Calculate The PH Of Each Of The Following Solutions At 25C. A. 0.140 M HONH2 (Kb = 1.1 X 10-8) PH = B.

Calculate The pH Of Each Of The Following Solutions At 25°C. A. 0.140 M $HONH_2$ (Kb = 1.1 x 10^{-8}) pH = B.

Understanding how to determine the pH of various solutions is essential in chemistry, especially when working with acids and bases. In this article, we will focus on calculating the pH of a specific solution: 0.140 M $HONH_2$ (hydroxylamine) at 25°C, given its base dissociation constant (Kb) of 1.1 x 10^{-8}. We will explore the step-by-step process involved in such calculations, including the relevant chemical principles, formulas, and methods. Whether you're a student preparing for exams or a professional working in a laboratory, mastering pH calculations for weak bases like hydroxylamine is fundamental to understanding solution chemistry.

Introduction to pH and Its Importance

pH is a measure of the acidity or basicity of an aqueous solution. It is defined as the negative logarithm of the hydrogen ion concentration:

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

A pH less than 7 indicates acidity, a pH of 7 is neutral, and a pH greater than 7 indicates alkalinity or basicity. Accurately calculating pH helps in various applications, from industrial processes to biological systems, where the pH can influence chemical reactions and biological functions.

Understanding Hydroxylamine ($HONH_2$) as a Base

Hydroxylamine ($HONH_2$) is a weak base. When dissolved in water, it reacts to produce hydroxide ions (OH^-):

$HONH_2 + H_2O \rightleftharpoons HONH_3^+ + OH^-$

The extent of this ionization determines the solution's pH. Because hydroxylamine is a weak base, its degree of ionization is small, and we can use the base dissociation constant (Kb) to calculate the concentration of OH^- ions, and subsequently, the pH.

Given Data and Objective

Let's restate the problem with all relevant data:

- Concentration of hydroxylamine, $[HONH_2]$ = 0.140 M
- Base dissociation constant, Kb = 1.1×10^{-8}
- Temperature, T = 25°C

Our goal is to calculate the pH of this solution at 25°C.

Step-by-Step Approach to pH Calculation

The calculation involves several steps:

1. Write the Ionization Equation

The base dissociation of hydroxylamine in water is represented as:

$HONH_2 + H_2O \rightleftharpoons HONH_3^+ + OH^-$

However, since hydroxylamine acts as a base, the relevant equilibrium is:

$HONH_2 + H_2O \rightleftharpoons HONH_3^+ + OH^-$

But for simplicity, and because the primary focus is on calculating pH via OH^- concentration, the key is the equilibrium:

$HONH_2 + H_2O \rightleftharpoons HONH_2^+ + OH^-$

Note: Hydroxylamine is a weak base, and the equilibrium involves the acceptor of a proton.

2. Use Kb to Find $[OH^-]$ Concentration

The dissociation constant Kb relates the concentrations at equilibrium:

$Kb = [HONH_3^+][OH^-] / [HONH_2]$

Assuming that the initial concentration of $HONH_2$ is 0.140 M, and that the amount that dissociates is 'x', then:

- $[OH^-]$ = x
- $[HONH_3^+]$ = x
- Remaining $[HONH_2]$ = 0.140 − x

Because Kb is small, x will be very small compared to the initial concentration, so we can approximate:

$[HONH_2] \approx 0.140$ M

3. Set Up the Expression and Solve for x

Substituting into the Kb expression:

$Kb = x^2 / 0.140$

Plugging in the known value:

$1.1 \times 10^{-8} = x^2 / 0.140$

$x^2 = (1.1 \times 10^{-8})\ 0.140$

$x^2 = 1.54 \times 10^{-9}$

$x = \sqrt{(1.54 \times 10^{-9})} \approx 3.92 \times 10^{-5}$ M

Thus, the hydroxide ion concentration $[OH^-]$ is approximately 3.92×10^{-5} M.

4. Calculate the pOH

$pOH = -\log [OH^-] = -\log (3.92 \times 10^{-5})$

$pOH \approx 4.41$

5. Calculate the pH

Given that pH + pOH = 14 at 25°C:

$pH = 14 - pOH = 14 - 4.41 = 9.59$

Therefore, the pH of the 0.140 M hydroxylamine solution at 25°C is approximately 9.59.

Summary of the Calculation

Step	Description	Result
1	Write the equilibrium equation	$HONH_2 + H_2O \rightleftharpoons HONH_3^+ + OH^-$
2	Use Kb to find $[OH^-]$	$[OH^-] \approx 3.92 \times 10^{-5}$ M
3	Calculate pOH	4.41
4	Calculate pH	9.59

This demonstrates that a 0.140 M solution of hydroxylamine at 25°C is mildly basic, with a pH of approximately 9.59.

Additional Considerations and Tips for pH Calculations

- Assumptions: The approximation that x <- Temperature Dependence: Kb and pH are temperature-dependent; calculations at temperatures other than 25°C require adjusting the constants.
- Significance of Kb: Smaller Kb values indicate weaker bases, leading to higher pH values approaching neutral.
- pH of Mixed Solutions: When dealing with mixtures or multiple solutes, consider interactions and potential shifts in equilibrium.

Conclusion

Calculating the pH of weak base solutions like hydroxylamine involves understanding the equilibrium chemistry, applying the appropriate constants, and making reasonable approximations. In this case, a 0.140 M hydroxylamine solution at 25°C has a pH of approximately 9.59, indicating a mildly basic environment. Mastery of such calculations is crucial for chemists, biochemists, and anyone involved in solution chemistry, enabling precise control and understanding of chemical environments.

Further Resources

For more detailed information on pH calculations, weak acids and bases, and equilibrium chemistry, consider exploring:

- "Chemistry: The Central Science" by Brown, LeMay, Bursten
- Online tutorials on pH and equilibrium calculations
- Academic articles on hydroxylamine chemistry and applications

By understanding and practicing these calculations, you enhance your proficiency in analytical chemistry and deepen your comprehension of solution behaviors.

Frequently Asked Questions

What is the first step to calculate the pH of a 0.140 M HONH2 solution at 25°C?

The first step is to write the equilibrium expression for the base HONH2 and determine the concentration of OH- ions using its Kb value.

How do you set up the expression to find the hydroxide ion concentration (OH-) for HONH2?

Use the expression Kb = [OH-]^2 / [HONH2], assuming x ≈ [OH-], leading to x = √(Kb × initial concentration).

What is the value of Kb for HONH2 given in the problem?

The given Kb for HONH2 is 1.1×10^{-8}.

How do you calculate the hydroxide ion concentration (OH-) from the given Kb and concentration?

Calculate x = √(Kb × initial concentration) = $\sqrt{(1.1 \times 10^{-8} \times 0.140\ M)}$.

What is the approximate value of [OH−] in this solution?

$x \approx \sqrt{(1.1 \times 10^{-8} \times 0.140)} \approx \sqrt{(1.54 \times 10^{-9})} \approx 3.92 \times 10^{-5}$ M.

How do you convert hydroxide ion concentration to pOH?

pOH = −log[OH−], so pOH = $-\log(3.92 \times 10^{-5})$.

What is the calculated pOH of the solution?

pOH ≈ 4.41.

How do you find the pH from the pOH in this case?

pH = 14 - pOH = 14 - 4.41 ≈ 9.59.

What is the final pH of the 0.140 M HONH2 solution at 25°C?

The pH is approximately 9.59.

Why is it important to assume x ≈ [OH−] in this calculation?

Because the initial concentration is much larger than the amount of dissociation, allowing the approximation for simplifying the calculation.

Additional Resources

Calculate the pH of each of the following solutions at 25°C: A. 0.140 M $HONH_2$ (Kb = 1.1×10^{-8}). pH =

Understanding how to calculate the pH of solutions is a fundamental skill in chemistry, particularly when dealing with weak bases and acids. In this guide, we will explore the step-by-step process to determine the pH of a solution containing $HONH_2$ (hydroxylamine) at a given molarity and temperature. This example demonstrates how principles of equilibrium chemistry, the use of base dissociation constants (Kb), and the relation between concentration and pH come together to provide a clear picture of solution acidity or alkalinity.

Introduction to $HONH_2$ and Its Basic Nature

Hydroxylamine ($HONH_2$) is a weak base, which means it does not completely dissociate in water but reaches an equilibrium where some $HONH_2$ molecules accept protons, leading to hydroxide ion formation. The strength of this base is characterized by its base dissociation constant, Kb, which for $HONH_2$ is given as 1.1×10^{-8} at 25°C.

Knowing the molarity of $HONH_2$ and its Kb allows us to determine the concentration of hydroxide ions ($[OH^-]$) in solution, which in turn enables us to calculate the pOH and, consequently, the pH.

Step-by-Step Guide to Calculating pH for $HONH_2$ Solution

1. Write the Relevant Equilibrium Equation

The base dissociation of hydroxylamine can be represented as:

$HONH_2 + H_2O \rightleftharpoons H_3O^+ + HONH^-$

However, since $HONH_2$ acts as a base, it accepts a proton:

$HONH_2 + H_2O \rightleftharpoons HONH^- + H_3O^+$

But for calculating pH, it's often simpler to consider the base's dissociation:

$HONH_2 + H_2O \rightleftharpoons HONH^- + OH^-$

Though the dissociation of $HONH_2$ is better represented as:

$HONH_2 + H_2O \rightleftharpoons HONH^- + H_3O^+$

Alternatively, the base dissociation constant Kb relates to the protonation:

$Kb = [HONH^-][OH^-] / [HONH_2]$

This relation is more straightforward to work with in calculating hydroxide ion concentration.

2. Determine the Initial Concentration

Given:

- Molarity of $HONH_2$ = 0.140 M

This is the initial concentration before dissociation.

3. Set Up an ICE Table

	$HONH_2$ (initial)	$HONH^-$ / OH^- (change)	$HONH^-$ (equilibrium)
Initial	0.140 M	0	0
Change	−x	+x	+x
Equilibrium	0.140 − x	x	x

Where:

- x = concentration of $HONH^-$ and OH^- at equilibrium.

4. Write the Expression for Kb

Using the equilibrium expression:

$Kb = [HONH^-][OH^-] / [HONH_2]$

Plugging in the expressions:

1. $x \approx [OH^-]$ (since $HONH^-$ forms from $HONH_2$ dissociation)
2. $[HONH_2]$ at equilibrium = $0.140 - x$

So:

$1.1 \times 10^{-8} = (x)(x) / (0.140 - x)$

Given the very small Kb value, it's safe to assume that x <
$1.1 \times 10^{-8} \approx x^2 / 0.140$

5. Solve for x (Hydroxide Ion Concentration)

Rearranged:

$x^2 = (1.1 \times 10^{-8})\ 0.140$

$x^2 \approx 1.54 \times 10^{-9}$

$x \approx \sqrt{(1.54 \times 10^{-9})} \approx 3.92 \times 10^{-5}$ M

This is the concentration of OH^- ions in the solution.

Calculating pOH and pH

6. Compute pOH

$pOH = -\log [OH^-]$

$pOH = -\log (3.92 \times 10^{-5})$

$pOH \approx 4.41$

7. Find pH

Since pH + pOH = 14 at 25°C:

$pH = 14 - pOH$

$pH \approx 14 - 4.41 \approx 9.59$

Final Result

The pH of the 0.140 M $HONH_2$ solution at 25°C is approximately 9.59.

This value indicates a solution that is weakly basic, as expected for hydroxylamine, which has a relatively low Kb value but still produces enough hydroxide ions to raise the pH above 7.

Additional Considerations & Tips for Accurate pH Calculations

- Assumption Validity: When Kb is very small, assuming x <- Temperature Effects: Kb values vary with temperature; always ensure the provided Kb corresponds to the temperature in question.
- Significance of Precision: Use appropriate significant figures to reflect the precision of your input data.
- pH of Acidic vs. Basic Solutions: Remember that for acids, you typically start with H_3O^+ concentrations; for bases, with OH^-.

Summary of Key Steps

- Write the dissociation equilibrium.
- Set up an ICE table.
- Use the Kb expression to relate x (OH^- concentration) to initial concentration.
- Solve for x; calculate pOH.
- Derive pH from pOH.

By following these systematic steps, you can accurately determine the pH of weak base solutions like hydroxylamine and similar compounds, enabling better understanding of solution behavior and properties in various chemical contexts.

In conclusion, calculating the pH of a weak base solution such as $HONH_2$ requires understanding equilibrium principles, applying the appropriate dissociation constants, and performing careful algebraic calculations. This approach can be extended to a wide range of weak acids and bases, making it an essential skill for chemists and students alike.

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