoint detection.
- Measurement inaccuracies.

Applications and Real-World Significance

Educational Value of Titration Experiments

Titration experiments like this one offer students practical experience with fundamental analytical techniques, reinforcing theoretical concepts learned in classroom lectures.

Industrial and Environmental Uses

Accurate concentration determination of acids and bases is crucial across various industries:

- Manufacturing of pharmaceuticals and chemicals.
- Water quality testing and environmental monitoring.
- Food industry, ensuring proper acidity levels.

Conclusion

Titration of a 25.00 mL sample of hydrochloric acid with 0.150 M NaOH, using phenolphthalein as an indicator, is a classic experiment that demonstrates fundamental principles of acid-base chemistry. By carefully measuring the volume of NaOH used to neutralize the acid, students can calculate the unknown concentration of HCl with high accuracy. This process emphasizes the importance of precise techniques, proper data analysis, and understanding the reaction stoichiometry. Whether for educational purposes or industrial applications, titrations remain an essential method for quantitative chemical analysis, providing vital insights into substance concentrations in various contexts.

Frequently Asked Questions

How is the concentration of HCl determined using titration with NaOH and phenolphthalein?

The concentration of HCl is determined by adding a known volume of NaOH of known molarity to the HCl sample until the endpoint is reached, indicated by a color change of phenolphthalein. Using the

volume of NaOH used and its molarity, the molarity of HCl can be calculated via the titration formula.

What is the significance of using phenolphthalein in this titration?

Phenolphthalein acts as an indicator that signals the endpoint of the titration by changing color from colorless to faint pink as the solution transitions from acidic to slightly basic, allowing precise detection of the titration completion.

How do you calculate the molarity of HCl after titration?

The molarity of HCl is calculated using the formula $M_1V_1 = M_2V_2$, where M_1 and V_1 are the molarity and volume of HCl, and M_2 and V_2 are the molarity and volume of NaOH used. Rearranged as $M_1 = (M_2 \times V_2) / V_1$, substituting the known values gives the unknown molarity of HCl.

Why is it important to record the volume of NaOH used at the endpoint?

Recording the volume of NaOH used is essential because it directly relates to the amount of HCl in the sample. Accurate volume measurement allows for precise calculation of the HCl's concentration.

What are common sources of error in this titration procedure?

Common errors include misreading the burette, overshooting the endpoint, inconsistent mixing, using impure reagents, or not accurately measuring the sample volume, all of which can affect the calculation accuracy.

How does the unknown concentration of HCl impact the titration process?

Since the concentration of HCl is unknown, titration helps determine it precisely. The process involves adding a titrant of known concentration until the endpoint, enabling calculation of the HCl's molarity based on the titration data.

What precautions should be taken during the titration with phenolphthalein?

Precautions include adding the titrant slowly near the endpoint, swirling the solution continuously for uniform mixing, accurately reading the burette, and ensuring the titrant and sample are at proper temperatures to avoid errors.

How can repeated titrations improve accuracy in determining HCl concentration?

Performing multiple titrations and calculating the average volume used reduces random errors, increases reliability of results, and provides a more precise determination of the HCl concentration.

Additional Resources

Student Titrates 25.00 mL Sample of HCl with 0.150 M NaOH Using Phenolphthalein: An Analytical Perspective

In the realm of analytical chemistry, titration remains one of the most fundamental and widely employed techniques for determining the concentration of an unknown solution. A student's experiment involving the titration of a 25.00 mL sample of hydrochloric acid (HCl) with a standardized sodium hydroxide (NaOH) solution exemplifies the core principles of acid-base chemistry, stoichiometry, and titration methodology. The use of phenolphthalein as an indicator further refines the process, providing a clear visual cue at the equivalence point. This article delves into the intricacies of this titration experiment, exploring its theoretical underpinnings, procedural execution, data analysis, and broader implications within chemical analysis.

Understanding the Fundamentals of Acid-Base Titration

The Core Principles of Titration

Titration is an analytical technique that involves the slow addition of a titrant— a solution of known concentration— to a analyte of unknown concentration until the reaction reaches a specific endpoint. The primary goal is to determine the unknown concentration based on the volume of titrant used.

In this scenario, the student aims to determine the molarity of hydrochloric acid (HCl)— a strong acid— by titrating it with a sodium hydroxide (NaOH) solution of known concentration, 0.150 M. The reaction involved is a classic acid-base neutralization:



This reaction proceeds rapidly and quantitatively, meaning the acid and base react in a 1:1 molar ratio, simplifying calculations.

The Significance of Phenolphthalein as an Indicator

Indicators are vital in titration as they signal the completion of the reaction— the endpoint— often through a color change. Phenolphthalein is a popular choice for acid-base titrations due to its clear color transition: colorless in acidic solutions and pink in basic solutions. Its pH transition range is approximately 8.3 to 10.0, making it ideal for titrations where the equivalence point occurs near a neutral pH, such as strong acid-strong base reactions.

In this experiment, as NaOH is added to the HCl sample, the solution remains clear until the neutralization point is approached. When the equivalence point is reached, a faint pink color persists,

indicating that the acid has been neutralized and the solution has turned slightly basic.

Experimental Procedure and Data Collection

Preparation and Execution of Titration

The student begins by pipetting exactly 25.00 mL of the HCl sample into a clean Erlenmeyer flask. A few drops of phenolphthalein are added as the indicator. The burette containing 0.150 M NaOH is filled and carefully positioned above the flask.

The titration proceeds with steady, controlled addition of NaOH, swirling the flask continuously to ensure thorough mixing. The goal is to approach the equivalence point gradually, avoiding overshooting. As the pink color begins to appear, the student carefully adds NaOH dropwise until a persistent pale pink hue remains for at least 30 seconds, signaling the endpoint.

Recording the Data

The critical data points recorded include:

- Initial volume of NaOH in the burette (usually zero at the start)
- Final volume of NaOH after reaching the endpoint

Suppose, for example, that the student records the initial volume as 0.00 mL and the final volume as 22.50 mL. This indicates that 22.50 mL of NaOH was used to neutralize the acid.

Data Analysis and Calculation of Unknown Concentration

Calculating Moles of NaOH Used

Given the known molarity of NaOH (0.150 M) and the volume used, the number of moles of NaOH consumed can be calculated:

$$\text{Moles of NaOH} = \text{Molarity} \times \text{Volume (L)}$$

Converting milliliters to liters:

$$[22.50\, \mathrm{mL} = 0.02250\, \mathrm{L}]$$

Thus:

$$[\text{Moles of NaOH} = 0.150\, \mathrm{mol/L} \times 0.02250\, \mathrm{L} = 0.003375\, \mathrm{mol}]$$

Determining the Moles of HCl in the Sample

Since the reaction between HCl and NaOH is a 1:1 molar ratio, the moles of HCl in the 25.00 mL sample are equal to the moles of NaOH used:

$$[\text{Moles, of, HCl} = 0.003375\, \mathrm{mol}]$$

Calculating the Molarity of the HCl Sample

The molarity of the HCl is derived by dividing the moles of HCl by the volume of the acid sample:

$$[\text{Molarity, of, HCl} = \frac{\text{Moles, of, HCl}}{\text{Volume, of, HCl, in, Liters}}]$$

$$[\text{Molarity, of, HCl} = \frac{0.003375\, \mathrm{mol}}{0.02500\, \mathrm{L}} = 0.135\, \mathrm{M}]$$

Therefore, the concentration of the unknown HCl sample is approximately 0.135 M.

Analytical Considerations and Sources of Error

Precision and Accuracy in Titration

The reliability of titration hinges on precise measurement and careful technique. The student's ability to accurately read volumes, control the rate of titrant addition, and identify the endpoint directly impacts the accuracy of the calculated concentration.

Potential sources of error include:

- Parallax errors when reading the burette scale
- Incomplete mixing during titration leading to inconsistent endpoint detection
- Over- or under-shooting the endpoint due to human reaction time
- Impurities or CO₂ absorption in the sample or titrant affecting pH

Impact of Indicators and pH Transition Range

While phenolphthalein is suitable for strong acid-strong base titrations, its pH transition range means the color change occurs slightly past the true equivalence point. This can result in a slight overestimation of the titrant volume, thereby affecting the calculated concentration. Multiple titrations can improve data reliability and help identify consistent results.

Replicability and Statistical Analysis

Performing multiple titrations allows for the calculation of an average concentration and assessment of experimental uncertainty. Standard deviation and percent error provide insights into the precision of measurements.

Broader Implications and Educational Significance

Understanding Acid-Base Equilibria

This titration exemplifies fundamental concepts such as molarity, stoichiometry, and pH behavior. Recognizing the equivalence point's pH and understanding the indicator's transition range deepen conceptual comprehension of acid-base interactions.

Application in Real-World Scenarios

Accurate determination of acid concentrations has practical applications in industries such as pharmaceuticals, environmental testing, and food chemistry. For instance, analyzing the acidity of water samples or the concentration of acids in products depends on precise titrations.

Educational Value and Skill Development

Engaging in titration experiments fosters critical laboratory skills, including precise measurement, proper titration technique, data analysis, and error minimization. These competencies are foundational for students pursuing careers in chemistry and related sciences.

Conclusion

The student titration of a 25.00 mL HCl sample with 0.150 M NaOH using phenolphthalein as an indicator encapsulates the core principles of analytical chemistry: precise measurement, stoichiometric calculation, and interpretation of results. Through careful execution and data analysis, an approximate concentration of 0.135 M for the unknown HCl sample was determined, illustrating the power and importance of titration as a quantitative analytical tool. This experiment not only reinforces theoretical concepts but also cultivates essential laboratory skills, fostering a deeper understanding of chemical reactions and measurement accuracy. As a fundamental technique, titration continues to serve as a cornerstone in scientific research, quality control, and educational laboratories worldwide.

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