A Buffer Solution Is 0.470 M In A Weak Acid (a=1.8106) And 0.210 M In Its Conjugate Base, What Is The

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Understanding buffer solutions is fundamental in chemistry, especially when controlling pH in various chemical and biological systems. When dealing with a buffer solution that contains a weak acid and its conjugate base, calculating the pH becomes a straightforward application of the Henderson-Hasselbalch equation. In this article, we will explore how to determine the pH of a buffer solution with given concentrations and acid dissociation constant (a), specifically focusing on a solution with 0.470 M weak acid and 0.210 M conjugate base, where the acid dissociation constant (Ka) is 1.8106.

Understanding Buffer Solutions

What Is a Buffer Solution?

A buffer solution is a mixture of a weak acid and its conjugate base (or vice versa) that resists changes in pH when small amounts of acids or bases are added. Buffers are vital in many biological systems and chemical processes because they maintain a stable pH environment.

Key characteristics of buffer solutions include:

- Composed of a weak acid and its conjugate base or vice versa.
- Capable of neutralizing added acids or bases.
- Have a specific pH range where they are most effective.

Role of Weak Acids and Conjugate Bases in Buffers

Weak acids and their conjugate bases work together to stabilize the pH. When an acid or base is added, the conjugate pair reacts to minimize pH fluctuations:

- Addition of acid: conjugate base reacts with H^+ ions.
- Addition of base: weak acid reacts with OH^- ions.

This dynamic equilibrium is crucial for maintaining pH stability.

Calculating pH of a Buffer Solution

Henderson-Hasselbalch Equation

The primary tool used to calculate the pH of a buffer solution is the Henderson-Hasselbalch equation:

$$ \boxed{\text{pH} = \text{p}K_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]} \right)} $$

Where:
- $[\text{A}^-]$ = concentration of conjugate base
- $[\text{HA}]$ = concentration of weak acid
- $\text{p}K_a$ = negative logarithm of the acid dissociation constant K_a

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Given Data and Parameters

Let's review the data provided:

- Concentration of weak acid, [HA]: 0.470 M
- Concentration of conjugate base, [A^-]: 0.210 M
- Acid dissociation constant, K_a: 1.8106

Note: Since $a = 1.8106$, it is likely referring to $\log_{10} K_a$. To clarify, in most chemistry contexts, a could represent the $\log_{10} K_a$, so:

$$ \text{p}K_a = a = 1.8106 $$

Thus,

$$ K_a = 10^{-1.8106} $$

Calculating the pKa of the Weak Acid

Using the provided a value:

$$ \text{p}K_a = 1.8106 $$

And calculating K_a:

$$ K_a = 10^{-\text{p}K_a} = 10^{-1.8106} $$

Using a calculator:

$$ K_a \approx 10^{-1.8106} \approx 0.0155 $$

This confirms that the acid's dissociation constant is approximately 0.0155.

Calculating the pH of the Buffer Solution

Applying the Henderson-Hasselbalch equation:

$$ \text{pH} = \text{p}K_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]} \right) $$

Substitute the known values:

$$ \text{pH} = 1.8106 + \log \left(\frac{0.210}{0.470} \right) $$

Calculate the ratio:

$$ \frac{0.210}{0.470} \approx 0.4468 $$

Calculate the logarithm:

$$ \log(0.4468) \approx -0.349 $$

Finally, compute the pH:

$$ \text{pH} = 1.8106 - 0.349 = 1.4616 $$

Final pH of the Buffer Solution

Based on the calculations, the pH of the buffer solution is approximately:

pH ≈ 1.46

This indicates that the buffer is quite acidic, consistent with the relatively high K_a value of the weak acid.

Implications and Applications of Buffer pH Calculations

Importance of Accurate pH Determination

Knowing the precise pH of a buffer solution is critical in various fields:

- Biology: Enzymatic activity depends heavily on pH.
- Medicine: Blood and other bodily fluids are maintained within narrow pH ranges.
- Chemistry: Reaction rates and equilibrium positions are pH-dependent.
- Environmental science: Buffer solutions help in studying and controlling environmental pH.

Factors Influencing Buffer Capacity

Buffer capacity refers to the ability of a buffer to resist pH change. It is influenced by:

- The concentrations of weak acid and conjugate base.
- The pKa relative to the target pH.
- The ratio of conjugate base to acid.

Maintaining optimal concentrations ensures effective buffering over the desired pH range.

Additional Considerations in Buffer Calculations

Effect of Adding Strong Acids or Bases

Adding strong acids or bases can shift the equilibrium, but a well-designed buffer minimizes pH changes. For example, adding HCl (a strong acid) will be neutralized by the conjugate base, while NaOH (a strong base) will react with the weak acid.

Buffer Range

The most effective pH range of a buffer is typically within 1 pH unit of its pKa. Since the pKa here is approximately 1.81, this buffer is most effective between pH 0.81 and 2.81.

Summary and Key Takeaways

- The pH of a buffer solution containing a weak acid and its conjugate base can be accurately calculated using the Henderson-Hasselbalch equation.
- Given concentrations and $a = 1.8106$ (interpreted as $\text{p}K_a$), the pH of the solution is approximately 1.46.
- Understanding buffer capacity and pH stability is essential in biological, chemical, and environmental applications.
- Precise calculations enable the design of buffers suitable for specific pH ranges, critical for experimental accuracy and system stability.

Conclusion

In conclusion, determining the pH of a buffer solution with known concentrations of weak acid and conjugate base involves understanding the fundamental principles of acid-base chemistry and correctly applying the Henderson-Hasselbalch equation. With a weak acid concentration of 0.470 M, a conjugate base concentration of 0.210 M, and a $\text{p}K_a$ of approximately 1.81, the pH is calculated to be around 1.46. This process underscores the importance of precise data and calculations in chemical analysis and practical applications, enabling scientists and engineers to design effective buffer systems for diverse purposes.

Keywords: buffer solution, pH calculation, weak acid, conjugate base, Henderson-Hasselbalch, pKa, acid dissociation constant, chemical equilibrium, pH stability, buffer capacity

Frequently Asked Questions

What is the pH of a buffer solution that is 0.470 M in a weak acid with a Ka of 1.81×10^{-6} and 0.210 M in its conjugate base?

Using the Henderson-Hasselbalch equation: $pH = pKa + \log([A^-]/[HA])$. First, find pKa: $pKa = -\log(Ka) \approx 5.74$. Then, $pH = 5.74 + \log(0.210/0.470) \approx 5.74 + \log(0.447) \approx 5.74 - 0.350 \approx 5.39$.

How do you calculate the pKa of a weak acid with a given Ka value?

pKa is calculated as $pKa = -\log(Ka)$. For $Ka = 1.81\times10^{-6}$, $pKa \approx 5.74$.

What is the significance of the ratio of conjugate base to acid in a buffer solution?

The ratio determines the pH of the buffer; a higher conjugate base concentration relative to the acid raises the pH, while a higher acid concentration lowers it, according to the Henderson-Hasselbalch equation.

How does increasing the concentration of the conjugate base affect the pH of the buffer?

Increasing the conjugate base concentration increases the ratio $[A^-]/[HA]$, resulting in a higher pH of the buffer solution.

Can you explain the Henderson-Hasselbalch equation in the context of this buffer?

Yes. The equation $pH = pKa + \log([A^-]/[HA])$ relates the pH to the pKa of the weak acid and the ratio of conjugate base to acid, helping to determine the buffer's pH based on their concentrations.

What is the primary purpose of a buffer solution in chemical systems?

A buffer solution resists changes in pH upon addition of small amounts of acid or base, maintaining a stable pH environment.

How would the pH change if the concentration of the weak acid increased while the conjugate base remained constant?

Increasing the acid concentration while keeping the conjugate base constant would decrease the ratio $[A^-]/[HA]$, leading to a lower pH.

What are common applications of buffer solutions with weak acids and their conjugate bases?

Buffer solutions are used in biological systems (like blood), chemical manufacturing, pharmaceuticals, and laboratory analysis to maintain consistent pH conditions.

Additional Resources

A Buffer Solution Is 0.470 M In A Weak Acid (a=1.8106) And 0.210 M In Its Conjugate Base, What Is The

In the realm of chemistry, understanding buffer solutions is fundamental to numerous scientific and industrial applications. From maintaining the pH of biological fluids to controlling reactions in chemical manufacturing, buffers serve as vital tools for managing acidity and alkalinity. Today, we delve into a specific scenario: a buffer solution comprising a weak acid and its conjugate base, characterized by known concentrations and a given acid dissociation constant. Our goal is to determine the pH of this buffer, a task that combines theoretical principles with practical calculations.

This detailed exploration will guide readers through the concepts of buffer systems, the significance of the acid dissociation constant (Ka), and the step-by-step calculation process to find the pH. Whether you're a student seeking clarity or a professional brushing up on foundational chemistry, this article aims to present complex ideas in a clear, accessible manner.

Understanding Buffer Solutions: The Basics

What Is a Buffer Solution?

A buffer solution is a mixture that resists changes in pH when small amounts of acid or base are added. This stability arises from the presence of a weak acid and its conjugate base (or a weak base and its conjugate acid). The weak acid can neutralize added bases, while the conjugate base can neutralize added acids, thereby maintaining a relatively constant pH.

For example, a typical buffer might contain acetic acid (a weak acid) and sodium acetate (its conjugate base). When small amounts of H^+ or OH^- ions are introduced, the buffer components react to neutralize these ions, preventing significant shifts in pH.

The Role of the Acid Dissociation Constant (Ka)

The strength of a weak acid is quantified by its acid dissociation constant, Ka. This value indicates how much the acid dissociates in water:

$$\mathrm{HA} \rightleftharpoons \mathrm{H^+} + \mathrm{A^-}$$

The expression for Ka is:

$$K_a = \frac{[\mathrm{H^+}][\mathrm{A^-}]}{[\mathrm{HA}]}$$

A smaller Ka means a weaker acid, dissociating less in solution, while a larger Ka indicates a stronger weak acid.

In our scenario, the weak acid has a Ka value derived from the provided a=1.8106, which is the activity coefficient. However, for dilute solutions, activities approximate concentrations, and we often use the pKa (the negative logarithm of Ka) for simplicity.

Given Data and Objective

The problem states:

- The concentration of the weak acid (HA) is 0.470 M.
- The concentration of its conjugate base (A^-) is 0.210 M.
- The activity coefficient (a) is 1.8106, but in dilute solutions, activities are approximated by concentrations for simplicity.
- The goal: To find the pH of this buffer solution.

Calculating the Acid Dissociation Constant (Ka)

The value of a=1.8106 suggests that activity coefficients are considered, but for practical calculations, especially in dilute aqueous solutions, we often approximate activities as concentrations, especially when the activity coefficient is close to 1.8106, indicating some ionic interactions or non-ideal behavior.

Assuming the provided 'a' refers to activity coefficient, we need to clarify whether the given value directly relates to Ka or if it’s a separate parameter. For simplicity, and common practice, if the activity coefficient is given and the solution is dilute, we can approximate:

$$ \text{Activities} \approx \text{Concentrations} \times \text{activity coefficient} $$

But in many cases, especially in basic buffer calculations, the activity coefficient is neglected unless specified otherwise.

For the purpose of this calculation, let's assume the activity coefficient pertains to the weak acid and base activities, but since the problem asks for pH, and the concentrations are provided, we proceed to calculate Ka assuming the activity coefficient's effect is minimal or that concentrations suffice.

If more precise calculations are needed, we could adjust concentrations by dividing by the activity coefficient.

Determining pKa from the Activity Coefficient

Given the activity coefficient (a), the effective concentration can be approximated as:

$$ [\mathrm{HA}]_{activity} \approx \frac{0.470}{1.8106} \approx 0.260 \, \text{M} $$

$$ [\mathrm{A^-}]_{activity} \approx \frac{0.210}{1.8106} \approx 0.116 \, \text{M} $$

Using these adjusted activities, the expression for Ka becomes:

$$ K_a = \frac{[\mathrm{H^+}][\mathrm{A^-}]}{[\mathrm{HA}]} $$

But since we're solving for pH, and the buffer equation relates pH, pKa, and the ratio of base to acid, it's more straightforward to use the Henderson-Hasselbalch equation.

Applying the Henderson-Hasselbalch Equation

The Henderson-Hasselbalch equation provides a direct way to calculate the pH of a buffer:

$$ \boxed{
\mathrm{pH} = \mathrm{p}K_a + \log \left(\frac{[\mathrm{A^-}]}{[\mathrm{HA}]} \right)
} $$

Given the concentrations:

- $[\mathrm{A^-}] = 0.210 \, \text{M}$
- $[\mathrm{HA}] = 0.470 \, \text{M}$

First, we need to find pKa.

Estimating pKa from the Given Data

Since the problem states the activity coefficient (a=1.8106), and assuming it impacts the activity, we can approximate the activity-adjusted concentrations:

$$ [\mathrm{HA}]_{act} \approx 0.470 / 1.8106 \approx 0.260 \, \text{M} $$

$$ [\mathrm{A^-}]_{act} \approx 0.210 / 1.8106 \approx 0.116 \, \text{M} $$

Now, the ratio needed for the Henderson-Hasselbalch equation is:

$$ \frac{[\mathrm{A^-}]}{[\mathrm{HA}]} = \frac{0.116}{0.260} \approx 0.446 $$

Taking the logarithm:

$$ \log(0.446) \approx -0.351 $$

Next, we need pKa. If the problem provided a value or a relation involving the activity coefficient and the acid strength, we could proceed. However, since no explicit Ka or pKa is provided, we might infer it from typical weak acids or through additional data.

If the initial activity coefficient or pKa is provided elsewhere, we could use it directly. For this illustration, let’s assume the weak acid’s pKa is around 4.74, a common value for weak acids like acetic acid, which often has a pKa close to this.

Calculating the pH of the Buffer

Using the assumed pKa of 4.74:

$$\mathrm{pH} = 4.74 + (-0.351) = 4.389$$

Thus, the approximate pH of the buffer is around 4.39.

Implications and Practical Considerations

This calculation demonstrates how buffer solutions maintain their pH through the balance of weak acid and conjugate base concentrations. It also underscores the importance of activity coefficients in real-world solutions, where ionic interactions and non-ideal behavior can affect calculations.

In industrial or biological contexts, knowing the pH allows scientists to predict how the buffer will perform under different conditions, ensuring system stability. For example, in pharmaceuticals, maintaining a stable pH is crucial for drug efficacy, while in agriculture, buffers help control soil acidity.

Conclusion: The Significance of Buffer Calculations

Determining the pH of a buffer solution from known concentrations involves understanding the interplay between acid-base chemistry, activity coefficients, and equilibrium principles. While simplified models often assume ideal behavior, real solutions require adjustments accounting for ionic interactions.

In this case, by applying the Henderson-Hasselbalch equation and considering the activity coefficient, we estimated the pH to be approximately 4.39. This demonstrates how fundamental chemical principles translate into practical calculations, enabling scientists to design and control systems with precision.

Understanding buffer systems not only deepens our grasp of chemical equilibria but also highlights their central role across scientific disciplines, from medicine to environmental science. As we continue to explore the chemistry of solutions, these foundational concepts remain essential tools in the chemist's toolkit, ensuring that the delicate balance of acidity and alkalinity is maintained across countless applications.

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