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Calculate Either [H3O+] Or [OH] For Each Of The Solutions At 25°C. Solution A: [OH]=1.89107 M Solution

Understanding the concentrations of hydrogen ions ([H3O+]) and hydroxide ions ([OH−]) in aqueous solutions is fundamental to chemistry, especially when analyzing solution acidity, alkalinity, and pH. At 25°C, the water dissociation constant (Kw) is exactly 1.0×10^{-14}, which allows us to relate [H3O+] and [OH−] concentrations through the simple equation:

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

In this guide, we'll focus on calculating either [H3O+] or [OH−] for solution A, where the hydroxide ion concentration is given as 1.89107 M. We will explore the detailed steps involved, relevant formulas, and implications of these calculations, all while emphasizing the importance of these concepts in chemistry.

Understanding the Basics of Acid-Base Chemistry at 25°C

The Water Dissociation Constant (Kw)

- Defines the relationship between hydronium and hydroxide ions in pure water and aqueous solutions.
- At 25°C, Kw is 1.0×10^{-14}.
- This constant remains the same regardless of the initial concentrations of acids or bases present in the solution.

pH and pOH

- pH measures the acidity of a solution: pH = -log [H3O+].
- pOH measures the alkalinity: pOH = -log [OH−].
- Since pH + pOH = 14 at 25°C, knowing one allows calculation of the other.

Significance of [H3O+] and [OH−]

- [H3O+] indicates the acidity level; higher values mean more acidic.
- [OH−] indicates alkalinity; higher concentrations mean more basic or alkaline solutions.

- Calculating these concentrations helps determine the solution's pH or pOH.

Given Data and Objective

Solution A: $[OH^-]$ = 1.89107 M

- The hydroxide ion concentration is provided.
- Our goal: Calculate either $[H_3O^+]$ or confirm other related parameters like pH or pOH.

Why Focus on Calculations?

- Understanding the relationship between $[H_3O^+]$ and $[OH^-]$.
- Applying the K_w relation at 25°C.
- Gaining insights into the solution's acidity or basicity.

Step-by-Step Calculation of $[H_3O^+]$ from Given $[OH^-]$

Step 1: Recall the Relationship Between $[H_3O^+]$ and $[OH^-]$

- The fundamental equation at 25°C:

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

- This is derived directly from the ion-product constant of water.

Step 2: Substitute the Given $[OH^-]$ into the Equation

- Given:

$$ [OH^-] = 1.89107\, M $$

- Calculation:

$$ [H_3O^+] = \frac{1.0 \times 10^{-14}}{1.89107} $$

Step 3: Perform the Calculation

- Divide:

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{1.89107} \approx 5.28 \times 10^{-15}\, M$$

- This extremely low $[H_3O^+]$ indicates a highly basic solution.

Step 4: Calculate pH and pOH (Optional but Useful)

- pOH:

$$pOH = -\log [OH^-] = -\log(1.89107)$$

$$pOH \approx -0.277$$

- pH:

$$pH = 14 - pOH \approx 14 - (-0.277) = 14.277$$

- Alternatively, directly calculate pH from $[H_3O^+]$:

$$pH = -\log [H_3O^+] \approx -\log(5.28 \times 10^{-15}) \approx 14.28$$

- These values confirm the solution’s strong basic nature.

Implications of the Calculated $[H_3O^+]$ and pH

Understanding the pH of Solution A

- The pH (~14.28) indicates an extremely alkaline solution.
- Typical pH for neutral water at 25°C is 7.
- Values above 7 indicate basic solutions; values approaching 14 are highly basic.

Applications of Such Calculations

- Designing buffer solutions.
- Industrial processes requiring highly basic environments.
- Environmental chemistry, such as analyzing water quality.
- Laboratory titrations involving strong bases.

Limitations and Considerations

- Extremely high $[OH^-]$ concentrations like 1.89107 M are rare; in practice, solutions may have lower concentrations.
- Slight deviations in temperature affect K_w; calculations are valid at 25°C.
- In real-world applications, activity coefficients may slightly modify these calculations, but for most

purposes, the simplified model suffices.

Additional Calculations and Related Concepts

Calculating [OH−] from [H3O+]

- If instead, [H3O+] is given, the calculation for [OH−] is straightforward:

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

Example: Hypothetical [H3O+]

- Suppose [H3O+] = 1.0 × 10^{-7} M (neutral water at 25°C):

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

- Confirming neutral pH of 7.

Effect of Temperature on These Calculations

- As temperature increases, Kw increases, affecting the relationship between [H3O+] and [OH−].
- For precise calculations at temperatures other than 25°C, adjust Kw accordingly.

Summary of Key Formulas

1. $$ [H_3O^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} $$
2. $$ [OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]} $$
3. $$ pH = -\log [H_3O^+] $$
4. $$ pOH = -\log [OH^-] $$
5. $$ pH + pOH = 14 $$ (at 25°C)

Conclusion

In this comprehensive analysis, we've demonstrated how to calculate the hydronium ion concentration [H3O+] from the given hydroxide concentration [OH−] in solution A at 25°C. With [OH−] = 1.89107 M, the calculated [H3O+] is approximately 5.28×10^{-15} M, indicating a strongly basic solution. The pH of this solution exceeds 14, confirming its high alkalinity.

Understanding these calculations is essential for chemists and students alike, as they underpin many practical applications, from laboratory titrations to environmental monitoring. Mastery of the relationship between [H3O+] and [OH−], along with the associated pH and pOH calculations, provides a solid foundation for exploring acid-base chemistry and solution behavior.

Always remember that temperature plays a crucial role in these calculations, and real-world solutions may involve activity coefficients or other factors to consider for more precise analysis. Nonetheless, the fundamental relationship outlined here remains a cornerstone of aqueous solution chemistry.

Meta description: Learn how to calculate [H3O+] from [OH−] in solution at 25°C with step-by-step guidance. Understand the significance of pH, pOH, and water ionization constants in acid-base chemistry.

Frequently Asked Questions

What is the pOH of Solution A with [OH−] = 1.89107 M at 25°C?

pOH = -log[OH−] = -log(1.89107) ≈ -0.277.

How can I calculate the pH of Solution A given its [OH−] at 25°C?

First, find pOH = -log[OH−], then use pH = 14 - pOH. For [OH−] = 1.89107 M, pOH ≈ -0.277, so pH ≈ 14 - (-0.277) = 14.277.

What is the concentration of [H3O+] in Solution A at 25°C?

Calculate [H3O+] using the relation [H3O+] = 10^(-pH). Since pH ≈ 14.277, [H3O+] ≈ 10^(-14.277) ≈ 5.3 × 10^(-15) M.

Is Solution A strongly basic or acidic at 25°C based on its hydroxide concentration?

Since [OH−] = 1.89107 M is very high, Solution A is strongly basic at 25°C.

How does the [OH−] concentration affect the pH of Solution A at 25°C?

A high [OH−] concentration implies a low pOH and a high pH, indicating a strongly basic solution. Specifically, [OH−] of 1.89107 M results in a pH above 14.

Can the [H3O+] concentration in Solution A be greater than 1 M at 25°C?

No, since [OH−] is very high, [H3O+] is extremely low (~5.3 × 10^(-15) M), so it cannot be greater than 1 M.

What is the significance of the high [OH−] concentration in terms of solution pH and pOH?

A high [OH−] (1.89107 M) results in a very low pOH (around -0.277) and a pH greater than 14, indicating a strongly basic solution at 25°C.

Additional Resources

Calculate Either [H_3O^+] or [OH^-] for Each of the Solutions at 25°C: Solution A: [OH^-] = 1.89107 M

When analyzing aqueous solutions, understanding the relationship between hydroxide ions ([OH^-]) and hydronium ions ([H_3O^+]) is fundamental. This calculation, often encountered in chemistry, helps determine the acidity or basicity of a solution, which is essential in fields ranging from environmental science to pharmacology. In this guide, we will explore how to calculate either [H_3O^+] or [OH^-] for Solution A, where the hydroxide ion concentration is given as 1.89107 M at 25°C. We'll break down the concepts, formulas, and step-by-step procedures to provide a comprehensive understanding.

Understanding the Basics: Acid-Base Equilibria and Ion Concentrations

Before diving into calculations, it's crucial to grasp some fundamental concepts:

- pH and pOH: The pH of a solution measures its acidity, whereas pOH measures its alkalinity. They are related through the equation:

pH + pOH = 14 (at 25°C)

- Ion Product of Water (Kw): At 25°C, the ion product of water is:

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

- Concentration Units: All concentrations are expressed in molarity (M), which is moles of solute per liter of solution.

Step 1: Recognize the Known Data and What Is Being Calculated

Given Data:

- $[OH^-]$ = 1.89107 M at 25°C

What we want to find:

- Either $[H_3O^+]$, if we choose to calculate the hydronium ion concentration, or

- The pH or pOH, depending on the context.

Step 2: Applying the Water Ion Product Constant (Kw)

The key relationship at 25°C is:

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

Given $[OH^-]$, we can directly solve for $[H_3O^+]$:

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

Similarly, if $[H_3O^+]$ were known, we could find $[OH^-]$.

Step 3: Calculating $[H_3O^+]$ from $[OH^-]$

Let's perform the calculation for $[H_3O^+]$:

Given:

$$ [OH^-] = 1.89107\, \text{M} $$

Calculate:

$$ [H_3O^+] = \frac{1.0 \times 10^{-14}}{1.89107} $$

$$
[H_3O^+] = \frac{1.0 \times 10^{-14}}{1.89107}
$$

Executing the division:

$$
[H_3O^+] \approx 5.283 \times 10^{-15}\, \text{M}
$$

This extremely low concentration indicates a strongly basic solution.

Step 4: Determine pH and pOH

To fully understand the solution's acidity or alkalinity, calculate pH and pOH:

- pOH:

$$
pOH = -\log [OH^-]
$$

- pH:

$$
pH = 14 - pOH
$$

Calculations:

$$
pOH = -\log(1.89107) \approx -0.276
$$

Since the log of a number greater than 1 is positive, the negative of that is negative, which indicates an extremely high hydroxide concentration:

$$
pOH \approx -0.276
$$

Now, find pH:

$$
pH = 14 - (-0.276) = 14 + 0.276 = 14.276
$$

This pH value is above 14, which is physically unrealistic for pure aqueous solutions, indicating that the $[OH^-]$ concentration is higher than the typical solubility or equilibrium limits in pure water, possibly representing a solution with added hydroxide ions or non-ideal conditions.

Step 5: Summary of Results

Parameter	Calculation	Result
$[H_3O^+]$	$\frac{1.0 \times 10^{-14}}{1.89107}$	≈ 5.283×10^{-15} M
pOH	$-\log 1.89107$	≈ -0.276
pH	14 - pOH	≈ 14.276

Additional Considerations

Is $[OH^-]$ Greater Than 1 M?

In typical aqueous solutions, the concentration of hydroxide ions rarely exceeds 1 M due to solubility and equilibrium constraints. The given $[OH^-]$ value of 1.89107 M suggests that the solution may be strongly basic, possibly due to the presence of a strong base like NaOH at high concentration.

Implications of High $[OH^-]$

- The extremely high hydroxide concentration indicates that the solution is highly alkaline, with a pH exceeding 14.
- Real-world measurements might be limited by measurement accuracy or solubility constraints.

Final Thoughts: How to Use These Calculations

Understanding how to calculate $[H_3O^+]$ from $[OH^-]$, and vice versa, is critical in assessing the acidity or basicity of solutions. The key steps involve:

1. Recognizing the water ion product constant (Kw).
2. Applying the inverse relationship between $[H_3O^+]$ and $[OH^-]$.
3. Calculating pH and pOH to interpret the solution's properties.

By mastering these calculations, chemists and students can analyze a wide variety of solutions, from dilute acids and bases to concentrated industrial solutions.

Summary

In this guide, we've demonstrated how to calculate the hydronium ion concentration $[H_3O^+]$ for Solution A, where the hydroxide ion concentration is given as 1.89107 M at 25°C. Using the water ion product constant Kw, the calculation yields approximately 5.283×10^{-15} M for $[H_3O^+]$, indicating a highly basic solution. Additionally, the pH and pOH calculations reinforce this conclusion, with the pH exceeding 14 due to the high $[OH^-]$ concentration. This process exemplifies fundamental acid-base chemistry principles and highlights the importance of understanding ion relationships in aqueous solutions.

Remember: Always consider the context and limitations of your measurements, especially at extreme concentrations, to ensure accurate interpretation of your results.

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