Consider A 1.0-L Solution That Is Initially 0.690 M NH3 And 0.540 M NH4Cl At 25 C. What Is The PH Of

Consider A 1.0-L Solution That Is Initially 0.690 M NH3 And 0.540 M NH4Cl At 25 °C. What Is The pH Of This Solution?

Understanding the pH of a solution containing both a weak base (NH3) and its conjugate acid (NH4Cl) requires a comprehensive analysis of the chemical equilibria involved. This type of solution is known as a buffer solution, which resists changes in pH upon addition of small amounts of acid or base. To determine the pH accurately, we need to examine the relevant equilibria, utilize the appropriate equilibrium constants, and consider the initial concentrations of the species involved. This article will guide you through the process step by step, providing insights into the underlying chemistry and calculations necessary to find the pH of this solution.

Understanding the Components of the Solution

1. The Weak Base: Ammonia (NH_3)

- Ammonia is a common weak base.
- It reacts with water to produce ammonium ions (NH_4^+) and hydroxide ions (OH^-).
- The equilibrium can be represented as:

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

2. The Conjugate Acid: Ammonium Chloride (NH_4Cl)

- NH_4Cl is a salt that dissociates completely in water to produce NH_4^+ and Cl^-.
- The presence of NH_4^+ establishes an acidic component in the buffer system.

3. Buffer Solution Characteristics

- Since the solution contains both NH_3 and NH_4Cl, it acts as a buffer.
- The pH is governed by the relative concentrations of the weak base and its conjugate acid, and their acid-base equilibrium.

Relevant Equilibrium Constants and Data

1. The Base Dissociation Constant (Kb) for NH_3

- Known as Kb, it quantifies the tendency of NH_3 to accept protons.
- Standard value at 25°C:

$K_b \approx 1.76 \times 10^{-5}$

2. The Acid Dissociation Constant (Ka) for NH_4^+

- The conjugate acid of NH_3, NH_4^+, has an associated Ka.
- Since NH_4^+ is the conjugate acid, Ka relates to Kb via the water ionization constant:

$K_a \times K_b = K_w$

- Where Kw is the ionization constant of water at 25°C:

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

- Therefore,

$K_a (NH_4^+) = K_w / K_b = (1.0 \times 10^{-14}) / (1.76 \times 10^{-5}) \approx 5.68 \times 10^{-10}$

3. pKa of NH_4^+

- Calculated from Ka:

$pK_a = -\log(K_a) \approx -\log(5.68 \times 10^{-10}) \approx 9.25$

Setting Up the Buffer Equation

1. The Henderson-Hasselbalch Equation

- The pH of a buffer solution is given by:

$pH = pK_a + \log([A^-]/[HA])$

- Where:
- $[A^-]$ is the concentration of the conjugate base (NH_3)
- $[HA]$ is the concentration of the conjugate acid (NH_4^+)

2. Identifying Species in the Solution

- Initial concentrations:
- $[NH_3]$ = 0.690 M
- $[NH_4Cl]$ provides $[NH_4^+]$ = 0.540 M

- Since NH_4Cl dissociates completely:
- $[NH_4^+] \approx 0.540$ M
- NH_3 is a weak base, and some of it will react with water, but if the initial concentrations are known, and the solution is at equilibrium, we can approximate the pH using the buffer equation.

Calculating the pH of the Buffer Solution

1. Applying the Henderson-Hasselbalch Equation

- Using the known values:

$$pH = pKa + \log([NH_3]/[NH_4^+])$$

- Plugging in values:

$$pH = 9.25 + \log(0.690 / 0.540)$$

2. Performing the Logarithmic Calculation

- Calculate the ratio:

$$0.690 / 0.540 \approx 1.278$$

- Find the logarithm:

$$\log(1.278) \approx 0.106$$

- Final pH:

$$pH \approx 9.25 + 0.106 \approx 9.356$$

Refining the Calculation: Considering

Equilibrium Dynamics

While the Henderson-Hasselbalch equation provides a robust approximation for buffer solutions, understanding the underlying equilibrium can help refine the pH estimate, particularly when initial concentrations are close or when precision is critical.

1. The Role of Ammonia's Ionization

- Some NH_3 molecules will accept protons from water, generating additional NH_4^+ and OH^-.
- The equilibrium:

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

- The extent of this reaction depends on the initial concentration of NH_3 and the Kb value.

2. Approximate Calculation of OH^- Concentration

- Set up an ICE table (Initial, Change, Equilibrium) for NH_3:

Species	Initial (M)	Change	Equilibrium (M)
NH_3	0.690	-x	0.690 - x
NH_4^+	0.540	+x	0.540 + x
OH^-	-	+x	x

- The equilibrium expression:

$$K_b = [NH_4^+][OH^-] / [NH_3]$$

- Substituting:

$$1.76 \times 10^{-5} = (0.540 + x)\, x / (0.690 - x)$$

- Since Kb is small and x is expected to be small, approximate:

$$0.540 + x \approx 0.540$$

$$0.690 - x \approx 0.690$$

- Simplify:

$$1.76 \times 10^{-5} \approx (0.540)\, x / 0.690$$

$$x \approx (1.76 \times 10^{-5})\, 0.690 / 0.540 \approx 2.25 \times 10^{-5}\ M$$

- Concentration of OH^-:

$[OH^-] \approx x \approx 2.25 \times 10^{-5}$ M

- Calculate pOH:

$pOH = -\log([OH^-]) \approx -\log(2.25 \times 10^{-5}) \approx 4.65$

- Finally, pH:

$pH = 14 - pOH \approx 14 - 4.65 \approx 9.35$

This refined calculation aligns closely with the approximation from the Henderson-Hasselbalch equation, confirming that the pH is approximately 9.36.

Conclusion: The pH of the Solution

Based on the calculations and the principles of buffer chemistry, the pH of a 1.0-L solution initially containing 0.690 M NH_3 and 0.540 M NH_4Cl at 25°C is approximately 9.36. This value reflects the buffer nature of the solution, where the weak base NH_3 and its conjugate acid NH_4^+ work together to maintain a relatively high pH. The close agreement between the Henderson-Hasselbalch approximation and the equilibrium-based calculation underscores the robustness of buffer equations in predicting pH in systems involving weak acids and bases.

Implications and Applications

Understanding how to determine the pH of such buffer systems is crucial in various fields:
- Chemical manufacturing: Designing buffers for industrial processes.
- Biology and medicine: Maintaining physiological pH stability.
- Environmental science: Assessing water quality and buffering capacity.
- Laboratory research: Preparing solutions with precise pH values.

Mastering these calculations enables chemists, biologists, and environmental scientists to predict and control pH effectively, ensuring optimal conditions for their specific applications.

Summary

- The key to calculating the pH of a buffer solution is understanding the relationship between the weak base and its conjugate acid.
- The Henderson-Hasselbalch equation provides a quick and reliable estimate.
- Equ

Frequently Asked Questions

What is the initial pH of a 1.0 L solution containing 0.690 M NH3 and 0.540 M NH4Cl at 25°C?

The initial pH can be calculated using the Henderson-Hasselbalch equation, considering NH_3 as the base and NH_4Cl as the conjugate acid. First, find the pKa of NH_4^+: pKa ≈ 9.25. Then, pH = pKa + log([NH_3]/[NH_4^+]) = 9.25 + log(0.690/0.540) ≈ 9.25 + 0.106 ≈ 9.36.

How does the presence of NH4Cl affect the pH of the solution compared to pure NH3?

NH_4Cl provides NH_4^+ ions, which act as a weak acid, lowering the pH relative to pure NH_3 solution. The buffer system of NH_3 and NH_4^+ maintains a pH around 9.36, which is slightly acidic for a basic solution.

Can the pH of this solution be determined using the Henderson-Hasselbalch equation?

Yes, because the solution contains a weak base (NH_3) and its conjugate acid (NH_4^+), the Henderson-Hasselbalch equation is appropriate for calculating the pH.

What role does temperature play in calculating the pH of this solution?

Temperature affects the acid dissociation constant (Ka) and pKa. At 25°C, the pKa of NH_4^+ is approximately 9.25, which is used in the calculation. Changes in temperature would alter these values and thus the pH.

What assumptions are made when calculating the pH of this buffer solution?

The calculation assumes the solution behaves as an ideal buffer, the concentrations are accurate, and activity coefficients are close to 1. It also assumes the initial concentrations of NH_3 and NH_4^+ are unchanged by any reactions or volume changes.

How would the pH change if additional NH3 were added to the solution?

Adding more NH3 would increase the [NH3], raising the ratio [NH3]/[NH4+], and thus increase the pH, making the solution more basic.

What is the significance of the pKa value in determining the pH of this buffer system?

The pKa indicates the pH at which the acid and conjugate base are at equal concentrations. It serves as a reference point for calculating the pH of the buffer solution using the Henderson-Hasselbalch equation.

Are there any limitations to using the Henderson-Hasselbalch equation in this scenario?

Yes, it assumes ideal behavior and ignores activity coefficients, which can be significant at higher concentrations. Also, it is most accurate for weak acid/base systems with moderate concentrations and may not account for all ionic interactions in complex solutions.

Additional Resources

Consider A 1.0-L Solution That Is Initially 0.690 M NH3 And 0.540 M NH4Cl At 25°C. What Is The pH Of

Understanding the pH of a solution containing both ammonia (NH_3) and ammonium chloride (NH_4Cl) is fundamental in acid-base chemistry. When examining such a mixture, it's crucial to consider the principles of buffer solutions, the behavior of weak bases and their conjugate acids, and how their equilibrium dynamics influence the pH. This review delves into the detailed process of calculating the pH of this specific solution, exploring the underlying chemistry, relevant equations, assumptions, and the implications of the findings.

Introduction to the System

The given solution comprises 0.690 M NH_3 (ammonia) and 0.540 M NH_4Cl (ammonium chloride) in a 1.0-liter container at 25°C. Ammonia is a weak base, and ammonium chloride is its conjugate acid salt. The mixture essentially forms a buffer system, resisting pH changes upon the addition of acids or bases. Understanding its equilibrium behavior requires analyzing the dissociation of ammonia in water and the protonation equilibrium of ammonium.

The key to determining the pH lies in recognizing this as a classic buffer scenario where both the weak base (NH_3) and its conjugate acid (NH_4^+ from NH_4Cl) coexist in solution. This makes the Henderson-Hasselbalch equation a suitable tool for estimating the pH.

Fundamental Chemistry Principles

Ammonia as a Weak Base

Ammonia (NH_3) reacts with water as follows:

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

The base dissociation constant (Kb) for ammonia at 25°C is approximately 1.8×10^{-5}. This value indicates the extent to which ammonia accepts a proton from water, producing ammonium (NH_4^+) and hydroxide ions (OH^-).

Ammonium as a Conjugate Acid

NH_4Cl is a salt that dissociates completely in water:

$NH_4Cl \rightarrow NH_4^+ + Cl^-$

The ammonium ion (NH_4^+) can donate a proton to water:

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$

The acid dissociation constant (Ka) for ammonium can be derived from Kw and Kb:

$Ka = Kw / Kb$

where $Kw = 1.0 \times 10^{-14}$ at 25°C.

Calculating the pH: Step-by-Step Approach

Step 1: Recognize the Buffer System

Since both NH_3 and NH_4^+ are present, the solution behaves as a buffer. The pH can be estimated using the Henderson-Hasselbalch equation:

$pH = pKa + \log([A^-]/[HA])$

In this context:
- $[A^-]$ = concentration of NH_3 (weak base)
- $[HA]$ = concentration of NH_4^+ (conjugate acid)

Given initial concentrations:
- $[NH_3] = 0.690$ M
- $[NH_4^+] = 0.540$ M (from NH_4Cl)

Step 2: Determine pKa of NH_4^+

Calculate Ka from Kb:

$$Ka = Kw / Kb = (1.0 \times 10^{-14}) / (1.8 \times 10^{-5}) \approx 5.56 \times 10^{-10}$$

Then,

$$pKa = -\log(Ka) \approx -\log(5.56 \times 10^{-10}) \approx 9.25$$

Step 3: Apply the Henderson-Hasselbalch Equation

Using the initial concentrations as approximate values for $[A^-]$ and $[HA]$, the pH estimate is:

$$pH \approx 9.25 + \log(0.690 / 0.540) \approx 9.25 + \log(1.278)$$

Calculating the logarithm:

$$\log(1.278) \approx 0.106$$

Therefore,

$$pH \approx 9.25 + 0.106 \approx 9.36$$

This is an initial estimate. However, because the weak base and conjugate acid are not in equal concentrations, a more precise calculation considers the equilibrium shifts.

Refined Calculation Using Equilibrium Expressions

While the Henderson-Hasselbalch provides a practical estimate, a more rigorous approach involves setting up the equilibrium equations explicitly, considering the initial concentrations, and solving for the hydroxide ion concentration.

Step 4: Set Up the Equilibrium Expressions

Let’s define:

- x = concentration of OH^- produced from ammonia dissociation
- The initial concentrations are:
- $[NH_3] = 0.690$ M
- $[NH_4^+] = 0.540$ M

The dissociation of ammonia:

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

At equilibrium:

- $[NH_3] = 0.690 - x$
- $[NH_4^+] = 0.540 + x$
- $[OH^-] = x$

Using Kb expression:

$K_b = [NH_4^+][OH^-] / [NH_3]$

Plugging in the known values:

$1.8 \times 10^{-5} = (0.540 + x) x / (0.690 - x)$

Given that x is small compared to initial concentrations, we can approximate:

$0.540 + x \approx 0.540$
$0.690 - x \approx 0.690$

Thus,

$x \approx (1.8 \times 10^{-5})\ 0.690 / 0.540 \approx (1.8 \times 10^{-5})\ 1.278 \approx 2.3 \times 10^{-5}$ M

This yields:

$pOH = -\log(2.3 \times 10^{-5}) \approx 4.64$

And:

$pH = 14 - pOH \approx 14 - 4.64 \approx 9.36$

This refined calculation confirms our initial estimate.

Implications and Features of the System

Buffer Capacity

The mixture’s pH (~9.36) indicates a basic environment, primarily due to excess ammonia. Because both NH_3 and NH_4^+ are present, the system exhibits significant buffer capacity, which resists drastic pH changes upon addition of small amounts of acids or bases.

Features:
- Effective in maintaining pH in the basic range
- Useful in applications requiring a stable alkaline environment

Limitations:
- Limited buffer capacity if large amounts of acid/base are added
- pH may shift significantly if initial concentrations change

Practical Applications

- Buffer solutions in chemical laboratories
- pH regulation in industrial processes
- Biological systems where ammonia-based buffers are relevant

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Pros and Cons of Using the Henderson-Hasselbalch Equation

Pros:
- Simple and quick calculations
- Requires only initial concentrations and pKa
- Suitable for weak acid/base conjugate pairs

Cons:
- Assumes negligible changes in concentrations (approximations)
- Less accurate for systems with large shifts or where x is not small
- Cannot account for complex equilibria beyond the simple buffer system

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Conclusion

The pH of the solution containing 0.690 M NH_3 and 0.540 M NH_4Cl at 25°C is approximately 9.36. This value reflects the buffer nature of the mixture, dominated by the equilibrium between ammonia and ammonium ions. The calculation underscores the importance of understanding equilibrium principles, dissociation constants, and the practical utility of the Henderson-Hasselbalch equation in estimating pH in buffer systems. While simplified models provide good estimates, more detailed equilibrium calculations can refine the pH value further, especially in systems where concentrations are comparable or shifts are significant.

Such analyses are essential in various scientific and industrial contexts, enabling precise control of pH-dependent processes. Recognizing the strengths and limitations of these approaches ensures accurate interpretations and effective application in real-world scenarios.

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