How Many Milliliters Of A 0.180 M Potassium Chloride Solution Should Be Added To 49.0 ML Of A 0.390 M

How Many Milliliters Of A 0.180 M Potassium Chloride Solution Should Be Added To 49.0 ML Of A 0.390 M is a common question in chemistry, especially when preparing solutions with desired concentrations. Whether you're a student working on a lab experiment or a professional chemist preparing solutions for analytical purposes, understanding how to dilute or mix solutions accurately is essential. This article provides a comprehensive guide to calculating the volume of a 0.180 M potassium chloride (KCl) solution needed to mix with 49.0 mL of a 0.390 M KCl solution, ensuring precise and effective results.

Understanding the Basics of Solution Concentrations

What Is Molarity?

Molarity (M) is a measure of concentration representing the number of moles of solute (in this case, potassium chloride) dissolved in one liter (1000 mL) of solution. It is expressed as:

- **Molarity (M) = Moles of solute / Volume of solution in liters**

For example, a 0.180 M KCl solution contains 0.180 moles of KCl per liter of solution.

Importance of Accurate Concentrations

Precision in solution preparation is vital in chemistry because:

- It ensures reproducibility of experiments.
- It maintains safety standards, especially with reactive substances.
- It guarantees the validity of analytical results.

Calculating the Required Volume of the 0.180 M KCl Solution

The Concept Behind the Calculation

To determine how much of the 0.180 M KCl solution should be added, we need to understand the principle of conservation of moles in solution mixing. When two solutions are mixed, the total amount of solute (moles of KCl) is conserved:

$$
\text{Moles before mixing} = \text{Moles after mixing}
$$

Assuming the final solution is a mixture of the initial solutions, the total moles of KCl are:

$$
n_{\text{total}} = n_1 + n_2
$$

where:

- n_1 = moles of KCl in the 49.0 mL, 0.390 M solution
- n_2 = moles of KCl in the added 0.180 M solution

The total volume after mixing will be:

$$
V_{\text{total}} = V_1 + V_2
$$

where:

- $V_1 = 49.0\, \text{mL}$
- $V_2 =$ volume of 0.180 M solution to be added (unknown)

Our goal is to find V_2 such that the resulting mixture has a desired concentration—often specified, but in this case, it seems to be a question of mixing solutions with given concentrations.

Step-by-Step Calculation

1. Calculate moles of KCl in the initial 49.0 mL of 0.390 M solution:

$$
n_1 = M_1 \times V_1
$$

Convert volume to liters:

$$
V_1 = 49.0\, \text{mL} = 0.0490\, \text{L}
$$

Calculate:

$$
n_1 = 0.390\, \text{mol/L} \times 0.0490\, \text{L} = 0.01911\, \text{mol}
$$

2. Express moles of KCl in the added solution:

$$
n_2 = M_2 \times V_2
$$

where:
- $M_2 = 0.180\, \text{mol/L}$
- V_2 in liters (unknown)

3. Determine the desired final concentration (if specified).

If the problem aims to prepare a solution with a specific final concentration M_f, then:

$$
n_{total} = M_f \times V_{total}
$$

and

$$
\text{Total moles} = n_1 + n_2
$$

and

$$
V_{total} = V_1 + V_2
$$

From these relationships, you can solve for V_2 once the final concentration M_f is specified.

Note: Since the problem as posed doesn't specify a target final concentration, the typical goal is to find V_2 such that the total amount of KCl in the mixture reaches a certain concentration or to understand the proportion of each solution needed for a particular purpose.

Practical Example: Preparing a Solution with a Target Concentration

Suppose you want to prepare a mixture with a final concentration of 0.300 M KCl. Here's how to find the necessary volume of the 0.180 M solution to add:

Step 1: Write the total moles after mixing:

$$
n_{total} = M_f \times V_{total}
$$

Expressed in mol:

$$
n_{total} = 0.300\, \text{mol/L} \times (V_1 + V_2)
$$

\]

Convert V_1 to liters:

$$
V_1 = 0.0490\, \text{L}
$$

Total moles:

$$
n_{total} = 0.300 \times (0.0490 + V_2)
$$

Step 2: Set up the equation with known moles:

$$
n_1 + n_2 = n_{total}
$$

$$
0.01911 + 0.180 \times V_2 = 0.300 \times (0.0490 + V_2)
$$

Express V_2 in liters.

Step 3: Solve for V_2:

$$
0.01911 + 0.180 V_2 = 0.300 \times 0.0490 + 0.300 V_2
$$

Calculate:

$$
0.300 \times 0.0490 = 0.0147
$$

So:

$$
0.01911 + 0.180 V_2 = 0.0147 + 0.300 V_2
$$

Bring like terms together:

$$
0.01911 - 0.0147 = 0.300 V_2 - 0.180 V_2
$$

$$
0.00441 = 0.120 V_2
$$

Solve:

$$
V_2 = \frac{0.00441}{0.120} \approx 0.03675\, \text{L}
$$

Convert to mL:

$$
V_2 \approx 36.75\, \text{mL}
$$

Therefore, approximately 36.75 mL of the 0.180 M KCl solution should be added to 49.0 mL of 0.390 M KCl solution to prepare a 0.300 M KCl solution.

Factors to Consider When Mixing Solutions

Temperature Effects

Chemical reactions and solution concentrations can be temperature-dependent. Always perform calculations at consistent temperatures and adjust if necessary.

Volume Measurement Precision

Use calibrated equipment such as micropipettes, burettes, or volumetric flasks to ensure accuracy.

Solution Homogeneity

Mix solutions thoroughly to ensure uniform distribution of solutes, critical for accurate concentration.

Summary and Best Practices

- Always convert volumes to liters when working with molarity calculations.
- Clearly define the desired final concentration or purpose of the solution.
- Use the molarity formula to relate moles, volume, and concentration.
- For target concentrations, set up an equation based on total moles and solve for the unknown volume.
- Double-check calculations and units to avoid errors.

In conclusion, calculating how many milliliters of a 0.180 M potassium chloride solution to add to a known volume of a 0.390 M solution involves understanding molarity, moles, and volume relationships. With the step-by-step approach outlined above, you can accurately prepare solutions tailored to your experimental needs, ensuring precise and reliable results in your chemistry endeavors.

Frequently Asked Questions

How do I calculate the volume of 0.180 M potassium chloride solution needed to dilute to a specific concentration in 49.0 mL of 0.390 M?

Use the dilution formula: $M_1V_1 = M_2V_2$. Rearranged, $V_2 = (M_1 \times V_1) / M_2$. Plug in the values: $V_2 = (0.390\ M \times 49.0\ mL) / 0.180\ M$, which equals approximately 106.17 mL.

What is the purpose of calculating the volume of potassium chloride solution added in this scenario?

To determine the exact amount of 0.180 M KCl solution needed to achieve a desired final concentration when mixed with a known volume of another solution.

Can I use the dilution formula directly when mixing solutions with different molarities?

Yes, the dilution formula $M_1V_1 = M_2V_2$ applies regardless of the solutions being mixed, as long as the total final volume is correctly determined.

If I want to prepare a specific final concentration, how do I determine the required volume of stock solution?

By rearranging the dilution formula: $V_1 = (M_2 \times V_2) / M_1$, where M_2 and V_2 are the desired final concentration and volume, and M_1 is the initial concentration of the stock solution.

What units should I use for volume when calculating the amount of solution needed?

Use consistent units, such as milliliters (mL). Ensure all volumes are in mL before performing calculations.

Is it necessary to account for the volume added when calculating the final concentration?

Yes, the total volume after mixing influences the final concentration, so you should sum the initial volumes to determine the final concentration accurately.

What is the significance of the molarity values 0.180 M and 0.390 M in this problem?

They represent the initial concentrations of the potassium chloride solutions involved in the dilution process, which are essential for calculating the required volume to achieve the desired concentration.

How precise do I need to be when measuring the volume of potassium chloride solution for accurate results?

Use calibrated equipment like pipettes or burettes for precise measurements, especially in laboratory settings, to ensure accurate and reproducible results.

What are common applications of preparing solutions with specific molarities like in this problem?

Such solutions are used in laboratory experiments, calibrations, chemical reactions, and preparing standards for analytical procedures.

If I add more than the calculated volume of potassium chloride solution, how will it affect the final concentration?

Adding more than the calculated volume will result in a higher final concentration than intended, potentially impacting the accuracy of your experiment or analysis.

Additional Resources

How Many Milliliters Of A 0.180 M Potassium Chloride Solution Should Be Added To 49.0 ML Of A 0.390 M

In the realm of chemistry, accurately preparing solutions with desired concentrations is fundamental to both research and practical applications. Whether in laboratories, industrial processes, or educational settings, understanding how to dilute or combine solutions to achieve specific molarities is a vital skill. Today, we explore a common, yet often nuanced question: How many milliliters of a 0.180 M potassium chloride (KCl) solution should be added to 49.0 mL of a 0.390 M KCl solution to reach a particular target concentration? This problem exemplifies the principles of solution dilution and mixing, highlighting the importance of molarity, volume calculations, and the conservation of moles.

The Context and Significance of Solution Mixing

Before diving into the calculations, it’s essential to understand why such problems matter. In analytical chemistry, preparing solutions with precise concentrations allows scientists to quantify substances accurately. Similarly, in biological experiments, maintaining specific ionic concentrations can influence cellular behavior. The question at hand involves mixing two solutions of potassium chloride with different concentrations and determining the volume of the more dilute solution needed to achieve a desired final concentration.

While the problem statement does not specify the target concentration, it often implies that the goal is to reach a certain molarity or to dilute the initial solution to a specific level. For clarity, we will assume the goal is to prepare a final solution with a known, target molarity (say, a specific concentration). If the problem is to determine how much of the 0.180 M solution should be added to the existing 49.0 mL of 0.390 M KCl to reach a certain final concentration, we can approach this systematically.

Understanding the Principles: Moles, Molarity, and Dilution

Molarity (M) is defined as moles of solute per liter of solution:

$$
\text{Molarity} (M) = \frac{\text{moles of solute}}{\text{liters of solution}}
$$

The key to solving solution mixing problems lies in the conservation of moles. When solutions are mixed (assuming no chemical reactions occur), the total moles of solute before mixing equal the total moles after mixing:

$$
\text{moles}_\text{initial} + \text{moles}_\text{added} = \text{moles}_\text{final}
$$

Expressed mathematically:

$$
C_1 V_1 + C_2 V_2 = C_f V_f
$$

Where:

- C_1, V_1 are the molarity and volume of the initial solution,
- C_2, V_2 are the molarity and volume of the added solution,
- C_f, V_f are the molarity and volume of the final mixture.

Step-by-Step Calculation Approach

1. Identify the known quantities:

- Initial solution:
- Molarity $C_1 = 0.390\, \text{M}$
- Volume $V_1 = 49.0\, \text{mL}$

- Added solution:
- Molarity $C_2 = 0.180\, \text{M}$
- Volume $V_2 = ?\, \text{mL}$

2. Determine the target concentration (if specified):

For this example, let's assume the goal is to prepare a final solution with a molarity of 0.300 M by mixing the two solutions.

3. Express the total moles before and after mixing:

- Moles in initial solution:

$$
n_1 = C_1 \times V_1
$$

Convert V_1 to liters:

$$
V_1 = 49.0\, \text{mL} = 0.049\, \text{L}
$$

$$
n_1 = 0.390\, \text{mol/L} \times 0.049\, \text{L} = 0.01911\, \text{mol}
$$

- Moles in added solution:

$$
n_2 = C_2 \times V_2
$$

Convert V_2 to liters:

$$
V_2\, \text{mL} = V_2 \, \text{mL} \Rightarrow V_2\, \text{L} = \frac{V_2}{1000}

\]

\[
n_2 = 0.180\, \text{mol/L} \times \frac{V_2}{1000}\, \text{L}
\]

4. Express the final concentration:

The total volume after mixing:

\[
V_f = V_1 + V_2
\]

The total moles after mixing:

\[
n_f = n_1 + n_2
\]

The final molarity:

\[
C_f = \frac{n_f}{V_f}
\]

Plugging in the values:

\[
0.300\, \text{M} = \frac{0.01911 + 0.180 \times V_2/1000}{0.049 + V_2/1000}
\]

5. Solve for V_2:

Multiply both sides by the denominator:

$$
0.300 \times (0.049 + V_2/1000) = 0.01911 + 0.180 \times V_2/1000
$$

Distribute:

$$
0.300 \times 0.049 + 0.300 \times \frac{V_2}{1000} = 0.01911 + 0.180 \times \frac{V_2}{1000}
$$

Calculate:

$$
0.0147 + \frac{0.300 V_2}{1000} = 0.01911 + \frac{0.180 V_2}{1000}
$$

Subtract 0.0147 from both sides:

$$
\frac{0.300 V_2}{1000} = 0.00441 + \frac{0.180 V_2}{1000}
$$

Rearrange:

$$
\frac{0.300 V_2}{1000} - \frac{0.180 V_2}{1000} = 0.00441
$$

$$
\frac{(0.300 - 0.180) V_2}{1000} = 0.00441
$$

$$
\frac{0.120 V_2}{1000} = 0.00441
$$

Multiply both sides by 1000:

$$
0.120 V_2 = 4.41
$$

Solve for V_2:

$$
V_2 = \frac{4.41}{0.120} = 36.75\, \text{mL}
$$

Interpreting the Result

Approximately 36.75 mL of the 0.180 M potassium chloride solution should be added to 49.0 mL of the 0.390 M solution to produce a final solution with a molarity of 0.300 M.

This calculation underscores several critical points:

- The importance of unit consistency: converting milliliters to liters is essential for molarity calculations.
- The principle of conservation of moles: the total moles after mixing equal the sum of initial moles.

- The influence of initial concentrations and volumes: higher initial molarity or volume affects how much of the dilute solution is needed.

Broader Applications and Practical Considerations

While the numerical example provides a clear pathway, real-world laboratory scenarios often involve additional factors:

- Precision and measurement errors: Pipettes and volumetric flasks have tolerances that can influence the accuracy of the final concentration.
- Solution stability: Some solutions may degrade over time; hence, calculations should consider freshness and storage.
- Chemical compatibility: The solutions being mixed should not react adversely unless the goal is to produce a new compound.
- Target concentration adjustments: Sometimes, the goal isn't a specific molarity but achieving a particular ionic strength or pH, which may require different calculations.

In industrial or pharmaceutical settings, these calculations become part of quality control protocols, ensuring that products meet exact specifications. In research, they allow scientists to replicate experiments reliably.

Extending the Approach: Other Scenarios

The methodology demonstrated can be adapted to various problems involving solution mixing:

- Dilution of stock solutions: Determining how much of a concentrated stock is needed to prepare a specific volume at a desired concentration.

- Serial dilutions: Calculating successive dilutions to reach very low concentrations.

- Reaction stoichiometry: Balancing react

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