If The Ka For Nh4 =5.61010, Find The Kb For Nh3. Assume That The Reaction Takes Place At 25c. Select

If The Ka For Nh4 = 5.6 × 1010, Find The Kb For Nh3. Assume That The Reaction Takes Place At 25°C. Select

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Introduction

Understanding the relationship between the acid dissociation constant (Ka) and the base dissociation constant (Kb) is fundamental in chemistry, especially when studying conjugate acid-base pairs. In this article, we will explore how to determine the Kb of ammonia (NH_3) given the Ka of its conjugate acid, ammonium ion (NH_4^+), under standard conditions at 25°C. This analysis is crucial for students and professionals working in fields such as analytical chemistry, environmental science, and chemical engineering.

The Relationship Between Ka and Kb

Conceptual Overview

In aqueous solutions, acids and bases exist in conjugate pairs. The conjugate acid-base pair is interconnected through the auto-ionization of water:

$$\mathrm{H_2O} \leftrightarrow \mathrm{H^+} + \mathrm{OH^-}$$

The dissociation constants of conjugate acid-base pairs are related by the ion product of water, Kw:

$$K_a \times K_b = K_w$$

where:

- K_a = acid dissociation constant of the acid
- K_b = base dissociation constant of the conjugate base
- K_w = ion product of water at 25°C, which is 1.0×10^{-14}

This fundamental relationship allows us to calculate one constant if the other and K_w are known.

Significance in Acid-Base Chemistry

Knowing the Ka of NH_4^+ helps in predicting the behavior of ammonium in solution, including its acidity and how it interacts with bases like NH_3. Conversely, the Kb of NH_3 determines its basicity and its ability to accept protons.

Given Data and Problem Statement

- Ka of NH_4^+: 5.6×10^{-10}
- Temperature: 25°C
- Objective: Find the Kb of NH_3

Step-by-Step Calculation

Step 1: Recognize the Relationship

Since NH_4^+ is the conjugate acid of NH_3, their dissociation constants are related by:

$$K_a \times K_b = K_w$$

At 25°C,

$$K_w = 1.0 \times 10^{-14}$$

Step 2: Rearrange the Equation

To find Kb of NH_3, rearranged as:

$$K_b = \frac{K_w}{K_a}$$

Step 3: Input the Known Values

$$K_b = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}}$$

Step 4: Perform the Calculation

Dividing the constants:

$$K_b = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}}$$

$$K_b = \frac{1.0}{5.6} \times 10^{-14 + 10}$$

$$K_b \approx 0.1786 \times 10^{-4}$$

$$K_b \approx 1.786 \times 10^{-5}$$

Final Answer:

$$\boxed{
\text{Kb of NH}_3 \approx 1.79 \times 10^{-5}
}$$

Additional Insights and Implications

Understanding the Magnitude of Kb

The calculated Kb value of approximately 1.79×10^{-5} indicates that ammonia is a weak base. This aligns with common knowledge that ammonia does not dissociate significantly in water but is still capable of accepting protons.

Acid-Base Strengths of NH_4^+ and NH_3

- NH_4^+ (Ammonium ion): Acidic in nature, with a Ka of 5.6×10^{-10}
- NH_3 (Ammonia): Basic in nature, with a Kb of approximately 1.79×10^{-5}

The strength of NH_4^+ as an acid and NH_3 as a base are reciprocal, demonstrating the conjugate relationship.

Broader Applications

Buffer Solutions

Understanding the Ka and Kb values of ammonia and ammonium ions is essential in designing buffer solutions. These solutions resist pH changes upon addition of acids or bases, which is vital in biological systems and industrial processes.

Environmental Chemistry

Ammonia plays a significant role in environmental chemistry, especially concerning water pollution and nitrogen cycling. Knowledge of its dissociation constants helps predict its behavior in natural waters, influencing toxicity and bioavailability.

Industrial Processes

Ammonia is utilized extensively in manufacturing fertilizers, cleaning agents, and in refrigeration. Accurate knowledge of its acid-base properties ensures safety and efficiency in these applications.

Summary of Key Points

- The relationship $K_a \times K_b = K_w$ at 25°C allows calculation of conjugate base or acid dissociation constants.
- Given K_a of NH_4^+ as 5.6×10^{-10}, the Kb of NH_3 is approximately 1.79×10^{-5}.
- NH_3 is a weak base, and its Kb value reflects its moderate basicity.
- Understanding these constants is crucial for applications ranging from buffer preparation to environmental monitoring.

Conclusion

In this comprehensive analysis, we've demonstrated how to determine the Kb of ammonia from the given Ka of ammonium, emphasizing the fundamental principles of acid-base chemistry. Recognizing the conjugate relationship between acids and bases and applying the ion product of water enables chemists to predict and manipulate chemical behaviors effectively. Whether in laboratory research, industrial processes, or environmental science, mastery of these concepts is essential for accurate chemical analysis and application.

References

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FAQs

Q1: Why is it important to know both Ka and Kb?

A: Knowing both constants helps predict the direction of proton transfer, strength of acids and bases, and the pH of solutions containing conjugate pairs.

Q2: Can the relationship $K_a \times K_b = K_w$ be applied at temperatures other than 25°C?

A: The relationship holds true at any temperature, but the value of K_w varies with temperature, so the calculation must account for the specific K_w at that temperature.

Q3: How does pKa relate to Ka?

A: pKa is the negative logarithm of Ka:

$$\mathrm{pK_a} = -\log K_a$$

Similarly, pKb relates to Kb.

End of Article

Frequently Asked Questions

Given that the pKa for NH_4^+ is 5.61 at 25°C, how do you calculate the Kb for NH_3?

First, find the Ka for NH_4^+ using pKa: Ka = 10^(-pKa) = 10^(-5.61) ≈ 2.45 × 10^(-6). Then, use the relation Kw = Ka × Kb, where Kw = 1.0 × 10^(-14) at 25°C. So, Kb = Kw / Ka ≈ 1.0 × 10^(-14) / 2.45 × 10^(-6) ≈ 4.08 × 10^(-9).

What is the significance of the pKa value in calculating Kb for NH_3?

The pKa value of NH_4^+ indicates the acidity of the ammonium ion. It allows us to calculate Ka, which is the acid dissociation constant, and from there, determine Kb for its conjugate base, NH_3, using the relationship Ka × Kb = Kw.

Why is it important to know the temperature (25°C) in this calculation?

Because the ionization constant Kw of water varies slightly with temperature, and at 25°C, Kw is standardly 1.0 × 10^(-14). This value is essential for accurately calculating Kb from Ka.

How do you convert pKa to Ka for NH_4^+?

Use the formula Ka = 10^(-pKa). For pKa = 5.61, Ka = 10^(-5.61) ≈ 2.45 × 10^(-6).

What assumptions are made in calculating Kb for NH_3 from pKa of NH_4^+?

The calculation assumes that the ionization of water is negligible and that the only significant dissociation is of NH_4^+ into NH_3 and H^+. It also assumes standard conditions at 25°C.

Can you explain the relationship between Ka, Kb, and Kw in this context?

Yes, for a conjugate acid-base pair, Ka (acid dissociation constant) and Kb (base dissociation constant) are related by Kw: Ka × Kb = Kw. At 25°C, Kw is 1.0 × 10^(-14).

If the pKa of NH_4^+ is 5.61, what is the pKb of NH_3?

pKb can be found using pKb = 14 - pKa. So, pKb = 14 - 5.61 = 8.39. Then, Kb = 10^(-8.39) ≈ 4.07 × 10^(-9).

What is the practical application of knowing the Kb of NH_3?

Knowing the Kb of NH_3 helps in understanding its basic strength, predicting pH in solutions, and calculating the degree of ionization and buffer capacities in chemical and biological systems.

Is the value of Kb for NH3 dependent on temperature? Why or why not?

Yes, Kb is temperature-dependent because the ionization equilibria are affected by temperature changes, which can alter the dissociation constants. The given value is applicable specifically at 25°C.

Additional Resources

Understanding the Relationship Between Ka and Kb for Ammonia (NH_3)

In the realm of chemical equilibria, particularly acid-base chemistry, the concepts of acid dissociation constants (Ka) and base dissociation constants (Kb) play a pivotal role. The problem at hand involves a classic calculation: given the acid dissociation constant (Ka) for ammonium ion (NH_4^+), find the base dissociation constant (Kb) for ammonia (NH_3) at 25°C. This task not only tests understanding of fundamental acid-base principles but also emphasizes the interconnectedness of conjugate acid-base pairs.

Fundamental Concepts: Acid-Base Equilibria and Dissociation Constants

What is Ka?

Ka, or the acid dissociation constant, quantifies the strength of an acid in solution. It expresses the extent to which an acid donates protons (H^+) to water:

$$ \text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^- $$

The expression for Ka is:

$$ K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} $$

- Higher Ka values indicate stronger acids.
- Lower Ka values indicate weaker acids.

What is Kb?

Kb is the base dissociation constant, representing the strength of a base:

$$ \text{B} + \text{H}_2\text{O} \rightleftharpoons \text{BH}^+ + \text{OH}^- $$

Its expression:

$$ K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]} $$

- Higher Kb values suggest stronger bases.
- Lower Kb values suggest weaker bases.

Relationship Between Ka and Kb

For conjugate acid-base pairs, the ion product of water (K_w) ties them together:

$$ K_w = K_a \times K_b $$

At 25°C, the value of K_w is:

$$ K_w = 1.0 \times 10^{-14} $$

Thus, for any conjugate acid-base pair:

$$ K_b = \frac{K_w}{K_a} $$

This fundamental relationship simplifies the calculation of Kb when Ka is known, and vice versa.

Given Data and Objective

- Given: K_a for NH_4^+ = 5.6×10^{10}
- Target: Find K_b for NH_3 at 25°C.

Note on the provided Ka value:
The value 5.6×10^{10} appears anomalously large for an acid dissociation constant. Typically, Ka values are less than 1 for weak acids. It's likely a typographical error, and the intended value might be 5.6×10^{-10}. This correction aligns with known acid constants of ammonium.

Assumption:
Assuming the correct Ka for NH_4^+ is 5.6×10^{-10}.

Step-by-Step Solution

Step 1: Confirm the Corrected Ka Value

Historical data indicates that:

$$ K_a (\text{NH}_4^+) \approx 5.6 \times 10^{-10} $$

This value is consistent with literature and typical for the ammonium ion.

Step 2: Use the Relationship $K_a \times K_b = K_w$

Given:

$$ K_w = 1.0 \times 10^{-14} $$

And:

$$ K_a (\text{NH}_4^+) = 5.6 \times 10^{-10} $$

The conjugate base of NH_4^+ is NH_3, whose K_b is:

$$ K_b (\text{NH}_3) = \frac{K_w}{K_a (\text{NH}_4^+)} $$

Substituting the known values:

$$ K_b (\text{NH}_3) = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} $$

Calculating Kb for NH_3

$$
K_b (\text{NH}_3) = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} = \left(\frac{1.0}{5.6}\right) \times 10^{-14 + 10} = 0.17857 \times 10^{-4}
$$

Simplify:

$$
K_b (\text{NH}_3) \approx 1.7857 \times 10^{-5}
$$

Final answer:

$$
\boxed{
K_b (\text{NH}_3) \approx 1.79 \times 10^{-5}
}
$$

This value aligns well with standard data for ammonia, which often cites K_b around 1.8×10^{-5}.

Deeper Insights into the Chemistry

Conjugate Acid-Base Pairs

Ammonia (NH_3) and ammonium (NH_4^+) are a classic conjugate base-acid pair. Recognizing this pairing is crucial:

- NH_4^+ (ammonium): conjugate acid.
- NH_3 (ammonia): conjugate base.

Their dissociation constants are related via K_w, reinforcing the idea that the strength of an acid is inversely related to that of its conjugate base.

Implications of Kb Values

The small value of K_b (~1.8×10^{-5}) for ammonia indicates:

- NH_3 is a weak base.
- It does not readily accept protons in aqueous solution.
- The equilibrium slightly favors the reactant (NH_3).

This is consistent with ammonia's known basicity and its ability to generate NH_4^+ and OH^- in solution.

Additional Considerations

Temperature Dependence of Ka and Kb

While the calculations assume temperature at 25°C, it's important to recognize that both Ka and Kb are temperature-dependent:

- Raising temperature may shift equilibria.
- Standard values are usually tabulated at 25°C for consistency.

pKa and pKb Calculations

Expressing acidity/basicity on a logarithmic scale (pKa and pKb) provides more intuitive insights:

\[

$$
pK_a = -\log K_a = -\log (5.6 \times 10^{-10}) \approx 9.25
$$

$$
pK_b = -\log K_b = -\log (1.79 \times 10^{-5}) \approx 4.75
$$

The sum:

$$
pK_a + pK_b \approx 9.25 + 4.75 = 14
$$

which aligns with the relation $pK_a + pK_b = pK_w = 14$.

Practical Applications and Context

Understanding these constants is essential in:

- Buffer solution design: Combining NH_3 and NH_4^+ to create buffers with desired pH.
- Industrial processes: Ammonia is extensively used in fertilizers, cleaning agents, and chemical syntheses.
- Environmental chemistry: Ammonia's behavior influences water quality and nitrogen cycles.

Summary and Key Takeaways

- The fundamental relationship $K_a \times K_b = K_w$ enables straightforward calculation of Kb from Ka.
- The corrected Ka value for NH_4^+ at 25°C is approximately 5.6×10^{-10}.
- Consequently, the Kb for NH_3 is approximately 1.79×10^{-5}.
- Recognizing conjugate pairs and their constants is vital for predicting solution behavior.
- The pKa and pKb values provide a more intuitive understanding of acid and base strengths.

Final Remarks

This problem encapsulates core principles of acid-base chemistry, illustrating how conjugate pairs are interconnected through fundamental constants. It underscores the importance of accurate data interpretation and reinforces the utility of the water ion product (K_w) as a bridge between acids and bases. Whether designing buffer systems or understanding environmental impacts, mastery of these concepts is crucial for chemists.

Always verify data units and values, especially when encountering anomalies like unusually large or small Ka/Kb figures, to ensure accurate and meaningful calculations.

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