Calculate [h3o+] Of The Following Polyprotic Acid Solution: 0.120 M H2co3.

Calculate [H_3O^+] Of The Following Polyprotic Acid Solution: 0.120 M H_2CO_3

Understanding how to determine the concentration of hydronium ions ([H_3O^+]) in a solution of a polyprotic acid like carbonic acid (H_2CO_3) is fundamental in acid-base chemistry. Polyprotic acids are acids that can donate more than one proton (H^+) per molecule during dissociation. Carbonic acid, in particular, is a common example because it plays a vital role in biological systems and environmental chemistry, especially in buffering systems of blood and natural waters.

This article provides an in-depth guide on how to calculate [H_3O^+] in a 0.120 M H_2CO_3 solution, emphasizing the principles of acid dissociation, equilibrium expressions, and the step-by-step process to arrive at the accurate concentration of hydronium ions.

Understanding Polyprotic Acids and Their Dissociation

What Are Polyprotic Acids?

Polyprotic acids are acids capable of donating multiple protons through successive dissociation steps. Each dissociation step has its own equilibrium constant (Ka), typically decreasing with each subsequent proton donation. For example, a generic polyprotic acid (H_nA) dissociates as follows:

1. First dissociation:
$H_nA \rightleftharpoons H^+ + H_{n-1}A^-$

2. Second dissociation:
$H_{n-1}A^- \rightleftharpoons H^+ + H_{n-2}A^{2-}$

and so on, until all protons are donated.

Carbonic Acid as a Polyprotic Acid

Carbonic acid (H_2CO_3) is a diprotic acid with two dissociation steps:

1. First dissociation:
$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$ ($Ka_1 \approx 4.3 \times 10^{-7}$)

2. Second dissociation:
$HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$ ($Ka_2 \approx 5.1 \times 10^{-11}$)

These dissociation constants indicate that the first proton dissociation is more significant than the second, which is comparatively weak.

The Importance of Calculating [H_3O^+]

Accurately calculating [H_3O^+] in solutions of polyprotic acids like H_2CO_3 is crucial for understanding pH, which influences biological functions, environmental chemistry, and industrial processes. Since pH is calculated as:

pH = -log [H_3O^+]

knowing [H_3O^+] allows for precise pH determination, which is essential in various scientific and practical applications.

Step-by-Step Approach to Calculate [H_3O^+] in 0.120 M H_2CO_3 Solution

The calculation involves understanding the dissociation equilibria of H_2CO_3 and applying equilibrium principles to determine the hydronium ion concentration.

Step 1: Recognize the Dominant Dissociation Step

Since Ka_1 ($\approx 4.3 \times 10^{-7}$) is much larger than Ka_2 ($\approx 5.1 \times 10^{-11}$), the majority of the hydronium ions originate from the first dissociation. Therefore, initial calculations can focus on the first dissociation, treating the second as a minor contribution, and then refining as necessary.

Step 2: Write the Equilibrium Expression for First Dissociation

For the first dissociation:

$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$

The equilibrium expression:

$Ka_1 = [H^+][HCO_3^-] / [H_2CO_3]$

Assuming initial concentration of H_2CO_3 is 0.120 M, and that x is the concentration of H_3O^+ produced:

- [H_3O^+] = x
- [HCO_3^-] ≈ x

- $[H_2CO_3]$ at equilibrium ≈ 0.120 - x

Since Ka_1 is small, x is expected to be small relative to initial concentration, allowing the approximation:

$[H_2CO_3] \approx 0.120$ M

Step 3: Set Up the Expression and Calculate x

Using the approximation:

$Ka_1 \approx x^2 / 0.120$

Solve for x:

$x^2 = Ka_1 \times 0.120$

$x = \sqrt{(Ka_1 \times 0.120)}$

Plugging in the values:

$x = \sqrt{(4.3 \times 10^{-7} \times 0.120)}$

$x \approx \sqrt{(5.16 \times 10^{-8})}$

$x \approx 7.19 \times 10^{-5}$ M

This value of x represents the concentration of $[H_3O^+]$ from the first dissociation.

Step 4: Consider the Contribution of the Second Dissociation

The second dissociation:

$HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$

has a much smaller Ka_2, indicating it contributes minimally to the overall $[H_3O^+]$. However, for more precise calculations, especially at higher accuracy, the second dissociation can be included via an equilibrium model.

Given its negligible contribution, for most practical purposes, the $[H_3O^+]$ is approximately 7.19×10^{-5} M.

Step 5: Calculate the pH

Using the $[H_3O^+]$ value:

$pH = -\log [H_3O^+]$

$pH \approx -\log (7.19 \times 10^{-5})$

$pH \approx 4.14$

Refining the Calculation: Including the Second Dissociation

For a more accurate calculation, especially if high precision is needed, the contribution of the second dissociation should be included. The second dissociation produces additional [H_3O^+], but since Ka_2 is very small, its effect is minimal at this concentration.

If you wish to incorporate it, you'd set up a system of equilibrium equations considering both dissociation steps, but typically, the first dissociation dominates the hydronium ion concentration in such solutions.

Practical Applications of Calculating [H_3O^+]

Understanding and calculating [H_3O^+] in polyprotic acid solutions has numerous applications:

- **Biology:** Maintaining blood pH within a narrow range is vital for physiological functions.
- **Environmental Chemistry:** Acidification of natural waters depends on carbonate equilibria.
- **Industrial Processes:** Buffer solutions and chemical manufacturing often rely on precise pH control.
- **Analytical Chemistry:** Titration of polyprotic acids requires knowledge of the dissociation equilibria.

Summary of Key Points

- Polyprotic acids like H_2CO_3 dissociate in multiple steps, each characterized by its own Ka.
- For dilute solutions, the first dissociation predominates in determining [H_3O^+].
- The approximation [H_3O^+] ≈ √(Ka_1 × initial concentration) is often sufficient for initial calculations.

- Refined calculations can include second dissociation steps for higher accuracy.
- The resulting [H_3O^+] allows calculation of pH, providing insights into the solution's acidity.

Conclusion

Calculating the [H_3O^+] in a 0.120 M H_2CO_3 solution demonstrates essential principles of acid-base chemistry and equilibrium analysis. Understanding the dominant dissociation steps and applying the appropriate equilibrium expressions enables accurate pH determination, which is crucial across biological, environmental, and industrial contexts. By following systematic approaches and considering the strengths of each dissociation step, chemists can effectively analyze complex polyprotic acid systems and their behaviors in various solutions.

Further Reading and Resources

- Chemistry Textbooks:
- "Principles of General Chemistry" by Peter Atkins
- "Quantitative Chemical Analysis" by Daniel C. Harris

- Online Resources:
- Khan Academy's Acid-Base Equilibrium Tutorials
- LibreTexts Chemistry Modules on Polyprotic Acids

- Tools:
- Online pH calculators
- Equilibrium solvers and simulation software

By mastering the calculation of [H_3O^+] in polyprotic acid solutions, students and professionals can deepen their understanding of complex chemical equilibria and improve their analytical skills in real-world scenarios.

Frequently Asked Questions

How do you determine the concentration of H_3O^+ in a 0.120 M H_2CO_3 solution?

You need to consider the dissociation steps of carbonic acid, calculate the equilibrium concentrations using Ka values, and solve for [H_3O^+] through equilibrium expressions.

What is the first dissociation step of carbonic acid, and how does it affect pH?

The first dissociation is $H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$, which primarily contributes to initial $[H_3O^+]$, influencing the solution's pH. The degree of dissociation depends on Ka_1 and the initial concentration.

Why do we need to consider multiple dissociation steps when calculating $[H_3O^+]$ for H_2CO_3?

Because H_2CO_3 is a polyprotic acid, it dissociates in multiple steps, and each step contributes to the overall $[H_3O^+]$. Accurate calculation requires accounting for all relevant dissociation equilibria.

What is the approximate pH of a 0.120 M H_2CO_3 solution?

The pH is approximately 4.3 to 4.5, but precise calculation involves solving equilibrium expressions using Ka values; the initial $[H_3O^+]$ corresponds to $[H^+]$ from the first dissociation step.

How do the Ka values for H_2CO_3 influence the calculation of $[H_3O^+]$?

Ka values determine the extent of dissociation at each step. For H_2CO_3, Ka_1 is about 4.3×10^{-7} and Ka_2 is about 5.6×10^{-11}, which affect the calculation of $[H_3O^+]$ by defining the equilibrium concentrations.

Can you directly assume complete dissociation of H_2CO_3 in dilute solutions? Why or why not?

No, because H_2CO_3 is a weak acid with partial dissociation. Assuming complete dissociation would overestimate $[H_3O^+]$ and lead to inaccurate pH calculations; equilibrium considerations are necessary.

Additional Resources

Calculating $[H_3O^+]$ of a 0.120 M H_2CO_3 Solution: An In-Depth Expert Analysis

---

Introduction to Polyprotic Acids and Their Significance in Chemistry

Polyprotic acids are a fascinating class of acids characterized by their ability to donate multiple protons (H^+) during their dissociation in aqueous solutions. Unlike monoprotic acids such as hydrochloric acid (HCl), which release only a single proton, polyprotic acids—like carbonic acid

(H_2CO_3), sulfuric acid (H_2SO_4), and phosphoric acid (H_3PO_4)—undergo multiple ionization steps, each with unique equilibrium constants.

Understanding how to calculate the concentration of hydronium ions ($[H_3O^+]$) for solutions of polyprotic acids is crucial in fields ranging from environmental chemistry to biochemistry and industrial processes. Accurate calculations aid in predicting pH, designing buffer solutions, and understanding reaction mechanisms.

This article provides an expert-level, detailed walkthrough of calculating the hydronium ion concentration for a 0.120 M H_2CO_3 (carbonic acid) solution, emphasizing the underlying principles, step-by-step procedures, and key considerations.

Fundamentals of Carbonic Acid Dissociation

Structure and Dissociation Pathways

Carbonic acid (H_2CO_3) is a diprotic acid, meaning it can donate two protons through two successive dissociation steps:

1. First dissociation:
$$
\mathrm{H_2CO_3} \leftrightarrow \mathrm{H^+} + \mathrm{HCO_3^-}
$$

2. Second dissociation:
$$
\mathrm{HCO_3^-} \leftrightarrow \mathrm{H^+} + \mathrm{CO_3^{2-}}
$$

Each step has its own equilibrium constant:

- K_{a1} for the first dissociation,
- K_{a2} for the second dissociation.

Typical values at 25°C:
- $K_{a1} \approx 4.3 \times 10^{-7}$
- $K_{a2} \approx 5.6 \times 10^{-11}$

These values indicate that the first dissociation is much more significant than the second, which is a key consideration in calculations.

Implications for [H_3O^+] Calculation

Because the first dissociation is dominant, the initial approximation is that the majority of H_2CO_3 dissociates into H^+ and HCO_3^-, with the second dissociation contributing minimally. Nonetheless, for precise pH calculations, especially at higher concentrations, both steps must be considered.

Methodologies for Calculating [H_3O^+]

Different approaches exist for calculating hydronium ion concentration in polyprotic acid solutions:

- Assumption-based approximation, where the dominant dissociation step is considered alone.
- Iterative or simultaneous solution of multiple equilibrium expressions, considering all dissociation steps.
- Using the Henderson-Hasselbalch equation for buffer solutions, if applicable.

Given the complexity and the goal of an in-depth, expert-level explanation, this article will focus on the comprehensive method involving equilibrium expressions and quadratic solutions.

Step-by-Step Calculation of [H_3O^+] for 0.120 M H_2CO_3

Step 1: Establish Initial Conditions and Assumptions

- Initial concentration of H_2CO_3: $\mathrm{[H_2CO_3]_0} = 0.120\, M$
- The major dissociation occurs in the first step; the second is less significant but still considered for accuracy.
- The initial concentration of HCO_3^- and CO_3^{2-} is zero.

Step 2: Write Equilibrium Expressions

First dissociation:

$$
\mathrm{H_2CO_3} \leftrightarrow \mathrm{H^+} + \mathrm{HCO_3^-}
$$

- Let x = concentration of H_3O^+ (or H^+) produced by the first dissociation.

Equilibrium concentrations:

- $\mathrm{[H_2CO_3]} \approx 0.120 - x$
- $\mathrm{[H^+]} \approx x$
- $\mathrm{[HCO_3^-]} \approx x$

Equilibrium expression:

$$
K_{a1} = \frac{[\mathrm{H^+}][\mathrm{HCO_3^-}]}{[\mathrm{H_2CO_3}]} \approx \frac{x \times x}{0.120 - x} = \frac{x^2}{0.120 - x}
$$

Given the small magnitude of K_{a1}, and for initial approximation, assume $x \ll 0.120$, thus:

$$
K_{a1} \approx \frac{x^2}{0.120}
$$

which simplifies to:

$$
x^2 \approx K_{a1} \times 0.120
$$

Substituting the known value:

$$
x^2 \approx 4.3 \times 10^{-7} \times 0.120 = 5.16 \times 10^{-8}
$$

$$
x \approx \sqrt{5.16 \times 10^{-8}} \approx 7.19 \times 10^{-5}\, M
$$

This initial approximation indicates the $[H_3O^+]$ is around 7.19×10^{-5} M, corresponding to a pH of approximately 4.14.

Step 3: Check the Validity of the Approximation

Since x is much less than 0.120 M, the assumption $0.120 - x \approx 0.120$ is valid. However, to improve accuracy, especially considering the second dissociation, a more detailed calculation involving quadratic equations is necessary.

Step 4: Incorporate the Second Dissociation Step

Second dissociation:

$$
\mathrm{HCO_3^-} \leftrightarrow \mathrm{H^+} + \mathrm{CO_3^{2-}}
$$

- Let y = concentration of $\mathrm{CO_3^{2-}}$ formed.

Equilibrium expression:

$$
K_{a2} = \frac{[\mathrm{H^+}][\mathrm{CO_3^{2-}}]}{[\mathrm{HCO_3^-}]} \approx \frac{(x + y) \times y}{x - y}
$$

Since K_{a2} is very small (5.6×10^{-11}) and x is already small, the second dissociation contributes negligibly to $[H_3O^+]$, but for precise calculations, the equilibrium must be solved simultaneously.

Simplification:

Assuming $y \ll x$ and $y \ll 0.120$, the second dissociation's contribution to $[H_3O^+]$ is minimal, and the dominant contribution remains from the first dissociation.

Step 5: Final Calculation of $[H_3O^+]$

Given the dominance of the first dissociation, the approximate $[H_3O^+]$ is:

$$
[\mathrm{H_3O^+}] \approx x \approx 7.19 \times 10^{-5}\, M
$$

pH Calculation:

$$
\mathrm{pH} = -\log [\mathrm{H_3O^+}] \approx -\log(7.19 \times 10^{-5}) \approx 4.14
$$

Refined Calculations and Considerations

While the above provides a good approximation, more refined calculations involve solving quadratic equations explicitly, considering the second dissociation, and possibly iterative numerical methods.

Quadratic solution for the first dissociation:

$$
x^2 + K_{a1} x - K_{a1} \times 0.120 = 0
$$

which yields:

$$
x = \frac{-K_{a1} + \sqrt{K_{a1}^2 + 4 \times K_{a1} \times 0.120}}{2}
$$

Plugging in the numbers:

$$
x = \frac{-4.3 \times 10^{-7} + \sqrt{(4.3 \times 10^{-7})^2 + 4 \times 4.3 \times 10^{-7} \times 0.120}}{2}
$$

Because K_{a1} is very small, the quadratic simplifies further, confirming the previous approximation.

Key insight: The dominant contributor to [H_3O^+] in this solution is from

Calculate H3o Of The Following Polyprotic Acid Solution 0 120 M H2co3

Calculate H3o Of The Following Polyprotic Acid Solution 0 120 M H2co3

Related Articles

- The Segments Shown Below Could Form A Triangle.CB911BCO A. TrueB. False
- An Example Of The Postponement Strategy For Improving Service Productivity
- Complete The Square To Re-write The Quadratic Function In Vertex Form

Back to Home