A Solution Has [h3o] = 6.4105 M . Use The Ion Product Constant Of Water Kw=[h3o][oh] To Find The [oh]

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Understanding the behavior of water as an aqueous solution is fundamental in chemistry, especially when analyzing pH, pOH, and the concentrations of hydrogen and hydroxide ions. In this article, we explore how to determine the hydroxide ion concentration ($[OH^-]$) in a solution where the hydronium ion concentration ($[H_3O^+]$) is known, specifically using the ion product constant of water, Kw. We will walk through the theoretical background, step-by-step calculations, and practical applications, providing a comprehensive guide for students, educators, and chemistry enthusiasts.

Introduction to Water's Ionization and Kw

Water (H_2O) is a unique substance that exhibits auto-ionization, meaning a small fraction of water molecules naturally dissociate into hydronium (H_3O^+) and hydroxide (OH^-) ions:

$H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + OH^-\ (aq)$

This equilibrium is characterized by the ion product constant of water, Kw, which at 25°C is:

$Kw = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$

This constant is crucial because it connects the concentrations of hydronium and hydroxide ions in any aqueous solution at a given temperature.

Understanding the Relationship Between $[H_3O^+]$, $[OH^-]$, and Kw

In aqueous solutions:

- $[H_3O^+]$ indicates the acidity.
- $[OH^-]$ indicates the alkalinity or basicity.
- The product of these concentrations at 25°C is always equal to Kw.

If you know one of these ion concentrations, you can easily find the other using Kw:

$[OH^-] = Kw / [H_3O^+]$

or

$[H_3O^+] = Kw / [OH^-]$

This relationship allows chemists to quantify the acidity or basicity of a solution accurately.

Calculating $[OH^-]$ from $[H_3O^+]$ in a Solution with $[H_3O^+] = 6.4105$ M

Given:

- $[H_3O^+] = 6.4105$ M
- $Kw = 1.0 \times 10^{-14}$ (at 25°C)

The goal is to find $[OH^-]$.

Step-by-Step Calculation

1. Write the formula:

$[OH^-] = Kw / [H_3O^+]$

2. Substitute known values:

$[OH^-] = (1.0 \times 10^{-14}) / 6.4105$

3. Perform the division:

$[OH^-] \approx 1.560 \times 10^{-15}$ M

Result: The hydroxide ion concentration is approximately 1.560×10^{-15} M.

This extremely low concentration indicates a highly acidic solution, which is consistent with the high $[H_3O^+]$.

Implications of the Calculated $[OH^-]$ Concentration

Understanding the $[OH^-]$ concentration helps in various applications:

- pOH Calculation: $pOH = -\log [OH^-]$
- pH and pOH Relationship: $pH + pOH = 14$ (at 25°C)
- Determining Acidic or Basic Nature: If $[H_3O^+] > 1.0 \times 10^{-7}$ M, the solution is acidic; if $< 1.0 \times 10^{-7}$ M, it is basic.

Applying these:

- $pOH = -\log(1.560 \times 10^{-15}) \approx 14 - \log(1.560) \approx 14 - 0.193 = 13.807$
- $pH = 14 - pOH \approx 0.193$

The pH close to 0 indicates an extremely acidic solution, consistent with the high $[H_3O^+]$.

Key Points to Remember

- The ion product constant of water, Kw, remains constant at a given temperature, allowing for straightforward calculations.
- Knowing $[H_3O^+]$ directly allows for the determination of $[OH^-]$ using the relationship $[OH^-] = Kw / [H_3O^+]$.
- The concentrations of ions influence the solution's pH and pOH, which are vital in chemical analysis and industry applications.

Practical Applications of Calculating $[OH^-]$

Understanding and calculating hydroxide ion concentration has numerous real-world applications:

- Chemical Manufacturing: Ensuring solutions have the correct pH for reactions.
- Environmental Chemistry: Monitoring water quality and pollution levels.
- Biological Systems: Maintaining proper pH for biological processes.
- Pharmaceuticals: Formulating solutions with precise pH levels.
- Food Industry: Ensuring safety and quality in food processing.

Advanced Considerations

While the calculations above are straightforward at 25°C, note that Kw varies with temperature:

- At higher temperatures: Kw increases, making water more ionized.
- At lower temperatures: Kw decreases, reducing ionization.

For precise calculations, always consider temperature effects, and consult temperature-dependent Kw values.

Summary

In this article, we've demonstrated how to determine the hydroxide ion concentration in a solution with a known hydronium ion concentration by utilizing the ion product constant of water, Kw. For a solution with $[H_3O^+]$ = 6.4105 M, the calculation yields an $[OH^-]$ of approximately 1.560×10^{-15} M, revealing a highly acidic environment. Understanding this relationship is essential for various scientific and industrial processes, emphasizing the importance of Kw as a fundamental constant in aqueous chemistry.

Final Thoughts

Mastering the relationship between $[H_3O^+]$, $[OH^-]$, and Kw enables chemists to analyze and manipulate solutions effectively. Whether you’re studying acids, bases, or complex chemical reactions, these calculations provide critical insights into solution behavior, pH balance, and chemical equilibria.

Keywords for SEO:

- Ion product constant of water
- Kw calculation
- Hydroxide ion concentration
- $[H_3O^+]$ and $[OH^-]$ relationship
- pH and pOH calculations
- Acidic solutions
- Basic solutions
- Water ionization
- Aqueous solution chemistry
- Chemistry calculations

Frequently Asked Questions

What is the given concentration of H_3O^+ in the solution?

The concentration of H_3O^+ in the solution is 6.4105 M.

What is the value of the ion product constant of water, Kw?

The ion product constant of water, Kw, is typically 1.0×10^{-14} at 25°C.

How can we find the OH^- concentration using Kw and H_3O^+ concentration?

We can use the relation $K_w = [H_3O^+][OH^-]$, so $[OH^-] = K_w / [H_3O^+]$.

What is the formula to calculate $[OH^-]$ in this problem?

The formula is $[OH^-] = K_w / [H_3O^+]$.

Calculate the hydroxide ion concentration $[OH^-]$ for the given H_3O^+ concentration.

Using $K_w = 1.0 \times 10^{-14}$ and $[H_3O^+] = 6.4105$ M, $[OH^-] = (1.0 \times 10^{-14}) / 6.4105 \approx 1.56 \times 10^{-15}$ M.

Is the solution acidic, basic, or neutral based on the given H_3O^+ concentration?

Since $[H_3O^+]$ is very high (6.4105 M), the solution is strongly acidic.

Why is calculating $[OH^-]$ important in understanding the solution's properties?

Calculating $[OH^-]$ helps determine the pOH and overall acidity or basicity of the solution.

What assumptions are made when using Kw to find $[OH^-]$?

It is assumed that the temperature is 25°C and that the solution is at equilibrium, allowing the use of $K_w = [H_3O^+][OH^-]$.

Additional Resources

A Solution Has $[H_3O^+] = 6.4105$ M: Using the Ion Product Constant of Water (Kw) to Find $[OH^-]$

Understanding the behavior of aqueous solutions is fundamental in chemistry, and one of the core concepts involves the ionization of water and the equilibrium established between hydrogen ions (H_3O^+) and hydroxide ions (OH^-). In this context, the problem of determining the concentration of hydroxide ions ($[OH^-]$) given a specific hydronium ion concentration ($[H_3O^+]$) is a classic example of applying the ion product constant of water, Kw. This article explores the process in detail, demonstrating how to leverage Kw to find $[OH^-]$, analyzing the underlying principles, and discussing the broader implications and applications.

Understanding the Ion Product Constant of Water (Kw)

What is Kw?

The ion product constant of water, denoted as Kw, is a fundamental parameter in aqueous chemistry representing the equilibrium constant for the self-ionization of water:

$$2H_2O(l) \leftrightarrow H_3O^+(aq) + OH^-(aq)$$

At 25°C (standard laboratory temperature), Kw has a value of:

$$Kw = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

This value indicates the extent of water's auto-ionization and serves as a basis for calculating pH, pOH, and the concentrations of ions in various aqueous solutions.

Features of Kw:
- Temperature Dependence: Kw varies with temperature, increasing as temperature rises.
- Constant at a Given Temperature: For calculations at 25°C, Kw is treated as a constant.
- Foundation for pH and pOH Calculations: Knowing Kw allows for the conversion between $[H_3O^+]$ and $[OH^-]$.

Pros & Cons:
- Pros: Provides a straightforward way to relate ion concentrations.
- Cons: Variability with temperature necessitates adjustments for non-standard conditions.

Given Data: $[H_3O^+] = 6.4105$ M

The problem specifies that the concentration of hydronium ions in the solution is 6.4105 molar (M). This is an unusually high concentration for hydronium ions, indicating a strongly acidic solution. Typically, pure water at 25°C has $[H_3O^+] \approx 1.0 \times 10^{-7}$ M; thus, 6.4105 M suggests either an experimental scenario or a hypothetical exercise.

Implications of a High $[H_3O^+]$:
- The solution is strongly acidic.
- The pH can be calculated directly using the $[H_3O^+]$ value.
- The $[OH^-]$ will be correspondingly low, owing to the ion product relationship.

Calculating $[OH^-]$ Using Kw

The Fundamental Equation

The core relationship used to find $[OH^-]$ given $[H_3O^+]$ is:

$$ Kw = [H_3O^+][OH^-] $$

Rearranged to solve for $[OH^-]$:

$$ [OH^-] = \frac{Kw}{[H_3O^+]} $$

Since Kw at 25°C is 1.0×10^{-14}, and $[H_3O^+]$ is 6.4105 M, the calculation becomes straightforward:

$$ [OH^-] = \frac{1.0 \times 10^{-14}}{6.4105} $$

Step-by-Step Calculation

1. Write the known values:
- $Kw = 1.0 \times 10^{-14}$
- $[H_3O^+] = 6.4105$

2. Calculate $[OH^-]$:

$$ [OH^-] = \frac{1.0 \times 10^{-14}}{6.4105} $$

3. Perform the division:

$$ [OH^-] \approx 1.561 \times 10^{-15} \; \text{M} $$

This extremely low hydroxide concentration confirms the highly acidic nature of the solution.

Interpreting the Results

The result, $[OH^-] \approx 1.561 \times 10^{-15}$ M, indicates an environment where hydroxide ions are exceedingly scarce compared to hydronium ions. This aligns with the concept of aqueous neutrality, where the product of the ion concentrations remains constant at a given temperature.

Additional Calculations:

- pH:

$$ pH = -\log [H_3O^+] = -\log 6.4105 \approx -0.807 $$

- pOH:

$$ pOH = -\log [OH^-] \approx -\log 1.561 \times 10^{-15} \approx 14.81 $$

- Verification:

$$ pH + pOH \approx 14 $$ (consistent with the water auto-ionization at 25°C)

Broader Context and Applications

Understanding how to compute $[OH^-]$ from $[H_3O^+]$ using Kw is vital in various chemical contexts.

Applications:
- Acid-Base Titrations: Determining concentrations during titrations.
- pH and pOH Calculations: For solutions with known ion concentrations.
- Environmental Chemistry: Assessing acidity or alkalinity of natural waters.
- Industrial Processes: Controlling pH in manufacturing or chemical reactions.

Features and Benefits:
- Predictive Power: Allows prediction of solution behavior.
- Versatility: Applicable across a range of temperatures with temperature adjustments.

- Educational Value: Fundamental for understanding aqueous equilibria.

Limitations:
- Temperature Sensitivity: Kw varies with temperature, requiring adjustments for non-standard conditions.
- Extremely Concentrated Solutions: At very high or low concentrations, assumptions of ideal behavior may falter.

Conclusion

The process of using the ion product constant of water, Kw, to find the hydroxide ion concentration from a given hydronium ion concentration exemplifies core principles in aqueous chemistry. In the case of $[H_3O^+]$ = 6.4105 M, the calculation reveals an extremely low $[OH^-]$, highlighting the strong acidity of the solution. This exercise underscores the importance of Kw as a fundamental constant that links ion concentrations and facilitates a comprehensive understanding of solution chemistry. Whether in academic settings, research, or industrial applications, mastery of these calculations is essential for accurate analysis and control of aqueous systems.

Key Takeaways:
- The relationship between $[H_3O^+]$ and $[OH^-]$ is governed by Kw.
- High $[H_3O^+]$ corresponds to very low $[OH^-]$, and vice versa.
- Precise calculations are crucial for understanding solution pH and related properties.
- Temperature considerations are vital for accurate assessments of Kw and ion concentrations.

By mastering these concepts, chemists and students can confidently analyze and interpret the behavior of acids, bases, and their solutions, forming a foundational skill in chemical education and practice.

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