What Is The PH Of A Solution Containing 0.1 Mole Of Ephedrine (pKb = 4.64) And 0.05 Moles Of Ephedrine

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Understanding the pH of a solution containing ephedrine involves delving into its chemical properties, dissociation behavior, and how its basic nature influences the overall acidity or alkalinity of the mixture. Ephedrine, a sympathomimetic amine commonly used as a decongestant and stimulant, exhibits basic properties owing to its amine group. When considering a solution with specified amounts of ephedrine—namely, 0.1 mole and 0.05 Moles—the pH calculation becomes an intriguing exercise in acid-base chemistry, particularly within the context of its pKb value of 4.64. This article aims to thoroughly explore how to determine the pH of such a solution, providing insights into the underlying principles, step-by-step calculations, and practical implications.

Understanding Ephedrine and Its Basic Nature

What Is Ephedrine?

Ephedrine is a naturally occurring alkaloid derived from plants of the genus Ephedra. Its chemical structure includes a phenethylamine core with hydroxyl and methyl groups attached, conferring both stimulant and decongestant properties. As an organic base, ephedrine can accept protons due to the lone pair of electrons on its amino group.

Basicity and pKb of Ephedrine

The basicity of ephedrine is characterized by its pKb value, which is 4.64 in this context. The pKb value indicates the strength of the base—the lower the pKb, the stronger the base. It also relates to the pKa of its conjugate acid through the relationship:

$$ pKa + pKb = 14 $$

Thus, the pKa of ephedrine's conjugate acid can be calculated as:

$$ pKa = 14 - pKb = 14 - 4.64 = 9.36 $$

This pKa value reflects the acidity of the conjugate acid and provides a basis for understanding the equilibrium involved in the dissociation of ephedrine in solution.

Initial Data and Assumptions

To accurately determine the pH, we need to clarify the nature of the solution and the forms of ephedrine present:

- Amount of ephedrine: 0.1 mole
- Concentration of ephedrine: 0.05 M (moles per liter)
- pKb of ephedrine: 4.64

Given that the solution contains 0.05 M ephedrine, and since the total amount is 0.1 mole, the volume of the solution can be deduced:

$$ \text{Volume} = \frac{\text{moles}}{\text{concentration}} = \frac{0.1\, \text{mol}}{0.05\, \text{mol/L}} = 2\, \text{L} $$

This indicates that the solution volume is 2 liters.

Determining the pH: Step-by-Step Approach

To find the pH of this solution, we need to analyze the equilibrium involving the base form of ephedrine and its conjugate acid, considering that the solution is likely to be basic due to ephedrine's nature.

Step 1: Write the Relevant Equilibrium

The core equilibrium involves the protonation of ephedrine:

$$ \text{Ephedrine (B)} + H_2O \rightleftharpoons \text{Ephedrine (BH}^+)\ + OH^- $$

However, since we're dealing with a basic compound, it’s more straightforward to analyze its protonation using the conjugate acid-base equilibrium:

$$ \text{B} + H_2O \rightleftharpoons \text{BH}^+ + OH^- $$

The base dissociation constant (Kb) is related to pKb:

$$ K_b = 10^{-pK_b} = 10^{-4.64} \approx 2.29 \times 10^{-5} $$

Step 2: Calculate the Concentration of Ephedrine and Its Conjugate Acid

In the solution, the initial concentration of free ephedrine (B) is 0.05 M, and the equilibrium will involve some degree of protonation forming BH^+.

Let:

- x = concentration of BH^+ formed at equilibrium

Since the initial concentration of B is 0.05 M, and assuming that the change in concentration due to ionization is small (which is generally true for weak bases), the equilibrium concentrations are:

- $[B] = 0.05 - x$
- $[\text{BH}^+] = x$

Step 3: Write the Expression for Kb

Using the equilibrium concentrations:

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

Assuming $[OH^-] \approx x$:

$$K_b = \frac{x \times x}{0.05 - x} \approx \frac{x^2}{0.05}$$

Since K_b is small, x is much less than 0.05, so:

$$0.05 - x \approx 0.05$$

Therefore:

$$x^2 = K_b \times 0.05$$

$$x^2 = (2.29 \times 10^{-5}) \times 0.05$$

$$x^2 = 1.145 \times 10^{-6}$$

$$x = \sqrt{1.145 \times 10^{-6}} \approx 1.07 \times 10^{-3} \, \text{M}$$

This x is the concentration of $[OH^-]$.

Step 4: Calculate the pOH and pH

Now, calculate the pOH:

$$pOH = -\log [OH^-] = -\log (1.07 \times 10^{-3}) \approx 2.97$$

Finally, the pH is:

$$ pH = 14 - pOH = 14 - 2.97 = 11.03 $$

Interpreting the Results and Practical Considerations

The calculated pH of approximately 11.03 indicates a strongly basic solution, consistent with the behavior of ephedrine as a weak base. This pH value is significant because it demonstrates that even at a relatively low concentration (0.05 M), ephedrine's basic properties substantially raise the solution's pH.

Practical implications include:

- Ephedrine solutions are typically alkaline, which is important for its pharmacological activity.
- The pH influences solubility, stability, and bioavailability of ephedrine in formulations.
- Understanding the pH helps in designing proper storage conditions and handling procedures.

Additional Factors Affecting the pH Calculation

While the above calculation provides a good estimate, certain factors can influence the actual pH:

1. Ionic Strength and Activity Coefficients

At higher ionic strengths, activity coefficients deviate from unity, which can slightly alter the equilibrium concentrations.

2. Temperature Variations

The pKb and related constants are temperature-dependent; deviations from standard conditions can impact the pH.

3. Presence of Other Ions or Solutes

Additional ions or solutes in the solution can shift the equilibrium, affecting the pH.

4. Complete Dissociation vs. Partial

Since ephedrine is a weak base, it does not fully dissociate, and the approximation used here remains valid for most practical purposes.

Conclusion

Determining the pH of a solution containing 0.1 mole of ephedrine and a concentration of 0.05 M involves understanding its basic properties, the equilibrium between the base form and its conjugate acid, and applying the principles of acid-base chemistry. Through the calculated equilibrium, we find that the solution exhibits a pH of approximately 11.03, confirming its basic nature.

This detailed approach provides valuable insights for chemists, pharmacologists, and formulation scientists working with ephedrine or similar amines. Recognizing how pKb and concentration influence pH allows for better control and optimization of pharmaceutical and chemical processes involving ephedrine.

In summary:

- Ephedrine, with a pKb of 4.64, behaves as a weak base.
- Its pH in a 0.05 M solution is approximately 11.03.
- The calculation hinges on the equilibrium involving the conjugate acid-base pair and the base dissociation constant.

Understanding these principles ensures effective handling, application, and formulation of ephedrine-based solutions in various scientific and medical contexts.

Frequently Asked Questions

What is the pH of a solution containing 0.1 mole of ephedrine with a pKb of 4.64 and 0.05 M ephedrine?

To determine the pH, first calculate the Kb from pKb: Kb = 10^(-4.64) ≈ 2.29 × 10^(-5). Using the initial concentration of ephedrine (0.05 M), set up the equilibrium expression for the base dissociation. Assuming x is the amount that dissociates: x^2 / (0.05 - x) ≈ Kb. Since Kb is small, approximate 0.05 - x ≈ 0.05. Then, x ≈ √(Kb × 0.05) ≈ √(2.29×10^(-5) × 0.05) ≈ √(1.145×10^(-6)) ≈ 0.00107 M. The concentration of OH^- is approximately 0.00107 M, so pOH ≈ 2.97. Therefore, pH ≈ 14 - 2.97 ≈ 11.03. The solution is mildly basic with a pH around 11.

How does the pKb value of 4.64 influence the basicity of ephedrine in solution?

A pKb of 4.64 indicates that ephedrine is a relatively weak base. The lower the pKb, the stronger the base. Since 4.64 is moderate, ephedrine exhibits mild to moderate basicity, contributing to a pH above 7 in solution, typically around 11 as calculated.

What is the significance of the 0.05 M concentration of ephedrine in calculating the solution's pH?

The 0.05 M concentration provides the initial amount of ephedrine in solution, which is essential for setting up the equilibrium expression. It determines the extent of dissociation and thus influences the concentration of OH□ ions, ultimately affecting the pH calculation.

If the solution contains 0.1 mole of ephedrine, how does this compare to the 0.05 M concentration for pH calculation?

The 0.1 mole of ephedrine and the 0.05 M concentration are related quantities; if the solution volume is known, they can be converted into each other. In the context of pH calculation, the 0.05 M concentration represents the initial molarity, which is used directly in equilibrium expressions. The 0.1 mole indicates the total amount present, assuming a specific volume, but for pH calculations, molarity is the key parameter.

Can we assume complete dissociation of ephedrine in solution given the pKb of 4.64?

No, with a pKb of 4.64 indicating a weak base, complete dissociation is unlikely. Instead, equilibrium dissociation occurs, and only a small fraction of ephedrine molecules dissociate, which is accounted for in pH calculations using the equilibrium approach.

How would increasing the amount of ephedrine affect the pH of the solution?

Increasing the amount of ephedrine increases the base concentration, leading to greater dissociation and higher OH□ ion concentration. Consequently, the pH would increase, making the solution more basic.

What assumptions are made in calculating the pH of this solution?

Key assumptions include: (1) The dissociation of ephedrine is small enough to approximate (0.05 - x) as 0.05; (2) The solution volume remains constant; (3) Temperature is standard (usually 25°C); and (4) No other acids or bases are present to influence the pH significantly.

Additional Resources

What Is The pH Of A Solution Containing 0.1 Mole Of Ephedrine (pKb = 4.64) And 0.05 Moles Of Ephedrine

Understanding the pH of a solution that contains a specific amount of ephedrine, especially when considering its basic properties such as pKb, is essential in fields ranging from pharmaceutical formulation to chemical analysis. In this article, we will explore what is the pH of a solution containing 0.1 mole of ephedrine (pKb = 4.64) and 0.05 Moles of ephedrine, breaking down the concepts, calculations, and implications involved in determining the acidity or basicity of such a solution. This comprehensive guide aims to clarify the process for students, researchers, or anyone interested in chemical solutions' pH dynamics.

Introduction to Ephedrine and Its Basic Properties

Ephedrine is a sympathomimetic amine commonly used in medications for asthma, nasal congestion, and weight loss. Chemically, it functions as a weak base, which can accept protons in aqueous solutions. Its basic nature is quantified through its basicity constants—namely pKb, which describes its affinity for protons.

In the context of solution chemistry:

- pKb is the negative base-10 logarithm of the base dissociation constant (Kb).
- The higher the pKb, the weaker the base; conversely, a lower pKb indicates a stronger base.

Given:
- pKb of ephedrine = 4.64
- Amount of ephedrine in the solution = 0.1 mole and 0.05 Moles (we'll clarify how these relate)

Clarifying the Scenario: Concentrations and Moles

Before proceeding to calculations, it’s crucial to understand the specifics:

- 0.1 Mole of ephedrine: This specifies the total number of moles present in the solution.
- 0.05 Moles of ephedrine: Usually, molarity (M) relates to moles per liter. If the solution volume is known, these quantities can be related.

Assuming the solution volume is 1 liter:
- The solution contains 0.1 mol of ephedrine in total.
- The concentration of ephedrine is 0.05 M, which implies the solution volume is 2 liters (since 0.05 mol in 2 liters equals 0.025 M, or perhaps the problem refers to two different solutions or conditions).

Alternatively, if the solution volume isn’t specified, we will focus on relative concentrations or assume standard conditions.

For simplicity, and to facilitate the calculations, we'll assume:
- The total volume of the solution is 1 liter.
- The initial concentration of ephedrine is 0.05 M (since molarity is moles per liter).

Thus, the 0.1 mole of ephedrine corresponds to 0.1 mol / 1 L = 0.1 M concentration.

However, because both quantities are given, it suggests a mixed scenario—potentially, a solution initially containing 0.1 mol total, with a concentration of 0.05 M, which indicates the solution volume is 2 liters.

Understanding Ephedrine as a Weak Base

Ephedrine's basicity can be characterized by its pKb = 4.64. To find the pH, we need to understand how much of the ephedrine exists in its protonated and unprotonated forms in solution.

The equilibrium involved:

$$ \text{Ephedrine (B)} + H_2O \rightleftharpoons \text{Ephedrine (BH}^+ \text{)} + OH^- $$

Or considering its conjugate acid:

$$ \text{B} + H_2O \rightleftharpoons \text{BH}^+ + OH^- $$

The relevant dissociation constant:

$$ K_b = 10^{-\text{p}K_b} = 10^{-4.64} \approx 2.29 \times 10^{-5} $$

Step-by-Step Calculation of pH

1. Calculating the Base Dissociation Constant (Kb)

Given:

- pKb = 4.64
- $K_b = 10^{-pK_b} \approx 2.29 \times 10^{-5}$

2. Determine the Initial Concentration of Ephedrine

Assuming:

- Total volume = 1 liter
- Initial concentration of ephedrine (B) = 0.05 M

3. Setting Up the Equilibrium Expression

Let:

- $[B]$ = concentration of unprotonated ephedrine at equilibrium
- $[BH^+]$ = concentration of protonated ephedrine at equilibrium
- x = amount of ephedrine that dissociates to form OH^-

Since ephedrine is a weak base:

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

Initially:

$$ [B]_0 = 0.05\, M $$

At equilibrium:

$$ [B] = 0.05 - x $$

$$ [BH^+] = x $$

$$ [OH^-] = x $$

Substituting into the Kb expression:

$$ 2.29 \times 10^{-5} = \frac{x \times x}{0.05 - x} $$

Assuming $x \ll 0.05$, which is valid for weak bases:

$$ 2.29 \times 10^{-5} \approx \frac{x^2}{0.05} $$

$$ x^2 = 2.29 \times 10^{-5} \times 0.05 $$

$$ x^2 \approx 1.145 \times 10^{-6} $$

$$ x \approx \sqrt{1.145 \times 10^{-6}} $$

$$ x \approx 1.07 \times 10^{-3} \, M $$

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

4. Calculate pOH and pH

- $\text{pOH} = -\log [OH^-]$
- $\text{pOH} = -\log (1.07 \times 10^{-3}) \approx 2.97$

Then:

$$\text{pH} = 14 - \text{pOH} = 14 - 2.97 = 11.03$$

Final pH of the Solution

Based on the assumptions:
- Initial concentration of ephedrine = 0.05 M
- pKb = 4.64

The estimated pH of the solution is approximately 11.03, indicating a basic solution due to the weak base nature of ephedrine.

Additional Considerations and Nuances

Effect of Total Moles and Concentration Variations

- If the total moles or volume change, the concentration of ephedrine will change accordingly, impacting the degree of ionization and thus the pH.
- In more concentrated solutions, the degree of ionization decreases slightly, but for weak bases, the pH remains in the basic range.

Protonation and Acid-Base Equilibria

- If the solution contains acids or other ions, they can influence the equilibrium.
- The pKb only describes the basicity; in real solutions, the pH will also depend on other factors.

Practical Implications

- Pharmacological formulations involving ephedrine must consider the pH to ensure stability and bioavailability.
- Understanding the pH helps in predicting the ionization state, which affects absorption and interaction.

Summary and Key Takeaways

- The pH of a solution containing ephedrine can be estimated using its pKb value.
- For 0.05 M ephedrine (assuming volume 1 L), the pH is approximately 11.03, indicating a basic solution.
- The calculation involves:
- Converting pKb to Kb
- Setting up an equilibrium expression
- Assuming negligible x relative to initial concentration
- Calculating $[OH^-]$, then pOH and pH

Understanding these principles enables chemists and pharmacists to predict solution behavior, optimize formulations, and interpret experimental data effectively.

Final Thoughts

Accurately determining the pH of a solution containing weak bases like ephedrine requires a clear understanding of acid-base equilibria, concentration effects, and the relevant equilibrium constants. By systematically applying these concepts, scientists can predict and control the pH environment, which is vital for both chemical research and practical applications.

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