Question 6 A 5.00 ML Aliquot Of A 0.20 M HCl Solution Is Diluted To A Final Volume Of 25.00 ML. What

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Understanding dilution calculations is fundamental in chemistry, especially when preparing solutions of specific concentrations for experiments or industrial processes. In this article, we will explore the problem of diluting a hydrochloric acid (HCl) solution, analyze the concepts involved, and provide step-by-step solutions. This comprehensive guide is structured to enhance your understanding of dilution principles, concentration calculations, and their practical applications in laboratory settings.

Introduction to Solution Dilution

Dilution is the process of reducing the concentration of a solute in a solution, typically by adding more solvent without adding more solute. It is a common procedure in laboratories for preparing solutions at desired concentrations for various experiments. The fundamental principle guiding dilution calculations is the conservation of moles of solute.

Key Concepts in Solution Dilution

- Concentration (Molarity, M): The number of moles of solute per liter of solution.

- Moles of Solute: The amount of solute in moles, calculated as concentration times volume in liters.
- Dilution Equation: The relationship between initial and final concentrations and volumes, expressed as:

$$
C_1 V_1 = C_2 V_2
$$

Where:

- C_1 = initial concentration
- V_1 = initial volume
- C_2 = final concentration
- V_2 = final volume

Analyzing the Given Problem

The problem states:
- An aliquot (a measured portion) of 5.00 mL of a 0.20 M HCl solution.
- This aliquot is diluted to a final volume of 25.00 mL.

The question is: What is the concentration of the diluted solution?

Step-by-Step Solution Approach

To solve this problem, we will utilize the dilution equation:

$$
C_1 V_1 = C_2 V_2
$$

Where:

- C_1 = initial concentration = 0.20 M
- V_1 = volume of initial solution used = 5.00 mL
- V_2 = final volume after dilution = 25.00 mL
- C_2 = unknown final concentration

Step 1: Convert Volumes to Consistent Units

Since concentrations are in molarity (mol/L), convert volumes from milliliters to liters:

$$
V_1 = 5.00\, \text{mL} = 0.00500\, \text{L}
$$

$$
V_2 = 25.00\, \text{mL} = 0.02500\, \text{L}
$$

Step 2: Calculate the Moles of HCl in the Aliquot

The initial moles of HCl in the aliquot:

\[

\text{Moles} = C_1 \times V_1 = 0.20\, \text{mol/L} \times 0.00500\, \text{L} = 0.00100\, \text{mol}
\]

Step 3: Determine the Final Concentration

Since moles of solute are conserved during dilution:

\[
\text{Moles before} = \text{Moles after}
\]

Therefore,

\[
C_2 = \frac{\text{Moles}}{V_2} = \frac{0.00100\, \text{mol}}{0.02500\, \text{L}} = 0.04\, \text{mol/L}
\]

Final answer:

\[
\boxed{
\textbf{The concentration of the diluted HCl solution is 0.04 M}
}
\]

Understanding the Significance of the Calculation

The calculation demonstrates how a small aliquot of a concentrated acid solution can be diluted to a lower concentration suitable for various applications, such as titrations or buffer preparations. Recognizing the relationship between initial and final concentrations helps chemists control solution properties precisely.

Practical Applications of Dilution Calculations

Dilution calculations are crucial for:

- Preparing standard solutions for titrations
- Adjusting reagent concentrations
- Ensuring safety by reducing hazardous concentrations
- Achieving desired pH levels in buffer solutions
- Maintaining accuracy in analytical chemistry

Additional Considerations in Dilution Procedures

While the basic dilution formula is straightforward, practical aspects include:

- Accuracy of measurements: Use calibrated pipettes and volumetric flasks.
- Contamination avoidance: Use clean equipment to prevent contamination.
- Proper mixing: Ensure solutions are mixed thoroughly after dilution.
- Solution stability: Be aware of the stability of chemicals over time.

Common Mistakes and Tips

- Incorrect unit conversions: Always convert mL to liters when working with molarity.
- Neglecting moles conservation: Remember that moles of solute remain constant during dilution.
- Mixing units: Keep units consistent throughout calculations.
- Rounding errors: Use appropriate significant figures based on measurement precision.

Practice Problems for Mastery

To reinforce your understanding, consider practicing with additional problems:

1. A 10.0 mL sample of a 0.15 M NaOH solution is diluted to 50.0 mL. What is the concentration after dilution?
2. How much of a 0.50 M H2SO4 solution is needed to prepare 250 mL of a 0.10 M solution?
3. If you dilute 2.00 mL of a 0.10 M solution to 100.00 mL, what is the resulting concentration?

Conclusion

Understanding dilution calculations is a cornerstone of quantitative chemistry. By mastering the use of the dilution equation and related concepts, you can accurately prepare solutions with desired concentrations for various scientific and industrial applications. Remember, careful measurement, unit consistency, and thorough understanding of the principle of conservation of moles are key to successful solution preparation.

Summary

- The problem involves diluting a 0.20 M HCl solution from 5.00 mL to a final volume of 25.00 mL.
- Using the dilution equation, the final concentration is calculated as 0.04 M.
- Proper unit conversions and understanding of mole conservation are essential.
- Dilution calculations are fundamental in ensuring solution accuracy for laboratory experiments.

Keywords: solution dilution, molarity, HCl concentration, dilution equation, laboratory solutions, molar concentration, chemical calculations, preparing solutions, titration solutions, analytical chemistry

Frequently Asked Questions

What is the initial amount of moles of HCl in the 5.00 mL aliquot?

The initial moles of HCl are calculated by multiplying the concentration by the volume: 0.20 M × 5.00 mL = 0.20 mol/L × 0.005 L = 0.001 mol.

How do you calculate the new concentration of HCl after dilution?

The new concentration is found using the dilution formula: $C_1V_1 = C_2V_2$. So, $C_2 = (C_1V_1) / V_2$ = (0.20 M × 5.00 mL) / 25.00 mL = 0.04 M.

What is the significance of the dilution factor in this problem?

The dilution factor is the ratio of the initial volume to the final volume, which in this case is 25.00 mL / 5.00 mL = 5, indicating the solution is diluted fivefold, reducing the concentration accordingly.

What is the final concentration of HCl after dilution?

The final concentration of HCl is 0.04 M.

How does dilution affect the pH of the solution?

Since HCl is a strong acid, the pH is calculated as $-\log[H^+]$. For a 0.04 M HCl solution, pH = $-\log(0.04) \approx 1.40$, indicating an acidic solution.

If you wanted to prepare 50 mL of a 0.10 M HCl solution, how much of the original solution would you need?

Using $C_1V_1 = C_2V_2$: $V_1 = (C_2 \times V_2) / C_1$ = (0.10 M × 50 mL) / 0.20 M = 25 mL.

Why is it important to perform dilutions carefully in laboratory settings?

Precise dilutions ensure accurate concentrations, which are critical for reproducible results and correct interpretations in experiments.

What assumptions are made in this dilution calculation?

Assumption that the solution is ideal, the volume measurements are accurate, and the acid is completely dissociated, which is valid for strong acids like HCl.

How would the concentration change if the initial aliquot volume was doubled to 10.00 mL before dilution?

The initial moles would double to 0.002 mol, and the final concentration after dilution to 25 mL would be 0.08 M, since C_2 = (0.20 M × 10 mL) / 25 mL = 0.08 M.

Additional Resources

Understanding the Concentration Changes in Dilution: A Step-by-Step Analysis of a 5.00 mL Aliquot of 0.20 M HCl Diluted to 25.00 mL

When working in chemistry, particularly in titrations, laboratory preparations, or analytical procedures, understanding how dilutions affect concentration is fundamental. A classic example involves taking a small, known volume of a concentrated solution and diluting it to a larger volume, thereby reducing its concentration proportionally. Here, we examine Question 6: A 5.00 mL aliquot of a 0.20 M HCl solution is diluted to a final volume of 25.00 mL. This scenario exemplifies the principles of dilution, concentration calculations, and the importance of precise measurements in chemical experiments.

The Basics of Dilution and Concentration

Before delving into the specifics of the problem, it’s essential to understand the core concepts:

- Concentration (Molarity, M): The number of moles of solute per liter of solution.
- Aliquot: A measured sub-volume taken from a larger solution, often used for analysis or further dilution.
- Dilution: The process of decreasing the concentration of a solute in solution, achieved by adding solvent.

Key principle: The number of moles of solute remains constant during dilution. This means:

$$\text{Initial moles} = \text{Final moles}$$

or mathematically:

$$C_1 V_1 = C_2 V_2$$

where:

- C_1 and V_1 are the initial concentration and volume,
- C_2 and V_2 are the final concentration and volume.

Step 1: Understanding the Given Data

Let's lay out the data provided:

- Initial solution concentration (C_1): 0.20 M
- Initial aliquot volume (V_1): 5.00 mL
- Final volume after dilution (V_2): 25.00 mL

The question (though incomplete in the prompt) generally asks for the final concentration (C_2) after dilution or how many moles of HCl are present in the aliquot or the diluted solution.

Step 2: Calculating the Moles in the Initial Aliquot

Since molarity relates to moles per liter, convert the initial volume into liters:

$$
V_1 = 5.00\, \text{mL} = 0.00500\, \text{L}
$$

Calculate the moles of HCl in this aliquot:

$$
\text{Moles of HCl} = C_1 \times V_1 = 0.20\, \text{mol/L} \times 0.00500\, \text{L}

\]

\[
\text{Moles of HCl} = 0.00100\, \text{mol}
\]

This tells us that the 5.00 mL aliquot contains 0.00100 moles of HCl.

Step 3: Applying Conservation of Moles to Find Final Concentration

Since no chemical reactions occur during dilution, the number of moles remains the same. The total moles in the diluted solution are still 0.00100 mol, but now dispersed in a larger volume.

Using the dilution formula:

\[
C_2 = \frac{\text{moles of solute}}{V_2}
\]

Convert the final volume to liters:

\[
V_2 = 25.00\, \text{mL} = 0.02500\, \text{L}
\]

Calculate the final concentration:

\[
C_2 = \frac{0.00100\, \text{mol}}{0.02500\, \text{L}} = 0.0400\, \text{M}

\]

Result: The concentration of HCl after dilution is 0.0400 M.

Summary of the Calculation

Step	Description	Calculation	Result
1	Moles in initial aliquot	$0.20\, \text{mol/L} \times 0.00500\, \text{L}$	0.00100 mol
2	Final concentration after dilution	$\frac{0.00100\, \text{mol}}{0.02500\, \text{L}}$	0.0400 M

Additional Considerations and Practical Applications

Why is this important? In laboratory settings, dilutions are routinely performed to prepare solutions of desired concentrations for titrations, calibration curves, or spectroscopic analyses. Accurate calculations ensure the reliability of experimental results.

Key points to remember:

- The moles of solute remain constant during dilution.
- Volume conversions are essential for accurate calculations.
- Small aliquots can be scaled up to larger volumes while controlling concentration.

Potential Variations and Related Problems

While the problem above asks for the final concentration, similar problems might involve:

- Determining the amount of HCl in the original solution based on titration data.
- Calculating the volume of stock solution needed to prepare a specific concentration.
- Understanding serial dilutions for multi-step concentration adjustments.

Conclusion

The process of diluting a concentrated solution, such as HCl, requires a clear understanding of the relationship between volume, molarity, and moles. Using the fundamental dilution equation $C_1 V_1 = C_2 V_2$, chemists can accurately determine the concentration of solutions post-dilution, ensuring precise experimental conditions. In this example, diluting 5.00 mL of 0.20 M HCl to 25.00 mL yields a final concentration of 0.0400 M, demonstrating how a simple proportional relationship governs the preparation of solutions in the lab.

Final Remarks

Mastering dilution calculations is crucial for any chemist or student working in analytical chemistry. It allows for precise control over solution concentrations, which is essential for reproducibility and accuracy. Whether preparing standard solutions, performing titrations, or analyzing unknowns, understanding the principles outlined here ensures confidence in your experimental results.

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