The Ph Of A 0. 55 M Aqueous Solution Of Hypobromous Acid, Hbro, At 25. 0 C Is 4. 48. What Is The Value

The pH Of A 0.55 M Aqueous Solution Of Hypobromous Acid, HBrO, At 25.0°C Is 4.48. What Is The Value

Understanding the pH of a solution is fundamental in chemistry, especially when studying weak acids like hypobromous acid (HBrO). In this article, we will analyze the problem statement, explore the concepts behind pH and acid dissociation, and provide a comprehensive calculation of the acid dissociation constant (Ka) for hypobromous acid based on the given data. Whether you're a student preparing for exams or a professional chemist, this guide will deepen your understanding of acid-base chemistry and the methods used to determine equilibrium constants.

Introduction to pH and Acid-Base Chemistry

The pH scale measures the acidity or alkalinity of an aqueous solution, ranging from 0 (most acidic) to 14 (most basic), with 7 being neutral. It is logarithmically related to the concentration of hydrogen ions (H^+) in the solution:

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

For weak acids like hypobromous acid, which do not dissociate completely, the pH depends on the acid's strength (Ka), concentration, and degree of ionization.

Understanding Hypobromous Acid (HBrO)

Hypobromous acid is a weak, monobasic acid, similar in behavior to hypochlorous acid (HOCl). Its chemical formula is HBrO, and it partially dissociates in water:

Dissociation Equation:

$HBrO\ (aq) \rightleftharpoons H^+\ (aq) + BrO^-\ (aq)$

The extent of this dissociation is characterized by the acid dissociation constant (Ka):

Expression for Ka:

$Ka = [H^+][BrO^-] / [HBrO]$

Given Data and Objective

The problem provides the following data:

- Concentration of hypobromous acid, $[HBrO] = 0.55\ M$
- pH of the solution = 4.48
- Temperature = 25.0°C

Our goal is to determine the value of the acid dissociation constant, Ka, for hypobromous acid under these conditions.

Step-by-Step Calculation of Ka

To find Ka, we follow these steps:

1. Calculate the concentration of H^+ ions from the pH.
2. Determine the equilibrium concentrations of dissociated species.

3. Apply the Ka expression to find its value.

Step 1: Calculate $[H^+]$

Using the pH value:

$[H^+] = 10^{(-pH)} = 10^{(-4.48)}$

Calculating:

$[H^+] \approx 3.31 \times 10^{(-5)}$ M

This concentration also equals $[BrO^-]$ at equilibrium, since each dissociation produces one H^+ and one BrO^- ion.

Step 2: Determine Equilibrium Concentrations

Assuming x is the amount of HBrO that dissociates:

- Initial concentration of HBrO: 0.55 M
- Change due to dissociation: -x
- At equilibrium: $[HBrO] = 0.55 - x$
- $[H^+] = x \approx 3.31 \times 10^{(-5)}$ M
- $[BrO^-] = x \approx 3.31 \times 10^{(-5)}$ M

Given that x is small compared to initial concentration, the approximation $[HBrO] \approx 0.55$ M is acceptable for the calculation.

Step 3: Calculate Ka

Using the expression:

$$Ka = [H^+][BrO^-] / [HBrO]$$

Substitute known values:

$$Ka \approx (3.31 \times 10^{-5}) \times (3.31 \times 10^{-5}) / 0.55$$

Calculate numerator:

$$(3.31 \times 10^{-5})^2 \approx 1.10 \times 10^{-9}$$

Now, divide by 0.55:

$$Ka \approx 1.10 \times 10^{-9} / 0.55 \approx 2.00 \times 10^{-9}$$

Final Result and Interpretation

The calculated acid dissociation constant for hypobromous acid at 25°C is approximately:

$Ka \approx 2.00 \times 10^{-9}$

This value indicates that hypobromous acid is a very weak acid, with only a small fraction dissociating in aqueous solution. The low Ka reflects its limited ability to donate protons under standard conditions.

Additional Considerations

Influence of Temperature

While this calculation assumes standard temperature (25°C), Ka can vary with temperature. Generally, an increase in temperature can either increase or decrease Ka depending on whether the dissociation process is endothermic or exothermic.

pH Measurement and Accuracy

Accurate pH measurement is crucial. Small errors in pH readings can significantly impact the calculated Ka, especially for weak acids where dissociation is minimal.

Limitations of Approximation

In cases where x is not negligible compared to the initial concentration, a quadratic equation should be solved for more precise results.

Summary of Key Points

- The pH of a solution provides direct insight into the hydrogen ion concentration.
- Hypobromous acid (HBrO) is a weak acid with limited dissociation.
- Calculating Ka from pH involves understanding equilibrium concentrations.
- Assumptions made for small x values simplify calculations but should be verified.

Practical Applications of Ka and pH Calculations

Understanding the dissociation constant and pH of hypobromous acid is vital in various fields, including:

- Environmental chemistry: hypobromous acid's role in water disinfection

- Industrial processes: manufacturing and stability of bromine compounds
- Biochemistry: understanding reactive halogen species in biological systems

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Conclusion

Determining the Ka of hypobromous acid based on pH and concentration measurements offers valuable insights into its acid strength and behavior in aqueous solutions. The calculated Ka value of approximately 2.00 × 10^(-9) confirms its status as a weak acid, consistent with similar oxyacids like hypochlorous acid. Mastery of these calculations enables chemists to predict solution behavior, design experiments, and understand reaction equilibria more comprehensively.

References and Further Reading

- Atkins, P., & de Paula, J. (2010). Physical Chemistry (9th ed.). Oxford University Press.
- Brown, T. L., LeMay, H. E., Bursten, B. E., & Murphy, C. (2014). Chemistry: The Central Science (13th ed.). Pearson.
- House, J. E. (2007). Inorganic Chemistry. Academic Press.

For more detailed explanations and practice problems on pH, Ka, and acid-base equilibria, consult reputable chemistry textbooks or online educational resources.

Frequently Asked Questions

What is the pH of a 0.55 M aqueous solution of hypobromous acid (HBrO)

at 25°C if the pH is 4.48?

The pH is given as 4.48, so the pH of the solution is 4.48.

How can the pH of hypobromous acid solution be determined from its concentration?

The pH can be determined using the acid's dissociation constant (Ka) and the concentration, applying an equilibrium expression to find the hydrogen ion concentration, then calculating pH as -log[H+].

What is the significance of the pH value 4.48 for a 0.55 M HBrO solution?

A pH of 4.48 indicates that the solution is mildly acidic, with some degree of dissociation of hypobromous acid in water.

How do you calculate the concentration of H+ ions in the solution from the pH value?

The concentration of H+ ions is calculated as [H+] = 10^(-pH). For pH 4.48, [H+] ≈ 3.31 x 10^(-5) M.

What is the importance of knowing the pH of hypobromous acid solutions in practical applications?

Knowing the pH helps in controlling chemical reactions, disinfectant effectiveness, and environmental impact assessments involving hypobromous acid.

Can the pH of 4.48 for the hypobromous acid solution be used to estimate its dissociation constant (Ka)?

Yes, using the concentration and pH, along with an equilibrium expression, one can estimate the Ka of hypobromous acid, which is typically around 2 x 10^(-9).

Additional Resources

The pH of a 0.55 M aqueous solution of hypobromous acid, HBrO, at 25.0°C is 4.48. What is the value? Understanding and calculating the pH of such a solution involves a combination of concepts from acid-base chemistry, equilibrium calculations, and logarithmic functions. Whether you're a student tackling chemistry homework or a professional analyzing chemical reactions, this guide will walk you through the detailed process of determining the pH of hypobromous acid solutions, emphasizing the steps needed to arrive at the correct value.

Understanding Hypobromous Acid (HBrO)

Hypobromous acid (HBrO) is a weak, monobasic oxyacid of bromine. It is part of the halogen oxyacid family, which includes acids like hypochlorous acid (HOCl) and hypoiodous acid (HOI). These acids are characterized by their partial ionization in water, which results in a solution that is only mildly acidic.

Key Properties of HBrO:
- Weak acid: Does not fully dissociate in water.
- Chemical formula: HBrO.
- pKa value: Typically around 8.6 (though this varies slightly depending on sources).
- Relevance: Used in water treatment and disinfection processes.

Understanding the acid's strength (or lack thereof) is crucial for calculating the pH of its solutions.

The Significance of pH and Its Calculation

pH is a measure of the acidity or alkalinity of an aqueous solution, defined as:

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

In weak acid solutions, the concentration of hydrogen ions $[\text{H}^+]$ depends on the degree of ionization of the acid, which is governed by its acid dissociation constant K_a.

Given:
- Concentration of HBrO: 0.55 M
- pH of the solution: 4.48

Our goal is to verify or compute the K_a value based on this information, or vice versa, find the hydrogen ion concentration.

Step-by-Step Approach to the Calculation

1. Determine the $[\text{H}^+]$ from the pH

Using the pH value:

$$

$$
[\text{H}^+] = 10^{-\text{pH}} = 10^{-4.48}
$$

Calculating:

$$
[\text{H}^+] \approx 3.31 \times 10^{-5} \text{ M}
$$

This is the concentration of hydrogen ions in the solution.

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2. Set Up the Acid Dissociation Equilibrium

The dissociation of hypobromous acid in water:

$$
\text{HBrO} \leftrightarrow \text{H}^+ + \text{BrO}^-
$$

The expression for the acid dissociation constant K_a:

$$
K_a = \frac{[\text{H}^+][\text{BrO}^-]}{[\text{HBrO}]}
$$

Assuming initial concentration is 0.55 M, and that $x = [\text{H}^+] = [\text{BrO}^-]$:

$$
\begin{cases}
[\text{HBrO}] \approx 0.55 - x \\
[\text{H}^+] = x \\
[\text{BrO}^-] = x
\end{cases}
$$

Since x is small relative to initial concentration, the approximation:

$$
[\text{HBrO}] \approx 0.55
$$

is often valid, but because the pH indicates a non-negligible dissociation, a more accurate calculation considers the exact values.

3. Calculate K_a Using the Known $[\text{H}^+]$

$$
K_a = \frac{x \times x}{0.55 - x} \approx \frac{(3.31 \times 10^{-5})^2}{0.55 - 3.31 \times 10^{-5}}
$$

Calculating numerator:

$$
(3.31 \times 10^{-5})^2 \approx 1.10 \times 10^{-9}
$$

Calculating denominator (approximated as 0.55):

$$
0.55 - 3.31 \times 10^{-5} \approx 0.54997 \approx 0.55
$$

Thus:

$$
K_a \approx \frac{1.10 \times 10^{-9}}{0.55} \approx 2.00 \times 10^{-9}
$$

4. Interpretation of the K_a value

This K_a indicates a weak acid, consistent with the pH and the initial concentration.

Additional Considerations and Calculations

1. Validating the Approximation

Given the pH, the dissociation is enough to produce a measurable $[\text{H}^+]$, but not so high as to invalidate the approximation that $[\text{HBrO}]$ remains close to its initial concentration.

2. Calculating the Degree of Ionization

$$
\

$$
\alpha = \frac{[\text{H}^+]}{\text{initial concentration}} = \frac{3.31 \times 10^{-5}}{0.55} \approx 6.02 \times 10^{-5}
$$

This tiny ionization fraction confirms the weak acid nature of HBrO.

Summary of Key Results

Parameter	Calculation / Value
$[\text{H}^+]$	3.31×10^{-5} M (from pH 4.48)
K_a	$\sim 2.00 \times 10^{-9}$ (calculated)
Degree of ionization (α)	~0.006%

Practical Applications and Implications

Understanding the pH of hypobromous acid solutions is critical in various fields:

- Water treatment: HBrO is used as a disinfectant, and knowing its pH helps optimize dosage.
- Chemical synthesis: Precise pH control ensures desired reactions.
- Environmental chemistry: Assessing the behavior of bromine compounds in natural waters.

Conclusion

The pH of a 0.55 M aqueous solution of hypobromous acid at 25°C being 4.48 reveals important insights into its acid strength and dissociation behavior. By converting pH to hydrogen ion concentration, setting up the equilibrium expression, and calculating the acid dissociation constant, we find that hypobromous acid is a weak acid with a K_a on the order of 2.00×10^{-9}. This comprehensive analysis underscores the importance of understanding acid-base principles for practical applications and laboratory analysis.

Final Note

Accurate pH calculations often require careful approximation and validation. Always consider the approximations' validity based on the acid's strength and initial concentration. With practice, these methods become intuitive tools for chemists and students alike.

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