For 490.0 Mlml Of Pure Water, Calculate The Initial Phph And The Final Phph After Adding 1.9102 Molmol

For 490.0 mL of pure water, calculate the initial pH and the final pH after adding 1.9102 mol of solute.

Introduction

Understanding the pH of water and how it changes upon the addition of substances is fundamental in chemistry, especially in acid-base chemistry. Pure water at 25°C is considered neutral, with a pH of 7.0. However, the addition of solutes—particularly acids or bases—can alter this pH. This article provides a comprehensive guide to calculating the initial pH of 490.0 mL of pure water and determines how this pH shifts after the addition of a specific amount of solute, 1.9102 mol, assuming it affects the water's acidity or basicity. We will methodically explore the concepts, calculations, and implications involved in this process.

Understanding pH and Its Calculation

What Is pH?

pH is a measure of the hydrogen ion concentration ($[H^+]$) in a solution. It is defined logarithmically as:

$$pH = -\log [H^+]$$

- A pH of 7.0 indicates neutrality, where $[H^+]$ equals $[OH^-]$.
- pH less than 7.0 indicates acidity.
- pH greater than 7.0 indicates alkalinity.

pH of Pure Water

Pure water at 25°C auto-ionizes slightly:

$2\ H_2O \rightleftharpoons H_3O^+ + OH^-$

The equilibrium concentration of $[H^+]$ and $[OH^-]$ in pure water is:

- $[H^+] = [OH^-] = 1.0 \times 10^{-7}$ M

Therefore, the initial pH of pure water is:

- pH = $-\log(1.0 \times 10^{-7}) = 7.00$

Calculating the Initial pH of 490.0 mL of Pure Water

Step 1: Convert Volume to Liters

Since molarity (mol/L) is used in calculations, convert the given volume:

- Volume (V) = 490.0 mL = 0.490 L

Step 2: Determine the Moles of Water

While water is the solvent, its molarity isn't usually expressed in mols per liter for pure water unless considering its auto-ionization. Instead, focus on the concentration of $[H^+]$ ions, which is 1.0×10^{-7} M at neutral pH.

- The volume of 0.490 L contains:

- Moles of H^+ in pure water = $[H^+] \times V = 1.0 \times 10^{-7}$ mol/L $\times$ 0.490 L = 4.9×10^{-8} mol

However, for pH calculations, this molar amount is implicit; the key point is the concentration.

Step 3: Calculate the Initial pH

Using the known $[H^+]$ in pure water:

- pH = $-\log(1.0 \times 10^{-7}) = 7.00$

Therefore, the initial pH of 490.0 mL of pure water at 25°C is 7.00.

Adding 1.9102 Mol of Substance to Water

Interpreting the Addition

The problem states adding 1.9102 mol, but it does not specify whether this is an acid, base, or neutral substance. The effect on pH depends on the nature of this solute:

- If it’s an acid, $[H^+]$ increases.
- If it’s a base, $[OH^-]$ increases.
- If it’s a neutral salt, the pH may remain unchanged unless it hydrolyzes.

For the purpose of this calculation, we assume the added substance is an acid that fully dissociates and contributes 1.9102 mol of H^+ ions to the solution, a common scenario in acid-base calculations.

Step 1: Calculate the Total Moles of H^+ After Addition

- Moles of H^+ initially: 4.9×10^{-8} mol (negligible compared to added amount)
- Moles of H^+ added: 1.9102 mol

Total moles of H^+ in solution:

- ≈ 1.9102 mol (since initial is negligible)

Step 2: Calculate the Final Concentration of H^+

Total volume after addition:

- The volume of water remains approximately 0.490 L (assuming volume change is negligible for this calculation).

Final $[H^+]$:

- $[H^+]$ = total moles of H^+ / volume (L) = 1.9102 mol / 0.490 L ≈ 3.8986 M

Step 3: Calculate the Final pH

Using the concentration:

- pH = -log $[H^+]$ = -log(3.8986)

Calculating:

- log(3.8986) ≈ 0.591
- pH ≈ -0.591 ≈ -0.59

This indicates a highly acidic solution, which is typical when a strong acid is added in large quantities.

Implications of the Calculation

Understanding the pH Shift

The initial pH of pure water was neutral at 7.00. After adding 1.9102 mol of strong acid, the pH drops dramatically to approximately -0.59. This negative pH value, while unconventional, mathematically indicates an extremely high concentration of hydrogen ions, far exceeding typical aqueous solutions.

In practical terms, such a solution would be highly corrosive and not stable in real-world conditions. Usually, pH values less than 0 are indicative of very concentrated strong acids, such as concentrated sulfuric or hydrochloric acid.

Limitations and Assumptions

- The calculation assumes complete dissociation of the added substance.
- Volume change upon addition is neglected; in reality, adding 1.9102 mol of solute to 0.490 L water would increase the volume, slightly decreasing the concentration.
- The calculation presumes the substance is a strong acid; if it's a base or a weak acid/base, the pH change would differ significantly.
- This example demonstrates the importance of knowing the nature of the added solute for accurate pH prediction.

Summary

- The initial pH of 490.0 mL of pure water at 25°C is 7.00.
- Adding 1.9102 mol of a strong acid (or a substance that contributes H^+ ions) results in a concentration of approximately 3.8986 M H^+.
- The final pH under these assumptions is roughly -0.59, indicating an extremely acidic solution.

Conclusion

Calculating the initial and final pH of water in response to added substances involves understanding the principles of molarity, auto-ionization of water, and acid-base chemistry. While the initial pH of pure water is straightforward at 7.00, the addition of a large amount of acid dramatically shifts the pH into highly acidic territory. These calculations exemplify the importance of knowing the nature of added substances and their dissociation characteristics. In real-world applications, precise pH measurements and considerations of volume changes are essential for accurate assessment and safety.

Further Considerations

- For weak acids or bases, the dissociation constants (Ka or Kb) would need to be considered, complicating calculations.

- Temperature variations affect water ionization and pH.
- Buffer solutions can mitigate pH changes, a factor not included in this simplified calculation.
- Practical laboratory measurements should always be performed to confirm theoretical predictions.

This comprehensive examination illustrates the fundamental principles involved in pH calculation and the profound impact of solute addition on water chemistry.

Frequently Asked Questions

What is the initial pH of 490.0 mL of pure water?

The initial pH of pure water is 7.0, since it is neutral with a pH of 7.

How do you calculate the initial pH of pure water in this problem?

Since the water is pure, its initial pH is 7.0 because the concentration of hydrogen ions (H^+) is 1.0×10^{-7} M.

What is the significance of adding 1.9102 mol of a substance to the water?

Adding 1.9102 mol of an acid or base will alter the hydrogen ion concentration, affecting the pH of the solution.

How do you determine the final pH after adding 1.9102 mol to 490.0 mL of water?

Convert the mL to liters, calculate the new concentration of H^+ ions after addition, and then find the pH using $pH = -\log[H^+]$.

What assumptions are made in calculating the pH after adding the substance?

Assumptions include complete dissociation of the added substance, no volume change upon addition, and the solution being dilute enough for ideal behavior.

If the added substance is an acid, how does that affect the pH?

Adding an acid increases the hydrogen ion concentration, decreasing the pH and making the solution more acidic.

What is the formula to calculate the new hydrogen ion concentration after addition?

New $[H^+]$ = (initial H^+ + moles of acid added) / total volume in liters.

How do you convert mL of water to liters for this calculation?

Divide the volume in milliliters by 1000; so, 490.0 mL = 0.490 L.

What is the final pH if the added mol of substance is a strong acid?

Calculate the new $[H^+]$ concentration by adding the moles of acid to the initial water, then find pH = $-\log[H^+]$.

Can the final pH be exactly 7 after adding 1.9102 mol of acid or base?

No, unless the added amount is exactly balanced by the initial water's H^+ or OH^- ions; generally, the pH will shift away from 7.

Additional Resources

For 490.0 Mlml Of Pure Water, Calculate The Initial Phph And The Final Phph After Adding 1.9102 Molmol

Understanding the pH scale and the influence of added substances on water's acidity or alkalinity is fundamental in chemistry, environmental science, and various industrial processes. In this article, we examine a specific scenario: starting with 490.0 milliliters of pure water, determining its initial pH, and then calculating the pH after introducing 1.9102 mol of a substance—presumably a strong acid or base—into the solution. The goal is to provide a comprehensive, yet accessible, explanation of the calculations involved, illustrating the principles underpinning pH determination and acid-base chemistry.

Introduction to pH and Its Significance

The pH scale, ranging from 0 to 14, serves as a measure of the hydrogen ion concentration ($[H^+]$) in a solution. It is logarithmic, meaning that each unit change in pH corresponds to a tenfold change in $[H^+]$ concentration. For pure water at 25°C, the pH is theoretically 7.0, indicating neutrality—where the concentrations of hydrogen ions (H^+) and hydroxide ions (OH^-) are equal.

Why is pH important?

- Environmental Monitoring: pH affects aquatic life, soil chemistry, and pollution levels.
- Industrial Processes: Many manufacturing processes depend on precise pH control.
- Biological Systems: Human blood, for example, must remain within a narrow pH range (~7.35–7.45).

Given its importance, calculating pH accurately under different conditions is crucial across disciplines.

Initial pH of Pure Water

Understanding the initial state

Pure water at 25°C undergoes auto-ionization:

$$ \mathrm{H_2O} \rightleftharpoons \mathrm{H^+} + \mathrm{OH^-} $$

with an equilibrium constant:

$$ K_w = [\mathrm{H^+}][\mathrm{OH^-}] = 1.0 \times 10^{-14} $$

At neutrality:

$$ [\mathrm{H^+}] = [\mathrm{OH^-}] = \sqrt{K_w} $$

Calculating initial [H^+]:

$$ [\mathrm{H^+}] = \sqrt{1.0 \times 10^{-14}} = 1.0 \times 10^{-7}\ \text{M} $$

And the initial pH:

$$ \mathrm{pH} = -\log [\mathrm{H^+}] = -\log (1.0 \times 10^{-7}) = 7.0 $$

Thus, the initial pH of pure water at 25°C is 7.0, reflecting neutrality.

Note: The pH of pure water can vary slightly with temperature, but for this analysis, we assume standard conditions.

Adding 1.9102 Mol of Substance: What Does It Mean?

The problem states that 1.9102 mol of a substance is added to 490.0 mL of water. To understand the impact, we need to clarify:

- Type of substance: Is it a strong acid (e.g., HCl), a strong base (e.g., NaOH), or a weak acid/base?
- Nature of the substance: The molar amount suggests we are dealing with a strong acid or base, which readily dissociates.

Given the typical context of pH calculations, it's common to assume that the added substance is either:

- A strong acid (e.g., HCl): releasing H^+ ions, lowering pH.
- A strong base (e.g., NaOH): releasing OH^- ions, raising pH.

For this example, let's assume the substance is hydrochloric acid (HCl), a common strong acid, which dissociates completely in water:

$$ \mathrm{HCl} \rightarrow \mathrm{H^+} + \mathrm{Cl^-} $$

This assumption simplifies calculations and aligns with typical chemistry problems.

Note: If the substance were a base, the calculation would follow similar principles but focus on hydroxide concentration.

Calculating the Initial Moles and Concentration of Water

Before adding the substance, the water volume is 490.0 mL:

$$ V = 490.0\ \text{mL} = 0.4900\ \text{L} $$

The number of moles of water:

$$ n_{\mathrm{H_2O}} = \frac{490.0\ \text{g}}{18.015\ \text{g/mol}} \approx 27.2\ \text{mol} $$

However, the number of water molecules or moles isn't directly crucial for pH calculations unless considering auto-ionization. Since pure water is neutral, the focus is on the concentration of added ions.

Concentration of added HCl:

$$ \text{Moles of HCl} = 1.9102\ \text{mol} $$

$$ \text{Concentration of HCl} = \frac{1.9102\ \text{mol}}{0.4900\ \text{L}} \approx 3.898\ \text{M} $$

This high concentration indicates a strongly acidic solution upon addition.

Key Point: For simplicity, the initial pH is 7.0, and by adding a large amount of HCl, we expect the solution's pH to decrease significantly.

Calculating the Final pH After Adding the Substance

Step 1: Determine the total moles of H^+ in the solution

Since HCl dissociates completely:

$$ \text{Moles of H}^+ \text{ added} = 1.9102\ \text{mol} $$

Step 2: Calculate total moles of H^+ in the solution

Initially, water contributes negligible H^+ (at pH 7.0, $[H^+] = 10^{-7}$ M), so the dominant contribution is from the added acid.

Total volume of the solution after addition:

- Assuming volume is additive with no volume contraction:

$$ V_{total} = 490.0\ \text{mL} + \text{volume of acid solution} $$

However, since the acid is added as a solid or concentrated solution, the exact volume contribution is negligible compared to the total volume; for simplicity, we assume the volume remains approximately 0.4900 L.

Step 3: Calculate the $[H^+]$ concentration after addition

$$ [\mathrm{H^+}] = \frac{\text{moles of H}^+}{V} = \frac{1.9102\ \text{mol}}{0.4900\ \text{L}} \approx 3.898\ \text{M} $$

Step 4: Calculate the pH

$$ \mathrm{pH} = -\log [\mathrm{H^+}] = -\log 3.898 $$

Using logarithm properties:

$$ \mathrm{pH} \approx -\log 3.898 \approx - (0.591) \approx -0.591 $$

A pH value less than 0 indicates an extremely acidic solution, which is physically meaningful in the context of concentrated acid solutions.

Summary of results:

- Initial pH: 7.0 (pure water)
- Final pH: approximately -0.59 after adding 1.9102 mol of HCl

This dramatic shift underscores how adding a significant amount of strong acid can drastically alter the acidity of water.

Additional Considerations and Real-World Implications

While the calculations above provide a theoretical framework, real-world scenarios involve complexities:

- Volume changes: Adding concentrated acid may slightly increase the total volume, affecting concentration calculations.
- Dilution effects: If the acid is added as a concentrated solution, initial volume and molarity calculations should account for the actual volume contributed.
- Temperature effects: pH varies with temperature, so precise calculations may require temperature correction.
- Buffering capacity: In natural waters or biological systems, buffering agents can mitigate pH changes, which is not considered here.

Industrial and environmental relevance:

Understanding these calculations is essential when:

- Designing chemical processes that require precise pH control
- Assessing the environmental impact of acidification
- Developing treatments for wastewater or contaminated water sources

Safety note: Handling large quantities of strong acids or bases requires appropriate safety measures.

Conclusion

Starting with 490.0 mL of pure water, the initial pH is 7.0, representing neutrality at standard conditions. Introducing 1.9102 mol of a strong acid like HCl results in an extraordinarily acidic solution with a pH approximately -0.59, highlighting the potent effect of large quantities of acids on water’s pH. These calculations underscore the importance of understanding molarity, ionization, and logarithmic scale properties in chemistry, especially when manipulating solutions for industrial, environmental, or research purposes.

By mastering these fundamental principles, scientists and engineers can predict and control the acidity of solutions, ensuring safety, efficiency, and environmental compliance in their work.

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