What Is The H(aq) Concentration In 0.05 M HCN(aq)? (K, For HCN Is 5.0 X 10-10) 5.0x10-10 M 5.0*10-4M

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Understanding the concentration of hydrogen ions (H^+ or H_3O^+) in a solution of weak acid like hydrocyanic acid (HCN) is fundamental in chemistry, particularly in acid-base equilibria. Given a solution of 0.05 M HCN and the acid dissociation constant (K_a) of 5.0×10^{-10}, determining the H^+ concentration involves applying principles of equilibrium chemistry. The question also presents options—5.0×10^{-10} M and 5.0×10^{-4} M—prompting an analysis to identify the correct value based on the chemical properties and the mathematical approach involving the acid dissociation constant.

Understanding Acid Dissociation Constant (K_a)

What Is K_a?

- The acid dissociation constant, denoted as K_a, measures the strength of an acid in solution.
- It is defined as the equilibrium constant for the dissociation of a weak acid (HA) into H^+ and A^-:

$HA \rightleftharpoons H^+ + A^-$

- Mathematically, $K_a = [H^+][A^-] / [HA]$

The Significance of K_a Values

- A small K_a (much less than 1) indicates a weak acid, which dissociates minimally.
- Conversely, a large K_a (close to 1 or greater) suggests a strong acid, which dissociates extensively.

In our case, HCN has a K_a of 5.0×10^{-10}, illustrating its weakness as an acid.

Analyzing the Problem: HCN Dissociation in Water

Given Data

- Initial concentration of HCN: $[HCN]_0 = 0.05$ M
- K_a of HCN: 5.0×10^{-10}
- Unknown: $[H^+]$ (or $[H_3O^+]$) at equilibrium

Assumptions for Weak Acid Calculations

- The initial concentration of HCN is significantly larger than the concentration of H^+ produced.
- The degree of dissociation is small, allowing us to use the approximation that $[HCN] \approx [HCN]_0$.

Setting Up the Equilibrium Expression

Let $x = [H^+]$ at equilibrium. Because HCN dissociates into H^+ and CN^-:

- $[H^+] = x$
- $[CN^-] = x$
- [HCN] at equilibrium = $[HCN]_0 - x \approx [HCN]_0$ (since x is small)

The equilibrium expression:

$$K_a = x^2 / ([HCN]_0 - x) \approx x^2 / [HCN]_0$$

Rearranged:

$$x^2 = K_a \times [HCN]_0$$

Calculating the H^+ Concentration

Applying the Approximation

Using the approximation that $x \ll [HCN]_0$, the calculation simplifies:

$$x = \sqrt{(K_a \times [HCN]_0)}$$

Substitute the known values:

$$x = \sqrt{(5.0 \times 10^{-10} \times 0.05)}$$

Calculate the product:

$$5.0 \times 10^{-10} \times 0.05 = 2.5 \times 10^{-11}$$

Now, take the square root:

$x = \sqrt{(2.5 \times 10^{-11})} \approx \sqrt{2.5} \times 10^{-11}$

$\sqrt{2.5} \approx 1.58$

Therefore:

$x \approx 1.58 \times 10^{-6}$ M

Interpreting the Result

- The equilibrium concentration of H^+ ions is approximately 1.58×10^{-6} M.
- This value is consistent with the expectation for a weak acid with such a small K_a.

Comparison With Provided Options

- 5.0×10^{-10} M — This is roughly the K_a value itself, not the H^+ concentration resulting from dissociation.

- 5.0×10^{-4} M — This is significantly larger than our calculated value and would suggest a much stronger acid or higher dissociation.

Given the calculation, the H^+ concentration in 0.05 M HCN with a K_a of 5.0×10^{-10} is approximately 1.58×10^{-6} M, which is much closer to 10^{-6} than to either of the options provided.

Understanding the Discrepancy in Options

Why Are the Options Confusing?

- The options listed may be typographical or conceptual misinterpretations.
- 5.0×10^{-10} M corresponds to the K_a, representing the maximum possible H^+ concentration if the acid dissociated completely, which it does not.
- 5.0×10^{-4} M suggests a much higher degree of dissociation than expected for such a weak acid.

Conclusion on Correct H^+ Concentration

- The accurate H^+ concentration in 0.05 M HCN with a K_a of 5.0×10^{-10} is approximately 1.6×10^{-6} M.
- Neither of the options provided exactly matches this value, but based on typical question structures,

the closest conceptual answer is that the actual H^+ concentration is very low, on the order of 10^{-6} M.

The Importance of Accurate Calculations in Acid-Base Chemistry

Why Precision Matters

- Small differences in K_a lead to large variations in pH calculations.
- Correct assumptions (e.g., negligible dissociation) are critical for reliable results.

Practical Applications

- Acid strength assessments.
- Buffer solution preparations.
- pH calculations in environmental and biological systems.

Summary and Final Remarks

- The dissociation of HCN in water is minimal due to its very small K_a.
- The H^+ concentration in a 0.05 M HCN solution is approximately 1.6×10^{-6} M, confirming its status as a weak acid.
- The options given (5.0×10^{-10} M and 5.0×10^{-4} M) do not precisely match the calculated value, but understanding the underlying principles clarifies that the actual concentration is closer to the former, very low value.

In conclusion, accurate application of equilibrium principles reveals that the hydrogen ion concentration in a 0.05 M HCN solution with a K_a of 5.0×10^{-10} is approximately 1.6×10^{-6} M, emphasizing the weak acidic nature of HCN and the importance of precise calculations in chemical equilibria.

Frequently Asked Questions

What is the H^+ concentration in a 0.05 M HCN solution given its K_a value?

The H^+ concentration in 0.05 M HCN is approximately 5.0×10^{-4} M.

How do you calculate the H^+ concentration in a weak acid like HCN?

Use the acid dissociation constant expression, setting up an equilibrium table to solve for $[H^+]$, considering the initial concentration and K_a.

Given that K_a for HCN is 5.0×10^{-10}, what is the approximate H^+ concentration in 0.05 M HCN?

The H^+ concentration is approximately 5.0×10^{-4} M.

Why is the H^+ concentration in HCN much lower than the initial concentration?

Because HCN is a weak acid with a small K_a value, it only partially dissociates, resulting in a low H^+ concentration.

Can the H^+ concentration in 0.05 M HCN be approximated directly from K_a?

Yes, since K_a is small, the H^+ concentration can be approximated as the square root of (K_a × initial concentration), which yields about 5.0×10^{-4} M.

What is the significance of the K_a value being 5.0×10^{-10} for HCN?

It indicates that HCN is a very weak acid, dissociating very minimally in aqueous solution.

Is the H^+ concentration in 0.05 M HCN affected significantly by the initial concentration?

Yes, for weak acids, the initial concentration influences the equilibrium H^+ concentration, but due to low dissociation, the change is minimal.

How do you verify the H^+ concentration calculation for HCN solution?

By applying the approximation $[H^+] \approx \sqrt{(K_a \times \text{initial concentration})}$ and confirming that the assumption of negligible ionization of the initial concentration holds true.

What is the correct H(aq) concentration in 0.05 M HCN given $K_a = 5.0 \times 10^{-10}$?

The H^+ concentration in the solution is approximately 5.0×10^{-4} M.

Additional Resources

Understanding the H(aq) Concentration in 0.05 M HCN(aq): A Comprehensive Analysis

When exploring the behavior of weak acids like hydrogen cyanide (HCN) in aqueous solutions, a fundamental question arises: What is the concentration of hydrogen ions (H^+) in a 0.05 M HCN solution? This question not only touches upon acid dissociation principles but also involves understanding equilibrium chemistry, calculating equilibrium concentrations, and interpreting the significance of the acid dissociation constant (K_a). In this detailed review, we will dissect this problem thoroughly, incorporating theoretical background, step-by-step calculations, and practical implications.

Fundamentals of Acid Dissociation and the Role of K_a

What Is HCN and Its Acidic Nature?

Hydrogen cyanide (HCN) is a weak, volatile acid that exists in aqueous solution as a molecular compound capable of donating protons (H^+). Its chemical behavior aligns with typical weak acids, characterized by partial ionization in water. The dissociation reaction can be represented as:

$$ \text{HCN} \ \rightleftharpoons \ H^+ + \text{CN}^- $$

Given the weak nature of HCN, it does not dissociate completely; instead, an equilibrium is established with a relatively small concentration of free H^+ ions.

The Acid Dissociation Constant (K_a)

The strength of a weak acid is quantified by its acid dissociation constant, K_a, which is defined as:

$$ K_a = \frac{[\text{H}^+][\text{CN}^-]}{[\text{HCN}]} $$

For HCN, the given K_a is:

$$ K_a = 5.0 \times 10^{-10} $$

This very small value indicates that only a tiny fraction of HCN molecules dissociate in water, resulting in a low concentration of H^+ ions.

Setting Up the Problem: Initial Concentration and Equilibrium Conditions

Initial Concentration of HCN

Before dissociation, the initial concentration of HCN in solution is:

$$ [\text{HCN}]_{initial} = 0.05 \ \text{M} $$

Since HCN is a weak acid, it's reasonable to assume that the initial concentrations of H^+ and CN^- are zero.

Change in Concentration During Dissociation

Let x be the concentration of H^+ (and CN^-) produced at equilibrium. Because one mole of HCN produces one mole of H^+ and CN^-, the equilibrium concentrations are:

- $[\text{H}^+] = x$
- $[\text{CN}^-] = x$
- $[\text{HCN}] = 0.05 - x$

Given that K_a is very small, we anticipate x to be much smaller than 0.05 M, simplifying calculations.

Mathematical Calculation of $[H^+]$

Applying the Equilibrium Expression

Using the definition of K_a:

$$
K_a = \frac{x \times x}{0.05 - x} \approx \frac{x^2}{0.05}
$$

The approximation $0.05 - x \approx 0.05$ holds because x is expected to be very small relative to 0.05 M.

Rearranged to solve for x:

$$
x^2 = K_a \times 0.05
$$

$$
x^2 = (5.0 \times 10^{-10}) \times 0.05
$$

$$
x^2 = 2.5 \times 10^{-11}
$$

$$
x = \sqrt{2.5 \times 10^{-11}} \approx 5.0 \times 10^{-6} \ \text{M}
$$

Thus, the hydrogen ion concentration in the solution, which is the [H(aq)] or $[H^+]$, is approximately 5.0×10^{-6} M.

Interpreting the Result

The calculated $[H^+]$ concentration:

- Is extremely low, reaffirming that HCN is a weak acid with minimal dissociation.
- Corresponds to a pH of:

$$
\text{pH} = -\log [\text{H}^+] = -\log (5.0 \times 10^{-6}) \approx 5.3
$$

This pH value places the solution in the mildly acidic range but not strongly acidic, consistent with weak acid behavior.

Comparing the [H(aq)] to Given Options: 5.0×10^{-10} M vs. 5.0×10^{-4} M

The initial question asks whether the H^+ concentration in 0.05 M HCN is approximately:

- 5.0×10^{-10} M
- 5.0×10^{-4} M

Based on calculations, the correct $[H^+]$ is approximately 5.0×10^{-6} M, which is significantly larger than 5.0×10^{-10} M but smaller than 5.0×10^{-4} M.

Implications:

- 5.0×10^{-10} M is too small; it would suggest almost no dissociation, inconsistent with the calculation.
- 5.0×10^{-4} M is larger than our calculated $[H^+]$, but considering approximate methods and real solution behavior, it is closer to the order of magnitude than 5.0×10^{-10} M.

Conclusion:

The most accurate estimate based on the equilibrium calculation is approximately 5.0×10^{-6} M, which does not exactly match either of the options but is closest to 5.0×10^{-4} M in the context of typical exam choices, indicating that the approximate $[H^+]$ is on the order of 10^{-6} M, significantly higher than 10^{-10} M but lower than 10^{-4} M.

Additional Factors and Considerations

Effect of Concentration on Dissociation

- The degree of ionization, α, can be approximated as:

$$
\alpha = \frac{[\text{H}^+]}{[\text{HCN}]_{initial}} \approx \frac{5.0 \times 10^{-6}}{0.05} = 1.0 \times 10^{-4}
$$

- Only about 0.01% of HCN molecules dissociate, confirming weak acid behavior.

Impact of Temperature and Ionic Strength

- The calculation assumes standard conditions; any changes in temperature could slightly alter K_a.
- Ionic strength effects, especially in more concentrated solutions, can influence dissociation, but at 0.05 M, this effect is minimal.

Practical Implications and Applications

- Understanding the $[H^+]$ concentration helps in predicting the pH of HCN solutions.
- Critical in contexts such as safety considerations, environmental chemistry, and industrial processes involving cyanide compounds.
- The weak dissociation emphasizes the toxicity of cyanide but also explains why dilute solutions are less immediately hazardous in terms of free H^+ ions.

Summary and Final Insights

- The key to solving the problem lies in applying the equilibrium expression for a weak acid, recognizing the small \(K_a \), and making appropriate approximations.
- The calculated $[H^+]$ in 0.05 M HCN is approximately 5.0×10^{-6} M.
- This value signifies a weakly acidic solution with a pH around 5.3.
- The options given in the original question — 5.0×10^{-10} M and 5.0×10^{-4} M — do not precisely match the calculated value but help contextualize the scope of dissociation.
- The most reasonable choice based on the calculation is closer to 5.0×10^{-4} M, considering typical approximation errors and the nature of weak acids.

Final Remarks

Understanding the dissociation of weak acids like HCN requires a firm grasp of equilibrium principles and careful calculations. While the theoretical approach yields an $[H^+]$ concentration of about 5.0×10^{-6} M, real-world factors may slightly adjust this value. Recognizing the order of magnitude and the relationship between \(K_a \), initial concentration, and ionization extent is crucial for chemists, environmental scientists, and industrial professionals dealing with cyanide chemistry.

This detailed analysis offers a thorough foundation for interpreting weak acid behavior, calculating equilibrium concentrations, and understanding the implications of \(K_a \) values in aqueous solutions.

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