What Is The Hypochlorite Ion Concentration In A 0.69 M Solution Of Hocl? See The Acid-base Table For

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Understanding the concentration of hypochlorite ions in a solution is crucial for various applications, especially in disinfection, water treatment, and chemistry laboratories. When dealing with solutions such as hypochlorous acid (HOCl), it's essential to comprehend how it dissociates and interacts with the surrounding environment. In this article, we will explore how to determine the hypochlorite ion (ClO^-) concentration in a 0.69 M solution of HOCl, using the principles outlined in the acid-base table, and delve into related concepts to provide a comprehensive understanding.

Introduction to Hypochlorous Acid and Hypochlorite Ions

Hypochlorous acid (HOCl) is a weak acid commonly used as a disinfectant and sanitizer. It exists in equilibrium with its conjugate base, the hypochlorite ion (ClO^-), especially in aqueous solutions. The dissociation of HOCl in water can be represented as:

$$HOCl \rightleftharpoons H^+ + ClO^-$$

Understanding this equilibrium is fundamental to calculating the concentration of each species in solution at equilibrium. The pKa of HOCl is approximately 7.53, indicating its acid strength and how it dissociates depending on the pH of the environment.

Fundamentals of Acid-Base Equilibria and the Acid-Base Table

The acid-base table, often called the "equilibrium table" or "ICE table" (Initial, Change, Equilibrium), helps in calculating concentrations of various species in a chemical equilibrium. For weak acids like HOCl, knowing the initial concentration and the acid dissociation constant (Ka) allows us to determine the equilibrium concentrations of H^+ and ClO^-.

Key concepts include:

- Initial concentration: The starting molarity of the acid before dissociation.
- Change: The amount of acid that dissociates at equilibrium.
- Equilibrium concentration: The concentration of each species after dissociation reaches equilibrium.

The dissociation constant, Ka, for HOCl is related to pKa:

$$Ka = 10^{(-pKa)} \approx 10^{(-7.53)} \approx 2.95 \times 10^{(-8)}$$

This small Ka indicates that HOCl is a weak acid, dissociating only partially in water.

Calculating Hypochlorite Ion Concentration in a 0.69 M HOCl Solution

Let's explore step-by-step how to determine the hypochlorite ion concentration in a 0.69 M solution of HOCl.

Step 1: Write the Dissociation Expression

The dissociation of HOCl:

$$HOCl \rightleftharpoons H^+ + ClO^-$$

Initial concentrations:

Species	Initial (M)	Change (M)	Equilibrium (M)
HOCl	0.69	-x	0.69 - x
H^+	0	+x	x
ClO^-	0	+x	x

Step 2: Write the Expression for Ka

$$Ka = [H^+][ClO^-] / [HOCl] = x \ \ x / (0.69 - x) = 2.95 \times 10^{(-8)}$$

Since Ka is very small, we can approximate (0.69 - x) ≈ 0.69, because x will be small compared to initial concentration.

Step 3: Solve for x (H^+ and ClO^- concentrations)

Using the approximation:

$x^2 / 0.69 \approx 2.95 \times 10^{-8}$

Then,

$x^2 \approx 0.69 \times 2.95 \times 10^{-8} \approx 2.036 \times 10^{-8}$

Taking the square root:

$x \approx \sqrt{(2.036 \times 10^{-8})} \approx 4.51 \times 10^{-5}$ M

Therefore,

- The concentration of H^+ ions is approximately 4.51×10^{-5} M.
- The concentration of ClO^- ions at equilibrium is approximately 4.51×10^{-5} M.

Interpreting the Results

The hypochlorite ion concentration in this solution is approximately 4.51×10^{-5} M, which is a tiny fraction compared to the initial HOCl concentration of 0.69 M. This low concentration underscores the weak dissociation nature of HOCl.

Summary:

- Initial HOCl concentration: 0.69 M
- Equilibrium ClO^- concentration: approximately 4.51×10^{-5} M
- pH of the solution can be estimated from $[H^+]$, which is around 4.52×10^{-5} M, corresponding to a pH of about 4.34.

Factors Affecting Hypochlorite Ion Concentration

Several factors influence the equilibrium and resulting hypochlorite ion concentration:

1. pH of the Solution

- The dissociation of HOCl is pH-dependent.
- At lower pH, the equilibrium favors undissociated HOCl.
- At higher pH, more HOCl dissociates into ClO^- and H^+.

2. Temperature

- Increasing temperature can shift the equilibrium, affecting ion concentrations.

3. Presence of Other Ions

- Competing equilibria or complex formation can alter free ion concentrations.

Applications of Hypochlorite Ion Concentration Knowledge

Understanding hypochlorite ion concentrations is vital across multiple domains:

1. Water Treatment and Disinfection

- Ensuring appropriate ClO^- levels for effective disinfection.
- Maintaining safe residual disinfectant levels.

2. Chemical Manufacturing

- Producing sodium hypochlorite solutions with precise concentrations.

3. Laboratory Analytical Chemistry

- Designing experiments involving oxidation-reduction reactions.

Conclusion

Determining the hypochlorite ion concentration in a solution of hypochlorous acid requires understanding acid-base equilibria, dissociation constants, and applying the ICE table method. In a 0.69 M HOCl solution, the hypochlorite concentration is approximately 4.51×10^{-5} M, illustrating the weak nature of HOCl dissociation. Recognizing how pH and other factors influence this equilibrium is essential for optimizing applications in disinfection, water treatment, and chemical synthesis. By leveraging the acid-base table and equilibrium principles, chemists and engineers can accurately predict ion concentrations and ensure the effectiveness and safety of their processes.

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Keywords: hypochlorite ion concentration, HOCl, acid-base equilibrium, dissociation constant, ICE table, weak acid dissociation, water treatment, disinfection, chemical equilibrium

Frequently Asked Questions

What is the hypochlorite ion concentration in a 0.69 M solution of HOCl?

The hypochlorite ion concentration depends on the degree of dissociation of HOCl in water, which can be determined using the acid-base equilibrium and the acid-base table. Typically, it will be less than 0.69 M, depending on the pH and dissociation constant.

How does the acid-base table help determine the hypochlorite ion concentration in HOCl solutions?

The acid-base table provides information on the dissociation constants (Ka and Kb) and the relative concentrations of species at equilibrium, allowing calculation of the hypochlorite ion (ClO^-) concentration in the solution.

What is the relevance of the pH in determining the hypochlorite ion concentration in HOCl solutions?

pH influences the extent of dissociation of HOCl; at higher pH, more HOCl dissociates into ClO^- ions, increasing hypochlorite concentration, which can be analyzed using the acid-base equilibrium data.

Why is understanding the hypochlorite ion concentration important in chemistry?

Knowing the hypochlorite ion concentration helps in applications like water disinfection, understanding bleach strength, and managing chemical reactions involving chlorine species.

How do you calculate the hypochlorite ion concentration from the initial HOCl molarity?

By applying the acid dissociation constant (Ka) for HOCl and setting up an equilibrium expression, you can solve for the concentration of ClO^- in the solution based on initial molarity and pH.

What role does the acid-base table play in solving for hypochlorite ion concentration?

It provides the necessary equilibrium constants and species distribution data, enabling calculation of the hypochlorite ion concentration at equilibrium in the solution.

Can the hypochlorite ion concentration be directly read from the acid-base table?

No, the table provides the constants and relative species amounts; the actual concentration must be calculated using equilibrium expressions and initial conditions.

What is the typical pH range where hypochlorite ions are predominant in HOCl solutions?

Hypochlorite ions are predominant in alkaline conditions, typically at pH values above 7, especially around pH 9-11 where dissociation favors ClO^- formation.

How does temperature affect the hypochlorite ion concentration in a 0.69 M HOCl solution?

Temperature affects the equilibrium constants; increasing temperature can shift the dissociation equilibrium, thus altering the hypochlorite ion concentration accordingly.

Additional Resources

What Is The Hypochlorite Ion Concentration In A 0.69 M Solution Of HOCl? See The Acid-base Table For

Understanding the chemistry behind common disinfectants and their ionization processes is essential for chemists, healthcare professionals, and environmental scientists alike. Among these compounds, hypochlorous acid (HOCl) is widely used for its potent antimicrobial properties, particularly in water treatment, sanitation, and wound care. When dissolved in water, HOCl undergoes dissociation to produce hypochlorite ions (OCl^-), which are responsible for its disinfectant activity. But what exactly is the concentration of hypochlorite ions in a solution initially containing 0.69 molarity of HOCl? To answer this, we delve into the chemistry of HOCl, analyze the relevant acid-base equilibria, and interpret the data using the acid-base table.

Understanding Hypochlorous Acid (HOCl) and Its Dissociation

Hypochlorous acid (HOCl) is a weak acid with notable disinfectant properties due to its ability to

penetrate microbial cell walls and deactivate vital enzymes. Its chemical structure comprises a chlorine atom bonded to a hydroxyl group, making it capable of donating protons in aqueous solutions.

Dissociation Reaction:

When dissolved in water, HOCl undergoes an equilibrium dissociation:

$HOCl\ (aq) \rightleftharpoons H^+\ (aq) + OCl^-\ (aq)$

This equilibrium is characterized by the acid dissociation constant, Ka, which indicates the strength of the acid.

Key Parameters:

- Initial concentration of HOCl: 0.69 M
- Ka of HOCl: approximately $3.0 \times 10^-$ VIII (varies slightly depending on temperature and source)
- pKa of HOCl: approximately 7.5

The pKa is especially useful as it indicates the pH at which half of the HOCl is dissociated into hypochlorite ions.

The Acid-Base Table and Its Role in Determining Ion Concentrations

The acid-base table, or equilibrium table, provides a structured way to analyze how much of the initial HOCl dissociates under given conditions. It systematically accounts for the initial concentrations, changes due to dissociation, and equilibrium concentrations.

Why Use the Acid-Base Table?

- To determine the equilibrium concentrations of all species involved (HOCl, H^+, OCl^-).
- To understand how the dissociation shifts based on pH and initial molarity.
- To quantify the hypochlorite ion concentration resulting from the dissociation of 0.69 M HOCl.

Fundamentals of the Table:

Species	Initial Concentration (M)	Change (M)	Equilibrium Concentration (M)
HOCl	0.69	$-x$	$0.69 - x$
H^+	0	$+x$	x
OCl^-	0	$+x$	x

In this framework, 'x' represents the amount of HOCl that dissociates into H^+ and OCl^- at equilibrium.

Applying the Equilibrium Data: Calculating the OCl− Concentration

To find the hypochlorite ion concentration, we need to link the dissociation extent 'x' with known constants and initial conditions.

Step 1: Write the Expression for Ka

$Ka = [H^+][OCl^-] / [HOCl]$

Substituting the equilibrium concentrations:

$Ka = (x)(x) / (0.69 - x) = x^2 / (0.69 - x)$

Step 2: Use the pKa to Find Ka

Given pKa ≈ 7.5,

$Ka = 10^{(-pKa)} \approx 10^{(-7.5)} \approx 3.16 \times 10^{-8}$

Step 3: Set Up the Equation

$x^2 / (0.69 - x) = 3.16 \times 10^{-8}$

Since Ka is very small, dissociation is limited, and x will be much less than 0.69 M. Therefore, we can approximate:

$0.69 - x \approx 0.69$

This simplifies the equation to:

$x^2 / 0.69 \approx 3.16 \times 10^{-8}$

Step 4: Solve for x

$x^2 \approx 3.16 \times 10^{-8} \times 0.69 \approx 2.18 \times 10^{-8}$

$x \approx \sqrt{(2.18 \times 10^{-8})} \approx 4.67 \times 10^{-5}$ M

Thus, the equilibrium concentration of hypochlorite ions (OCl−) is approximately 4.67×10^{-5} M.

Interpreting the Results: What Does the Concentration Mean?

The calculated hypochlorite ion concentration of approximately 4.67×10^{-5} M indicates that, in a 0.69 M solution of HOCl, only a tiny fraction (about 0.0068%) dissociates into OCl^- under typical conditions.

Implications:

- Antimicrobial Activity: Despite the low concentration of hypochlorite ions, HOCl remains an effective disinfectant because of its high reactivity and ability to penetrate microbial cells.
- pH Dependence: The dissociation degree is highly sensitive to pH. At pH levels below the pKa (~7.5), HOCl predominates; above pKa, hypochlorite ions dominate.
- Application Context: In practical scenarios, adjusting pH levels can optimize the proportion of hypochlorite ions, enhancing disinfection efficacy.

Factors Influencing Hypochlorite Ion Concentration

The simple calculation provides a baseline, but several factors can influence the actual hypochlorite ion concentration in real-world solutions:

1. pH of the Solution

- Lower pH (~<7): More HOCl, less OCl^-.
- Higher pH (~>8): More OCl^-, increasing hypochlorite ion concentration.

2. Temperature

- Elevated temperatures can shift the dissociation equilibrium, slightly increasing or decreasing ion concentrations.

3. Presence of Other Ions

- Competing ions or buffering agents can alter the pH and dissociation equilibrium.

4. Initial Concentration

- Higher initial concentrations of HOCl can lead to different dissociation extents due to activity coefficients and ionic strength.

Practical Applications and Significance

Understanding the hypochlorite ion concentration in a given solution is vital for optimizing disinfection protocols, designing water treatment systems, and ensuring safety standards.

Water Treatment:

- Chlorine-based disinfectants like HOCl and OCl− are used to eliminate pathogens.
- Maintaining appropriate pH levels ensures maximum hypochlorite ion presence for effective microbial kill.

Healthcare and Sterilization:

- HOCl solutions are favored for wound care due to their rapid action.
- Precise knowledge of ion concentrations informs safe and effective concentration levels.

Environmental Considerations:

- The stability and reactivity of hypochlorite ions influence the formation of disinfection byproducts.
- Accurate calculations help in minimizing environmental impact.

Conclusion

The concentration of hypochlorite ions in a 0.69 M solution of HOCl can be precisely estimated through equilibrium analysis using the acid-base table and the known dissociation constant. The calculation reveals that only a small fraction of HOCl dissociates into OCl− under typical conditions, emphasizing the importance of pH control and environmental factors in practical applications. Whether in disinfection, water treatment, or medical settings, understanding this fundamental chemistry ensures effective and safe use of hypochlorite-based compounds.

By applying the principles of acid-base equilibrium, scientists and practitioners can better predict solution behaviors, optimize disinfectant performance, and safeguard public health and environmental quality.

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