For Weak Acid, HX, Ka = 6.9 10-6. Calculate The PH Of A 0.13 M Solution Of HX.A) 0.89B) 3.02C) 6.05D)

For Weak Acid, HX, Ka = 6.9 × 10-6. Calculate The pH Of A 0.13 M Solution Of HX. A) 0.89 B) 3.02 C) 6.05 D)

Introduction to Acid-Base Chemistry and the Concept of pH

Understanding Acids and Bases

In chemistry, acids are substances that release hydrogen ions (H^+) when dissolved in water, while bases release hydroxide ions (OH^-). The strength of an acid is determined by its ability to ionize in water. Strong acids completely dissociate, whereas weak acids only partially ionize.

The pH Scale

pH is a logarithmic scale used to measure the acidity or alkalinity of a solution. It is defined as:

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

where $[\text{H}^+]$ is the molar concentration of hydrogen ions in the solution.

Understanding Weak Acids and Their Dissociation

Properties of Weak Acids

Weak acids, such as HX with a given dissociation constant (Ka), do not fully dissociate in water. Instead, they establish an equilibrium between the undissociated acid and the ions produced.

The Dissociation Equation for HX

The dissociation of a weak acid HX can be represented as:

$$
\text{HX} \leftrightarrow \text{H}^+ + \text{X}^-
$$

The equilibrium expression (Ka) is:

$$
K_a = \frac{[\text{H}^+][\text{X}^-]}{[\text{HX}]}
$$

where:
- $[\text{H}^+]$ = concentration of hydrogen ions
- $[\text{X}^-]$ = concentration of the conjugate base
- $[\text{HX}]$ = concentration of undissociated acid

Given Data and Objective

- Acid: HX (weak acid)
- Acid dissociation constant: $K_a = 6.9 \times 10^{-6}$
- Concentration of HX solution: $[\text{HX}]_0 = 0.13\, M$

Objective: Calculate the pH of this solution and determine which of the options (A, B, C, D) is correct.

Step-by-Step Calculation of pH

Step 1: Set Up the ICE Table

To analyze the dissociation, we use an ICE table, which stands for Initial, Change, Equilibrium.

	HX (initial)	HX (change)	HX (equilibrium)	H^+ and X^- (initial)	H^+ and X^- (change)	H^+ and X^- (equilibrium)
Concentration	0.13 M	-x	0.13 - x	0	+x	x

- Initial concentration: 0.13 M
- Change: Dissociation of HX produces x mol/L of H^+ and X^-
- Equilibrium concentrations: $0.13 - x$ for HX, x for H^+ and X^-

Step 2: Write the Ka Expression

Using the equilibrium concentrations:

$$
K_a = \frac{x \times x}{0.13 - x} = \frac{x^2}{0.13 - x}
$$

Given that $K_a = 6.9 \times 10^{-6}$, and assuming x is small compared to 0.13, we can approximate:

$$
0.13 - x \approx 0.13
$$

which simplifies the calculation.

Step 3: Solve for x

Substituting into the Ka expression:

$$
6.9 \times 10^{-6} = \frac{x^2}{0.13}
$$

Therefore,

$$
x^2 = 6.9 \times 10^{-6} \times 0.13
$$

Calculating:

$$
x^2 = 8.97 \times 10^{-7}
$$

$$
x = \sqrt{8.97 \times 10^{-7}} \approx 9.48 \times 10^{-4}
$$

This value of x represents the concentration of H^+ ions in the solution.

Step 4: Calculate the pH

Using the definition:

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

$$
\text{pH} = -\log (9.48 \times 10^{-4}) \approx -\left[\log 9.48 + \log 10^{-4}\right]
$$

$$
\text{pH} = -(0.977 - 4) = 4 - 0.977 = 3.02
$$

Conclusion and Final Answer

The calculated pH of the 0.13 M HX solution is approximately 3.02, which matches option B.

Additional Insights and Considerations

Why Is the Approximation Valid?

The assumption that x is small relative to 0.13 M is justified because K_a is quite small, indicating weak acid behavior. When x is significantly less than 0.13, the approximation holds, simplifying calculations.

Impact of Different Concentrations

If the initial concentration of HX were much lower or higher, the value of x and subsequently the pH would change. For very dilute solutions, the approximation becomes less accurate, and solving the quadratic equation explicitly is necessary.

Understanding the Options

- Option A (0.89): Corresponds to a pH that indicates a strongly basic environment, unlikely for a weak acid.
- Option B (3.02): Matches the calculated pH, confirming the correctness.
- Option C (6.05): Slightly acidic, but not consistent with the calculations.
- Option D: Not specified here, but would likely be outside the expected range.

Summary

- The weak acid HX with $K_a = 6.9 \times 10^{-6}$ dissociates partially in water.
- For a 0.13 M solution, the equilibrium concentration of H^+ ions is approximately 9.48×10^{-4} M.
- The resulting pH is approximately 3.02, aligning with option B.
- This calculation demonstrates the importance of understanding equilibrium principles, approximations, and the relationship between Ka and pH in weak acid solutions.

Final Remarks

Understanding how to calculate pH for weak acids is fundamental in chemistry, especially in fields like environmental science, medicine, and chemical engineering. Mastery of these concepts allows scientists and engineers to predict the behavior of various chemical systems accurately.

References:

- Atkins, P., & de Paula, J. (2010). Physical Chemistry (9th ed.). Oxford University Press.
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Frequently Asked Questions

What is the formula to calculate the pH of a weak acid solution like HX?

pH = -log $[H^+]$, where $[H^+]$ is the concentration of hydrogen ions, which can be found using the acid dissociation constant (Ka) and the initial concentration of the acid.

How do you determine the concentration of H^+ ions in a weak acid solution using Ka?

Set up an expression based on the dissociation of HX: Ka = $[H^+]^2$ / [initial concentration - $[H^+]$]. Since Ka is small, $[H^+]$ is usually much less than the initial concentration, allowing for approximation.

Given Ka = 6.9×10^{-6} and initial concentration of HX = 0.13 M, what is the approximate $[H^+]$ in the solution?

Using the approximation $[H^+] \approx \sqrt{(Ka \times \text{initial concentration})} = \sqrt{(6.9 \times 10^{-6} \times 0.13)} \approx \sqrt{8.97} \times 10^{-7} \approx 9.48 \times 10^{-4}$ M.

What is the calculated pH of the 0.13 M HX solution with Ka = 6.9×10^{-6}?

pH ≈ $-\log(9.48 \times 10^{-4}) \approx 3.02$.

Which of the provided options matches the calculated pH for the weak acid solution?

Option B) 3.02 matches the calculated pH.

Why is the approximation valid for calculating the pH of this weak acid solution?

Because Ka is small, indicating the acid only partially dissociates, and the concentration of H^+ ions is much less than the initial concentration, making the approximation accurate.

Additional Resources

Weak Acid, HX, Ka = 6.9×10^{-6}: Calculating the pH of a 0.13 M Solution

Introduction

In the field of acid-base chemistry, understanding the behavior of weak acids and their dissociation in aqueous solutions is fundamental. The calculation of pH—a measure of acidity—is essential for numerous scientific, industrial, and environmental applications. When dealing with weak acids like HX, which only partially dissociate in water, the equilibrium constants and concentration calculations become vital tools. This article provides a comprehensive analysis of how to determine the pH of a 0.13 M solution of HX, given its acid dissociation constant (Ka) of 6.9×10^{-6}, exploring the underlying principles, step-by-step calculation process, and implications for broader chemical understanding.

Understanding Weak Acids and Their Dissociation

What Is a Weak Acid?

Weak acids are substances that do not completely ionize or dissociate in aqueous solutions. Unlike strong acids such as hydrochloric acid (HCl) or sulfuric acid (H_2SO_4), which dissociate entirely, weak acids establish an equilibrium between their undissociated form and their ions. The degree of dissociation depends on the acid's strength, concentration, temperature, and other solution conditions.

Acid Dissociation Constant (Ka)

The strength of a weak acid is quantitatively expressed through its acid dissociation constant, Ka. This equilibrium constant reflects the ratio of dissociated ions to undissociated acid at equilibrium:

$$ \text{HX} \rightleftharpoons \text{H}^+ + \text{X}^- $$

$$ K_a = \frac{[\text{H}^+][\text{X}^-]}{[\text{HX}]} $$

A smaller Ka indicates a weaker acid with less dissociation; a larger Ka signifies a stronger weak acid. For HX, Ka = 6.9×10^{-6} suggests it is relatively weak compared to strong acids.

Setting Up the Problem: Calculating pH

Given Data

- Concentration of HX, $[HX]_0 = 0.13$ M
- Ka = 6.9×10^{-6}

Objective

Determine the pH of this solution, which requires calculating the concentration of H^+ ions produced by the dissociation of HX.

Theoretical Framework for pH Calculation

Assumptions and Approximations

Calculating pH for weak acids involves solving an equilibrium expression. However, the process can be simplified using logical assumptions:

1. Initial concentration of HX is known (0.13 M).
2. Dissociation produces H^+ and X^- ions.
3. Change in concentration of undissociated HX is small relative to initial concentration, allowing approximations.

The typical approach involves:

- Denoting the concentration of dissociated ions as x, which equals $[\text{H}^+]$ at equilibrium.
- Expressing the equilibrium concentration of HX as $[\text{HX}] = [\text{HX}]_0 - x$.

Given the weak nature of HX, and the small Ka, the approximation $[\text{HX}] \approx [\text{HX}]_0$ often holds, simplifying calculations.

Step-by-Step Calculation of pH

1. Write the Equilibrium Expression

$$ K_a = \frac{[H^+][X^-]}{[HX]} $$

Since each mole of HX that dissociates produces one H^+ and one X^-:

$$ K_a = \frac{x \times x}{[HX]_{initial} - x} $$

$$ K_a = \frac{x^2}{0.13 - x} $$

Given that x is small compared to 0.13, approximate:

$$ 0.13 - x \approx 0.13 $$

This simplifies the expression to:

$$ K_a \approx \frac{x^2}{0.13} $$

---

2. Solve for H^+ Concentration (x)

Rearranged:

$$ x^2 = K_a \times 0.13 $$

$$ x^2 = (6.9 \times 10^{-6}) \times 0.13 $$

Calculate:

$$ x^2 = 8.97 \times 10^{-7} $$

$$ x = \sqrt{8.97 \times 10^{-7}} $$

$$ x \approx 9.48 \times 10^{-4} \, \text{M} $$

This value represents the concentration of hydrogen ions, $[\text{H}^+]$.

---

3. Calculate the pH

pH is defined as:

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

Substitute:

$$ \text{pH} = -\log (9.48 \times 10^{-4}) $$

Calculate:

$$ \text{pH} \approx - (\log 9.48 + \log 10^{-4}) $$

$$ \text{pH} \approx - (0.977 + (-4)) $$

$$ \text{pH} \approx 4 - 0.977 $$

$$ \text{pH} \approx 3.02 $$

This is consistent with the initial approximations, confirming the validity of the assumption.

Analysis of Results and Significance

Comparison with Multiple Choice Options

The calculated pH of approximately 3.02 aligns exactly with option B. This indicates that the solution is mildly acidic, as expected for a weak acid with this Ka and concentration.

Implications of the Calculation

- The pH value demonstrates the extent of dissociation of HX in water. Despite being a weak acid, at a concentration of 0.13 M, HX still produces a noticeable concentration of H^+ ions.
- The small Ka value indicates limited dissociation, but enough to produce a pH below 7, characteristic of acids.

- The calculation underscores the importance of equilibrium assumptions in acid-base chemistry, especially the approximation that simplifies the algebra without significant error.

Broader Context and Applications

Relevance in Industrial and Environmental Chemistry

Understanding weak acid dissociation is crucial in numerous real-world contexts:

- Environmental Monitoring: Acidic rain involves weak acids such as carbonic acid and sulfurous acid, whose pH can be predicted using similar calculations.
- Pharmaceuticals: Drug formulations often involve weak acids or bases, with pH influencing solubility and absorption.
- Chemical Manufacturing: Buffer solutions rely on weak acids and bases to maintain stable pH levels.

Educational Significance

This calculation serves as a fundamental example in chemistry education for illustrating equilibrium principles, approximations, and logarithmic calculations. It emphasizes critical thinking in applying assumptions and understanding limitations.

Conclusion

The process of calculating the pH of a weak acid solution, such as HX with a Ka of 6.9×10^{-6}, involves understanding the dissociation equilibrium, applying appropriate approximations, and performing logarithmic calculations. For a 0.13 M solution, the resulting pH of approximately 3.02 confirms the solution's moderate acidity. This analytical approach not only exemplifies core concepts in acid-base chemistry but also illustrates the practical utility of equilibrium constants in solving real-world problems. Recognizing the balance between dissociation extent and concentration is vital for chemists, environmental scientists, and engineers alike, highlighting the enduring importance of fundamental chemical principles in diverse fields.

References

- Brown, T. L., LeMay, H. E., Bursten, B. E., & Murphy, C. J. (2014). Chemistry: The Central Science. Pearson Education.
- Atkins, P., & de Paula, J. (2014). Chemical Principles. Oxford University Press.

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For Weak Acid Hx Ka 6 9 10 6 Calculate The Ph Of A 0 13 M Solution Of Hx A 0 89b 3 02c 6 05d

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