

A Scientist Has A 5.00 Liter Solution Of 4.00 M NaCl. If The Scientist Wanted To Dilute The Solution

A Scientist Has A 5.00 Liter Solution Of 4.00 M NaCl. If The Scientist Wanted To Dilute The Solution, understanding the principles of molarity and dilution is essential for accurate laboratory work. Whether preparing solutions for experiments or industrial applications, the ability to calculate the necessary volume of solvent to achieve a desired concentration is fundamental in chemistry. This article explores the process of dilution, including the underlying chemistry, step-by-step calculations, and practical considerations for scientists working with sodium chloride (NaCl) solutions.

Understanding Molarity and Dilution

What Is Molarity?

Molarity (M) is a measure of concentration that indicates the number of moles of solute (in this case, NaCl) dissolved in one liter of solution. It is expressed as:

- $M = \text{moles of solute} / \text{liters of solution}$

For example, a 4.00 M NaCl solution contains 4.00 moles of NaCl in every liter of solution.

Why Dilute a Solution?

Dilution is performed to reduce the concentration of a solution without changing the total amount of solute present. Reasons include:

- Preparing solutions at specific concentrations for experiments
- Achieving safer or more manageable solution concentrations
- Adjusting concentration to meet protocol requirements

The Principle of Dilution: The Dilution Equation

The key relationship governing dilution is expressed as:

- $C_1V_1 = C_2V_2$

Where:

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

This equation states that the product of initial concentration and volume equals the product of final concentration and volume, assuming no solute is lost or gained during dilution.

Analyzing the Given Scenario

Initial Conditions

The scientist starts with:

- Volume of stock NaCl solution (V_1): 5.00 liters
- Concentration of stock solution (C_1): 4.00 M

Calculating the total amount of NaCl in the stock solution:

- Moles of NaCl = concentration $\times$ volume
- Moles = 4.00 mol/L $\times$ 5.00 L = 20.00 mol

This means the stock solution contains 20.00 moles of NaCl.

Desired Dilution

Suppose the scientist wants to prepare a new solution with a different, lower concentration, C_2 . The goal could be to prepare a specific volume at this new concentration, V_2 .

For example, if the scientist needs:

- Final concentration (C_2): 1.00 M
- Final volume (V_2): To be determined based on experimental needs

The key question is: How much of the stock solution (V_1) should be diluted to reach this concentration?

Calculating the Dilution: Step-by-Step

Step 1: Identify Known Values

- $C_1 = 4.00$ M
- $V_1 = ?$ (to be calculated)
- $C_2 = 1.00$ M (example target concentration)
- $V_2 =$ desired final volume (for example, 2.00 L)

Step 2: Apply the Dilution Equation

Using $C_1V_1 = C_2V_2$:

- $V_1 = (C_2 \times V_2) / C_1$

Plugging in the numbers:

- $V_1 = (1.00 \text{ mol/L} \times 2.00 \text{ L}) / 4.00 \text{ mol/L} = 0.50 \text{ L}$

This means the scientist needs to take 0.50 liters (or 500 milliliters) of the stock solution and dilute it with solvent (water) to reach a final volume

of 2.00 liters.

Step 3: Determine the Volume of Water to Add

- Final volume (V_2): 2.00 L
- Volume of stock solution used (V_1): 0.50 L
- Volume of water to add = $V_2 - V_1 = 2.00 \text{ L} - 0.50 \text{ L} = 1.50 \text{ L}$

Therefore, the scientist should:

- Measure 0.50 liters of the 4.00 M NaCl stock solution
- Add 1.50 liters of distilled water
- Mix thoroughly to produce 2.00 liters of 1.00 M NaCl solution

Practical Considerations in Dilution

Accuracy of Measurements

- Use proper lab equipment such as graduated cylinders, pipettes, or volumetric flasks for precise measurements.
- Ensure thorough mixing to achieve uniform concentration.

Avoiding Common Mistakes

- Always add solute to water, not water to solute, to prevent splashing or incomplete dissolution.
- Use appropriate safety precautions when handling concentrated solutions.

Adjusting for Different Final Concentrations and Volumes

The process remains the same regardless of the target concentration or volume:

- Use the dilution equation to calculate the volume of stock solution needed.
- Adjust the volume of solvent added accordingly.

Additional Examples of Dilution Calculations

Example 1: Preparing a 0.50 M Solution

Suppose the scientist wants 1 liter of 0.50 M NaCl:

- $V_1 = (0.50 \text{ mol/L} \times 1.00 \text{ L}) / 4.00 \text{ mol/L} = 0.125 \text{ L}$
- Measure 125 mL of stock solution
- Add water to reach 1.00 L total volume

Example 2: Creating a Small Volume at a Specific Concentration

For 100 mL of 2.00 M NaCl:

- $V_1 = (2.00 \text{ mol/L} \times 0.10 \text{ L}) / 4.00 \text{ mol/L} = 0.05 \text{ L} = 50 \text{ mL}$
- Take 50 mL of stock solution
- Add 50 mL of water to reach 100 mL total volume

Impact of Dilution on Chemical Solutions

Maintaining Solute Quantity

Regardless of dilution, the total amount of solute (NaCl) remains the same. Dilution simply redistributes the solute in a larger volume.

Applications in Laboratory and Industry

- Preparing buffer solutions
- Standardizing titrations
- Diluting stock solutions to working concentrations

Importance of Precise Calculations

Even small errors in measurement can lead to significant deviations in concentration, potentially affecting experimental outcomes. Therefore, meticulous calculations and measurements are crucial.

Conclusion

Understanding how to dilute a solution accurately is a fundamental skill in chemistry. In the scenario where a scientist starts with a 5.00-liter, 4.00 M NaCl solution and wishes to dilute it to a lower concentration, the process involves applying the dilution equation, measuring the correct volume of stock solution, and adding an appropriate amount of solvent. Proper technique ensures the desired concentration is achieved, enabling reliable and reproducible experimental results. Mastery of these calculations not only enhances laboratory efficiency but also underpins the scientific integrity of chemical research and applications.

Frequently Asked Questions

How much water should the scientist add to dilute the

NaCl solution from 4.00 M to a lower concentration?

The amount of water to add depends on the desired final concentration. Using the dilution formula ($C_1V_1 = C_2V_2$), the scientist can calculate the required volume after dilution and subtract the initial volume to find the amount of water to add.

What is the resulting concentration if the scientist adds enough water to double the volume of the original solution?

Doubling the volume from 5.00 L to 10.00 L will halve the concentration from 4.00 M to 2.00 M, since dilution reduces molarity proportionally to volume increase.

If the scientist wants a final concentration of 1.00 M NaCl, how much water must be added to the original 5.00 L of 4.00 M solution?

Using the dilution formula: $V_2 = (C_1 \times V_1) / C_2 = (4.00 \text{ M} \times 5.00 \text{ L}) / 1.00 \text{ M} = 20.00 \text{ L}$. Therefore, the total volume needed is 20.00 L, meaning 15.00 L of water should be added.

Why is it important to properly dilute NaCl solutions in laboratory experiments?

Proper dilution ensures accurate concentrations for experiments, prevents errors in results, and maintains safety by avoiding overly concentrated solutions that may be hazardous.

What are the steps involved in diluting a 4.00 M NaCl solution to a desired lower concentration?

First, determine the final desired concentration, then use the dilution formula to calculate the final volume. Next, add the appropriate amount of water to the original solution to reach this volume, ensuring thorough mixing.

Can the scientist use the dilution formula for any concentration change in NaCl solutions?

Yes, the dilution formula ($C_1V_1 = C_2V_2$) applies universally for preparing solutions at different concentrations, as long as the solution is dilute and the solute does not react or decompose during dilution.

Additional Resources

A Scientist Has A 5.00 Liter Solution Of 4.00 M NaCl. If The Scientist Wanted To Dilute The Solution

In the world of chemistry and laboratory research, precise control over solution concentrations is essential. Whether preparing reagents for

experiments or scaling up processes, understanding how to accurately dilute solutions is a foundational skill. Imagine a scientist who starts with a 5.00-liter stock solution of sodium chloride (NaCl) at a concentration of 4.00 molarity (M). The question then arises: how can they dilute this solution to achieve a desired lower concentration? This scenario, though seemingly straightforward, involves key concepts of molarity, volume, and the principles of dilution. This article explores the scientific process behind diluting solutions, focusing on the specific case of NaCl, and offers insights into practical calculations and considerations for laboratory work.

Understanding the Basics: Molarity and Solution Concentration

Before diving into the specifics of dilution, it's crucial to understand what molarity (M) signifies in chemistry. Molarity is a measure of concentration, defined as the number of moles of solute (in this case, NaCl) dissolved in one liter of solution.

Key Points About Molarity:

- Definition: Molarity (M) = moles of solute / liters of solution
- Units: Moles (mol), Liters (L)
- Implication: A 4.00 M NaCl solution contains 4.00 moles of NaCl per liter of solution.

Given these parameters, the initial solution contains:

- Total volume: 5.00 L
- Concentration: 4.00 M

Calculating Total Moles of NaCl:

Total moles = Molarity × Volume

= 4.00 mol/L × 5.00 L = 20.00 mol

This total quantity of NaCl remains constant regardless of how we dilute the solution; dilution only involves increasing the volume, which decreases concentration.

The Principle of Dilution: How It Works

Dilution involves adding solvent—water in most cases—to decrease the concentration of a solution without altering the total amount of solute. The fundamental principle governing this process is the dilution equation:

$$C_1V_1 = C_2V_2$$

Where:

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

This equation states that the product of concentration and volume before dilution equals the product after dilution, assuming no solute is lost or added.

Key Assumptions:

- The solute (NaCl) remains evenly distributed.
- No chemical reactions occur during dilution.
- The solution is homogeneous.

Practical Implication: Knowing any three of these variables allows calculation of the fourth, enabling precise control over the final concentration.

Calculating the Dilution: Step-by-Step Approach

Suppose the scientist wants to prepare a less concentrated NaCl solution—say, 1.00 M—from the existing 4.00 M stock. How should they proceed?

Step 1: Identify Known Values

- Initial concentration, $C_1 = 4.00 \text{ M}$
- Final concentration, $C_2 = 1.00 \text{ M}$
- Total amount of NaCl remains the same: 20.00 mol

Step 2: Use the Dilution Equation

Rearranged to find the final volume V_2 :

$$V_2 = (C_1 / C_2) \times V_1$$

But first, determine the volume of the stock solution needed (V_1):

$$V_1 = (\text{Desired amount of NaCl}) / (\text{initial concentration})$$

Since total moles = 20.00 mol,

$$V_1 = 20.00 \text{ mol} / 4.00 \text{ mol/L} = 5.00 \text{ L}$$

Note: The entire stock solution contains 5 liters; thus, to dilute to 1.00 M, the scientist could use part of this stock solution and dilute it further.

Step 3: Decide on the Final Volume

Suppose the scientist wants 10.00 liters of 1.00 M NaCl solution:

$$V_2 = 10.00 \text{ L}$$

Using the dilution equation:

$$V_1 = (C_2 \times V_2) / C_1$$

$$V_1 = (1.00 \text{ mol/L} \times 10.00 \text{ L}) / 4.00 \text{ mol/L} = 2.50 \text{ L}$$

Interpretation: The scientist should take 2.50 liters of the 4.00 M stock solution and dilute it with water until the total volume reaches 10.00 liters.

Practical Considerations in Laboratory Dilution

While calculations provide the theoretical framework, practical execution requires meticulous attention to ensure accuracy and safety.

1. Accurate Measurement of Volumes

- Use calibrated volumetric flasks, graduated cylinders, or pipettes.
- For large volumes, volumetric flasks ensure precision.
- For small volumes, micropipettes provide accuracy.

2. Adding Solvent Gradually

- Add water gradually to the concentrated solution.
- Mix thoroughly after each addition to ensure homogeneity.
- Avoid splashing or aerosol formation, especially with concentrated solutions.

3. Temperature Effects

- Molarity can be temperature-dependent due to volume expansion or contraction.
- Conduct dilution at room temperature or note the temperature for correction.

4. Safety Precautions

- Wear appropriate personal protective equipment (PPE).
- Handle NaCl solutions carefully to avoid ingestion or skin contact.

5. Documentation and Labeling

- Clearly label diluted solutions with concentration, date, and preparation details.
- Maintain records for reproducibility and quality control.

Scaling Up and Down: Practical Examples

Suppose the scientist needs even more dilute solutions for sensitive assays, or perhaps a more concentrated stock for other experiments. Here are scenarios illustrating the flexibility of dilution calculations.

Example 1: Preparing a 0.50 M Solution from the 4.00 M Stock

- Desired concentration (C_2): 0.50 M
- Final volume (V_2): 4.00 L

Calculate V_1 :

$$V_1 = (C_2 \times V_2) / C_1 = (0.50 \text{ mol/L} \times 4.00 \text{ L}) / 4.00 \text{ mol/L} = 0.50 \text{ L}$$

Action: Take 0.50 liters of the 4.00 M NaCl stock and dilute with water to reach 4.00 liters.

Example 2: Concentrating a Dilute Solution

If starting with a 0.10 M solution and aiming to reach 4.00 M, this is not feasible via dilution—it's a concentration process, not dilution. To achieve higher concentrations, techniques like evaporation or crystallization are necessary.

Conclusion: Mastering Dilution for Precise Laboratory Work

Dilution, while straightforward in theory, is a fundamental practice that requires precision and understanding of underlying principles. Whether reducing the concentration of a sodium chloride solution from 4.00 M to a lower value or preparing specific volumes for experiments, the key lies in applying the dilution equation accurately.

In the scenario of a scientist with a 5.00-liter 4.00 M NaCl solution, the process involves calculating the desired final volume and concentration, then carefully measuring and adding water to achieve the target solution. Attention to detail, proper measurement techniques, and safety protocols are essential components of successful dilution.

By mastering these concepts, scientists and laboratory technicians ensure reproducibility, accuracy, and safety in their experiments, ultimately advancing scientific knowledge and discovery.

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