

Consider The Reaction $\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$. K_b For NH_3 Is 1.8×10^{-5} At 25°C . What Is K_a

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Understanding the relationship between acids and bases is fundamental in chemistry, especially when analyzing reactions involving ammonia (NH_3). In this article, we will explore the equilibrium involved in the reaction of ammonia with water, analyze the given base dissociation constant (K_b), and determine the corresponding acid dissociation constant (K_a) for the ammonium ion (NH_4^+). This comprehensive discussion is essential for students and professionals seeking a deeper understanding of acid-base chemistry, equilibrium constants, and their interrelationships.

Introduction to Acid-Base Equilibria

Acid-base reactions are characterized by the transfer of protons (H^+ ions) between species. In aqueous solutions, acids donate protons, while bases accept protons. The strength of acids and bases is quantified by their respective dissociation constants: K_a for acids and K_b for bases.

Key Concepts:

- Acid dissociation constant (K_a): Measures the extent to which an acid dissociates in solution.
- Base dissociation constant (K_b): Measures the extent to which a base accepts protons in solution.
- Conjugate acid-base pairs: When an acid donates a proton, it forms its conjugate base; vice versa, a base accepting a proton forms its conjugate acid.

The relationship between K_a and K_b for conjugate acid-base pairs is fundamental because it allows us to calculate one constant if the other is known, given the ion product of water, K_w .

Understanding the Ammonia-Water Equilibrium

The reaction under consideration:



is a typical base dissociation process. Ammonia acts as a weak base, accepting a proton from water, which acts as an acid in this context. The equilibrium constant for this process is K_b , the base dissociation constant for ammonia.

Known Data:

- K_b for ammonia, 1.8×10^{-5} at 25°C .

- The ion product of water, $(\mathrm{K_w})$, at 25°C is (1.0×10^{-14}) .

The goal is to find the corresponding acid dissociation constant, K_a , for the ammonium ion $(\mathrm{NH_4^+})$.

Relationship Between K_a , K_b , and K_w

In aqueous chemistry, for a conjugate acid-base pair, the following relationship holds:

$$(\mathrm{K_a} \times K_b = K_w)$$

where:

- $(\mathrm{K_a})$ is the acid dissociation constant of the conjugate acid.
- $(\mathrm{K_b})$ is the base dissociation constant of the conjugate base.
- $(\mathrm{K_w})$ is the ion product of water, at 25°C, (1.0×10^{-14}) .

Using this relationship, if we know $(\mathrm{K_b})$ for ammonia, we can determine $(\mathrm{K_a})$ for ammonium:

$$(\mathrm{K_a} = \frac{(\mathrm{K_w})}{(\mathrm{K_b})})$$

This fundamental equation allows us to link the properties of conjugate acid-base pairs and understand their strengths in aqueous solutions.

Calculating K_a for $\mathrm{NH_4^+}$

Given:

- $(\mathrm{K_b})$ for ammonia = (1.8×10^{-5})
- $(\mathrm{K_w})$ at 25°C = (1.0×10^{-14})

The calculation:

$$(\mathrm{K_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}})$$

Performing the division:

$$(\mathrm{K_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = \frac{1.0}{1.8} \times 10^{-14 + 5} = 0.555 \times 10^{-9})$$

Expressed properly:

$$\boxed{\mathrm{K_a} \approx 5.56 \times 10^{-10}}$$

Thus, the acid dissociation constant for $\mathrm{NH_4^+}$ at 25°C is approximately 5.56×10^{-10} .

Implications of Ka and Kb Values

Understanding the relative magnitudes of $\mathrm{K_a}$ and $\mathrm{K_b}$ helps in predicting the behavior of conjugate acid-base pairs in solution:

- A small $\mathrm{K_a}$ indicates a weak acid, meaning $\mathrm{NH_4^+}$ does not readily donate protons.
- A small $\mathrm{K_b}$ indicates a weak base, meaning $\mathrm{NH_3}$ does not readily accept protons.

Since $\mathrm{K_a}$ for ammonium is extremely small, it shows that ammonium is a weak acid, consistent with expectations for its conjugate relationship with ammonia.

Applications:

- Buffer solutions involving ammonia and ammonium.
- pH calculations for solutions containing ammonia or ammonium.
- Understanding equilibrium shifts in industrial and biological processes.

pKa Calculations and pH Considerations

The pKa of ammonium provides insight into the acidity of its solution:

$$\mathrm{pK_a} = -\log_{10}(\mathrm{K_a})$$

Calculating:

$$\mathrm{pK_a} = -\log_{10}(5.56 \times 10^{-10}) \approx 9.25$$

This indicates that ammonium has a pKa of approximately 9.25, reflecting its weak acidic nature in aqueous solution.

Significance:

- pKa around 9.25 means that in solutions with pH below 9.25, ammonium remains largely in its protonated form.
- The pH of a solution containing ammonium can be estimated using this pKa value and the concentration of ammonium ions.

Practical Examples and Applications

Understanding the relationship between K_a and K_b in the context of ammonia and ammonium has several practical applications:

1. Buffer Solutions:

- Ammonia and ammonium form an effective buffer system in the pH range around 9.25.
- These buffers are used in laboratory settings, in biological systems, and industrial processes.

2. Environmental Chemistry:

- Ammonia's behavior in water bodies impacts water quality and aquatic life.
- The equilibrium between ammonia and ammonium affects nitrogen cycling and pollution control.

3. Industrial Processes:

- Ammonia is used in manufacturing fertilizers, cleaning agents, and refrigeration.
- Controlling pH and understanding dissociation constants are critical for process optimization.

4. Biological Systems:

- The ammonium ion plays a role in nitrogen metabolism.
- The weak acidity of ammonium influences its transport across cell membranes and enzyme activities.

Summary and Key Takeaways

- The reaction of ammonia with water is a classic example of a weak base reacting with water to produce ammonium and hydroxide ions.
- The base dissociation constant K_b for ammonia is 1.8×10^{-5} .
- Using the relationship $K_a \times K_b = K_w$, the acid dissociation constant K_a for ammonium is approximately 5.56×10^{-10} .
- The pKa of ammonium is approximately 9.25, indicating its weak acid nature.
- Understanding these constants aids in predicting solution behavior, designing buffer systems, and analyzing chemical equilibria.

Final Note:

The interplay between K_a and K_b underscores the importance of conjugate acid-base pairs in chemistry. Accurate knowledge of these constants enables chemists to

manipulate pH, design effective buffers, and understand biological and environmental processes involving nitrogen compounds.

References:

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- Zumdahl, S. S., & Zumdahl, S. A. (2014). Chemistry: An Atoms First Approach. Cengage Learning.
- Lyman, J. (2014). Principles of Chemistry. OpenStax.

Frequently Asked Questions

What is the relationship between K_b for NH_3 and K_a for NH_4^+ ?

K_a and K_b are related through the water ionization constant K_w , such that $K_a \times K_b = K_w$. For NH_3 and NH_4^+ , $K_a = K_w / K_b$.

Given K_b for NH_3 is 1.8×10^{-5} at 25°C , what is the K_a for NH_4^+ ?

Using $K_w = 1.0 \times 10^{-14}$, $K_a = K_w / K_b = (1.0 \times 10^{-14}) / (1.8 \times 10^{-5}) \approx 5.56 \times 10^{-10}$.

Why is it important to know K_a and K_b values for conjugate acid-base pairs?

Knowing K_a and K_b helps determine the strengths of acids and bases, predict the direction of reactions, and calculate pH in buffer solutions.

How does the value of K_a for NH_4^+ influence its acidity in aqueous solution?

A smaller K_a value indicates a weaker acid; for NH_4^+ , a K_a of approximately 5.56×10^{-10} suggests it is a weak acid that only slightly dissociates in water.

What is the significance of the water ionization constant K_w in calculating K_a from K_b ?

K_w represents the auto-ionization of water (1.0×10^{-14} at 25°C), and it links the acid and base dissociation constants through the relationship $K_a \times K_b = K_w$.

Can you explain how to derive K_a from K_b for a conjugate

acid-base pair?

Yes, $K_a = K_w / K_b$. For NH_4^+ , since K_b for NH_3 is known, you divide K_w by K_b to find K_a .

If you know the K_b for NH_3 , how can you predict the pH of a solution containing NH_4^+ ?

By calculating the K_a of NH_4^+ and then its corresponding hydrogen ion concentration, you can determine the pH of the solution.

What does a very small K_a value indicate about the conjugate base's strength?

A small K_a for the conjugate acid indicates that its conjugate base is relatively strong, meaning it can accept protons more readily.

How does temperature affect the values of K_a and K_b for a conjugate acid-base pair?

Temperature can influence the dissociation constants; generally, increasing temperature can increase or decrease K_a and K_b depending on whether the reactions are endothermic or exothermic.

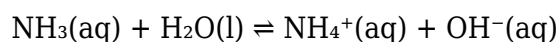
Additional Resources

Consider the Reaction $\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$: Determining K_a from K_b at 25°C

Introduction

In the realm of aqueous chemistry, understanding the acid-base properties of substances is fundamental. Among these, ammonia (NH_3) stands out as a classic example of a weak base. Its ability to accept protons from water leads to a dynamic equilibrium characterized by a specific equilibrium constant, K_b . Conversely, the conjugate acid of ammonia, the ammonium ion (NH_4^+), exhibits acidic properties, described by its own equilibrium constant, K_a . This interconnectedness of acid-base equilibria allows chemists to derive one constant from the other, providing deeper insights into the behavior of these species in aqueous solutions.

In this article, we delve into the specific reaction:



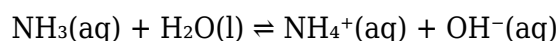
with a known base dissociation constant, K_b , for ammonia at 25°C of 1.8×10^{-5} . Our objective is to determine the acid dissociation constant, K_a , for the ammonium ion, leveraging the fundamental relationship that links these two equilibrium constants through the ion product of water, K_w . The process involves a detailed exploration of acid-base equilibria, the role of K_w , and the calculations

that connect K_b to K_a , thereby showcasing the interconnected nature of aqueous acid-base chemistry.

Understanding the Equilibrium: Ammonia and Water

The Role of Ammonia as a Weak Base

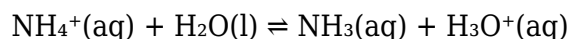
Ammonia (NH_3) is classified as a weak base because it does not fully dissociate in water but undergoes partial proton acceptance from water molecules. Its base dissociation equilibrium is represented as:



The equilibrium constant for this process, K_b , quantifies the extent to which ammonia accepts protons, and at 25°C , this value is given as 1.8×10^{-5} .

The Conjugate Acid: Ammonium Ion

The ammonium ion (NH_4^+), formed when ammonia accepts a proton, exhibits acidic behavior. It can donate a proton back to water, establishing its own equilibrium:

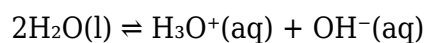


The corresponding acid dissociation constant, K_a , describes this propensity to donate protons.

Relationship Between K_a and K_b : The Fundamental Link

The Ion Product Constant of Water (K_w)

At 25°C , pure water undergoes auto-ionization:



with an equilibrium constant, K_w , equal to 1.0×10^{-14} . This value is crucial because it connects the acid and base equilibria of conjugate pairs.

The Mathematical Relationship

For conjugate acid-base pairs, the relationship between K_a and K_b is expressed as:

$$K_a \times K_b = K_w$$

This fundamental equation means that knowing one constant allows the calculation of the other, provided the temperature is constant (here, 25°C).

Calculating K_a for NH_4^+

Given that:

- K_b for $\text{NH}_3 = 1.8 \times 10^{-5}$
- K_w at $25^\circ\text{C} = 1.0 \times 10^{-14}$

We can determine the K_a for NH_4^+ using the relation:

$$K_a = K_w / K_b$$

Substituting the known values:

$$K_a = (1.0 \times 10^{-14}) / (1.8 \times 10^{-5})$$

Calculating the quotient:

$$K_a \approx (1.0 / 1.8) \times 10^{(-14 + 5)} = 0.555... \times 10^{-9}$$

Expressed in proper scientific notation:

$$K_a \approx 5.56 \times 10^{-10}$$

This value indicates that ammonium is a weak acid, with a K_a significantly less than 1, consistent with its weak acid behavior in aqueous solution.

Significance of the K_a Value

Understanding the magnitude of K_a for NH_4^+ provides insights into its acidity. A K_a of approximately 5.56×10^{-10} suggests that only a small fraction of ammonium ions donate protons to water at equilibrium, emphasizing its weak acid nature. This low K_a value also influences the pH of ammonium salt solutions and affects buffer systems containing ammonium and ammonia.

Broader Implications in Acid-Base Chemistry

pKa and pKb Calculations

From the K_a value, one can derive the pKa:

$$\text{p}K_a = -\log(K_a) \approx -\log(5.56 \times 10^{-10}) \approx 9.25$$

Since the pKb of ammonia is known as:

$$\text{p}K_b = -\log(K_b) \approx -\log(1.8 \times 10^{-5}) \approx 4.74$$

and the sum of pKa and pKb for conjugate acid-base pairs equals 14 at 25°C , this confirms the consistency:

$$\text{p}K_a + \text{p}K_b \approx 9.25 + 4.74 \approx 13.99 \approx 14$$

which aligns with the theoretical expectation.

Buffer Systems and Practical Applications

The ammonium/ammonia buffer system plays a vital role in biological and industrial processes. For example, the blood buffer system maintains pH around 7.4, involving ammonium salts. Precise knowledge of K_a and K_b values enables chemists and medical professionals to predict the behavior of these buffers in various conditions.

Analytical Methods and Experimental Determination

While the theoretical calculations based on known constants are straightforward, experimental validation is essential. Techniques such as potentiometric titrations, spectrophotometry, and pH measurements allow accurate determination of K_a and K_b values.

1. Potentiometric Titration: Involves titrating ammonium salts with a strong base or acid and measuring pH changes to find the equivalence point, from which K_a or K_b can be deduced.
2. Spectrophotometry: Monitors changes in absorbance related to protonation states to analyze equilibria.
3. pH Measurement: Using precise pH meters to track the buffer capacity and infer equilibrium constants.

These methods reinforce the theoretical values and provide insights into real-world variations due to temperature, ionic strength, and other factors.

Conclusion

The interrelationship between K_b and K_a is a cornerstone of acid-base chemistry, exemplified by the ammonium/ammonia system. Starting from the known K_b of ammonia (1.8×10^{-5} at 25°C), we have demonstrated how to derive the K_a of the ammonium ion, arriving at approximately 5.56×10^{-10} . This calculation underscores the weak acid nature of NH_4^+ and highlights the elegant symmetry of conjugate pair equilibria governed by K_w .

Understanding these constants not only enriches our theoretical grasp of chemical equilibria but also informs practical applications ranging from buffer design in laboratories to physiological pH regulation in biological systems. As aqueous chemistry continues to evolve, the fundamental principles exemplified by the ammonia-ammonium system remain central to deciphering complex biochemical and industrial processes.

This analytical approach exemplifies the power of foundational constants and relationships in chemistry, enabling scientists to predict and manipulate the behavior of substances in aqueous environments with confidence and precision.

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