Given The [H+] = 3.89x10^-7M, What Is The [OH-]?

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Understanding the relationship between hydrogen ion concentration ([H+]) and hydroxide ion concentration ([OH-]) is fundamental in chemistry, especially in the study of acids, bases, and pH. When given a specific [H+], calculating the corresponding [OH-] involves understanding the concept of ion product of water (Kw). This article explores how to determine [OH-] from [H+], delves into the underlying principles, discusses the significance of pH and pOH, and provides practical examples and tips for solving related problems.

Fundamentals of [H+], [OH-], and Water Ionization

The Autoionization of Water

Water (H_2O) is a neutral molecule that can undergo autoionization, a process where a small fraction of water molecules split into hydrogen ions (H+) and hydroxide ions (OH-):

```plaintext
H₂O (l) ⇌ H+ (aq) + OH- (aq)
```

This equilibrium is essential because it maintains the neutrality of pure water and defines the relationship between [H+] and [OH-].

The Ion Product of Water (Kw)

At 25°C (298 K), the ion product of water, denoted as Kw, is a constant:

```plaintext
Kw = [H+][OH-] = 1.0 × 10^-14
```

This means that in pure water at this temperature, the product of the molar concentrations of hydrogen and hydroxide ions always equals 10^-14.

Calculating [OH-] from [H+]

The Relationship Between [H+] and [OH-]

Given the [H+], you can determine [OH-] using the Kw value:

```plaintext
[OH-] = Kw / [H+]
```

This is derived directly from the equilibrium expression:

```plaintext
[OH-] = (1.0 × 10^-14) / [H+]
```

Applying the Formula to the Given Data

Given:

```plaintext
[H+] = 3.89 × 10^-7 M
```

Calculate [OH-]:

```plaintext
[OH-] = (1.0 × 10^-14) / (3.89 × 10^-7) ≈ 2.57 × 10^-8 M
```

This calculation shows that the hydroxide ion concentration is approximately 2.57 × 10^-8 M when the hydrogen ion concentration is 3.89 × 10^-7 M.

Interpreting the Result

- The [OH-] value is lower than pure water’s typical [OH-], indicating an acidic solution since [H+] > 10^-7 M.
- The relatively small [OH-] reflects the balance maintained in aqueous solutions, where [H+] and [OH-] concentrations are inversely related.

Understanding pH and pOH

Definition of pH and pOH

pH and pOH are logarithmic measures of [H+] and [OH-], respectively:

```plaintext
pH = -log [H+]
pOH = -log [OH-]
```

They serve as more manageable indicators of acidity and alkalinity.

Calculating pH from [H+]

Given [H+] = 3.89 × 10^-7 M:

```plaintext
pH = -log (3.89 × 10^-7) ≈ 6.41
```

This pH value indicates a slightly acidic solution since it's below 7.

Calculating pOH from [OH-]

Using the calculated [OH-] = 2.57 × 10^-8 M:

```plaintext
pOH = -log (2.57 × 10^-8) ≈ 7.59
```

Since pH + pOH ≈ 14 at 25°C, these values are consistent:

```plaintext
6.41 + 7.59 ≈ 14
```

This confirms the calculation’s accuracy and reflects the fundamental relationship between pH and pOH.

Impact of Temperature on [H+], [OH-], and Kw

Kw and Temperature Dependency

While Kw is 1.0 × 10^-14 at 25°C, it varies with temperature:

Temperature	Kw Value
25°C	1.0 × 10^-14
50°C	~5.5 × 10^-14
100°C	~1.0 × 10^-13

An increase in temperature increases Kw, affecting the [H+] and [OH-] concentrations in water.

Implications for pH and pOH

- At higher temperatures, pure water becomes slightly more ionized.
- Therefore, pH of pure water at higher temperatures is slightly below 7.
- When calculating [OH-] at temperatures other than 25°C, use the appropriate Kw value.

Practical Examples and Applications

Example 1: Acidic Solution

Suppose a solution has [H+] = 3.89 × 10^-7 M:

- Determine if the solution is acidic, neutral, or basic.
- Calculate [OH-].

Solution:

- pH ≈ 6.41 (from earlier calculation).
- Since pH < 7, the solution is acidic.
- [OH-] ≈ 2.57 × 10^-8 M.

Example 2: Basic Solution

If [OH-] is known to be 1.0 × 10^-5 M, find [H+].

Solution:

- Use Kw:

```plaintext
[H+] = Kw / [OH-] = 1.0 × 10^-14 / 1.0 × 10^-5 = 1.0 × 10^-9 M
```

- pH = -log(1.0 × 10^-9) = 9.00, indicating a basic solution.

Applications in Real-World Scenarios

- Environmental Chemistry: Monitoring water acidity to ensure safety and ecological balance.
- Industrial Processes: Adjusting pH levels in manufacturing, agriculture, and pharmaceuticals.
- Biological Systems: Maintaining optimal pH for enzyme activity and metabolic processes.

Common Mistakes and Tips for Accurate Calculations

- **Always confirm temperature conditions:** Kw varies with temperature, affecting calculations.
- **Use proper logarithmic functions:** Ensure your calculator is set to the correct mode (usually standard or scientific mode).
- **Check units:** Concentrations are molar (M), and calculations assume molarity unless specified otherwise.
- **Remember the inverse relationship:** As [H+] increases, [OH-] decreases, and vice versa.
- **Be mindful of significant figures:** Maintain appropriate precision based on input data.

Summary

- Given [H+] = 3.89 × 10^-7 M, the corresponding [OH-] is approximately 2.57 × 10^-8 M.
- The relationship is derived from the ion product of water: Kw = [H+][OH-] = 1.0 × 10^-14 at 25°C.
- Calculating pH and pOH provides insights into the acidity or alkalinity of the solution.
- Temperature influences Kw, thus affecting [H+], [OH-], pH, and pOH values.
- These calculations are vital in various scientific fields, including environmental science, medicine, and industrial chemistry.

Conclusion

Understanding how to calculate [OH-] from a given [H+] is an essential skill in chemistry. It enables scientists and students to analyze the acidity or basicity of solutions accurately. By mastering the relationship between [H+], [OH-], pH, and pOH, and recognizing the importance of temperature effects, one can confidently interpret and manipulate aqueous solutions across diverse applications.

Frequently Asked Questions

Given the [H+] = 3.89×10^{-7} M, what is the [OH−] in the solution?

The [OH−] can be calculated using the water dissociation constant Kw = 1.0×10^{-14}. First, find pH: pH = -log[H+] = -log(3.89×10^{-7}) ≈ 6.41. Then, pOH = 14 - pH ≈ 7.59. Finally, [OH−] = 10^(-pOH) ≈ 2.57×10^{-8} M.

How do you determine the [OH−] from the given [H+] in an aqueous solution?

You can determine [OH−] by using the water dissociation constant Kw = [H+][OH−] = 1.0×10^{-14}. Rearranged, [OH−] = Kw / [H+].

What is the relationship between [H+] and [OH−] in water at 25°C?

In pure water at 25°C, the product of [H+] and [OH−] is always 1.0×10^{-14} M^2, i.e., [H+][OH−] = 1.0×10^{-14}.

Calculate the pH of a solution with [H+] = 3.89×10^{-7} M.

pH = -log(3.89×10^{-7}) ≈ 6.41.

If [H+] = 3.89×10^{-7} M, is the solution acidic, neutral, or basic?

Since pH ≈ 6.41 (less than 7), the solution is slightly acidic.

Why is it important to know [OH−] in a solution with a given [H+]?

Knowing [OH−] helps determine the solution's alkalinity or acidity, and is essential for understanding its pH and chemical behavior.

What assumptions are made when calculating [OH−] from [H+] in dilute solutions?

The main assumption is that the solution behaves ideally, and that activity coefficients are approximately 1, which is valid for dilute solutions.

How does temperature affect the calculation of [OH−] from [H+], given the Kw value?

Kw varies with temperature; at higher temperatures, Kw increases, so the relationship between [H+] and [OH−] changes accordingly. Calculations are accurate at 25°C unless adjusted for temperature.

Can you determine the [OH−] directly from [H+] without calculations?

No, direct measurement of [OH−] typically requires titration or specialized sensors; calculations using Kw provide an estimated value based on [H+].

What is the significance of the [OH−] value in biological or chemical processes?

[OH−] influences enzyme activity, solubility, and reaction rates; maintaining proper pH and hydroxide levels is critical in biological systems and industrial processes.

Additional Resources

Given The [H+] = 3.89×10^{-7} M, What Is The [OH-]?

Introduction

In the realm of aqueous chemistry, understanding the concentration of hydrogen ions ([H+]) and hydroxide ions ([OH-]) within a solution is fundamental to deciphering its acidity, alkalinity, and overall chemical behavior. The question, "Given the [H+] = 3.89×10^{-7} M, what is the [OH-]?", exemplifies a core concept: the intrinsic relationship between these ions governed by the water dissociation equilibrium.

This investigation aims to meticulously explore the principles, calculations, and implications behind determining hydroxide ion concentration from a known hydrogen ion concentration, providing a comprehensive review suitable for scientific audiences, educators, and students alike.

Background: The Water Equilibrium and Ion Product Constant

The Autoprotolysis of Water

Pure water undergoes a reversible self-ionization process:

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

This equilibrium is characterized by the ion product constant of water, K_w, which at standard temperature (25°C) is:

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

Significance of K_w

The value of K_w is temperature-dependent but remains constant at a given temperature under standard conditions. It forms the foundation for calculating the pH and pOH of solutions and understanding their acid-base properties.

Determining [OH-] from Given [H+]

The Fundamental Relationship

Given:

$$ [\mathrm{H^+}] = 3.89 \times 10^{-7} \text{ M} $$

and knowing:

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

the hydroxide ion concentration can be derived directly:

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

Calculation Process

Substituting known values:

$$ [\mathrm{OH^-}] = \frac{1.0 \times 10^{-14}}{3.89 \times 10^{-7}} $$

Performing the division:

$$
\begin{aligned}
[\mathrm{OH^-}] &= \frac{1.0 \times 10^{-14}}{3.89 \times 10^{-7}} \\
&= \frac{1.0}{3.89} \times 10^{-14 + 7} \\
&= 0.257 \times 10^{-7} \\

&= 2.57 \times 10^{-8} \text{ M}
\end{aligned}
\]

Final Result

\[
\boxed{
[\mathrm{OH^-}] \approx 2.57 \times 10^{-8} \text{ M}
}
\]

This indicates that in a solution where the hydrogen ion concentration is \(3.89 \times 10^{-7}\) M, the hydroxide ion concentration is approximately \(2.57 \times 10^{-8}\) M.

Interpreting the Results: pH, pOH, and Acid-Base Character

Calculating pH and pOH

The pH of the solution is:

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

\[
pH = -\log (3.89 \times 10^{-7}) \approx -(\log 3.89 + \log 10^{-7}) \approx -(0.590 + (-7)) = 6.41
\]

Similarly, the pOH:

\[pOH = -\log [\mathrm{OH^-}] \]

\[
pOH = -\log (2.57 \times 10^{-8}) \approx -(0.410 + (-8)) = 7.59
\]

Note that:

\[pH + pOH \approx 14 \]

which aligns with the water ion product constant at 25°C.

Acidic, Neutral, or Basic?

Since:

\[pH \approx 6.41 < 7 \]

the solution is mildly acidic, consistent with the higher [H+].

Deeper Analysis: Temperature Dependence and Practical Implications

Temperature Effects on K_w

While the calculations above assume standard temperature (25°C), K_w varies with temperature. For example:

- At 20°C, $K_w \approx 6.8 \times 10^{-15}$
- At 30°C, $K_w \approx 1.2 \times 10^{-14}$

Adjusting calculations accordingly is necessary for precise laboratory or industrial applications.

Practical Applications

Understanding the relationship between [H+] and [OH-] is critical in:

- Pharmaceutical formulations
- Environmental monitoring
- Chemical manufacturing
- Biological systems

Accurate pH control is essential for enzyme activity, solubility, and reaction specificity.

Common Pitfalls and Considerations

Assumption of Pure Water

The calculations assume an ideal, pure water solution without other acids, bases, or ionic interferences.

Ionic Strength Effects

Real-world solutions often contain salts and other ions, which can affect activity coefficients and thus the effective concentrations.

Temperature Control

Accurate measurements require temperature control, as K_w is temperature-dependent.

Broader Context: The Significance of Acid-Base Equilibria

Historical Perspective

The understanding of water dissociation and pH emerged from early 20th-century research, revolutionizing analytical chemistry and biochemistry.

Modern Developments

Advances include:

- Ion-selective electrodes for direct ion measurement
- Spectrophotometric techniques for pH determination
- Computational chemistry models predicting ion activity in complex solutions

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Conclusion

The simple calculation of hydroxide ion concentration from hydrogen ion concentration exemplifies the elegance and utility of fundamental chemical principles. From the initial data point of $[H+] = 3.89\times10^{-7}$ M, deriving [OH-] as approximately 2.57×10^{-8} M underscores the interconnectedness of water's autoionization and the concept of pH.

This analysis not only solidifies core acid-base theory but also highlights the importance of precise measurements and considerations in real-world applications. Whether in lab settings or environmental assessments, understanding these relationships is vital for advancing scientific knowledge and technological innovation.

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