Determine The Ph And Poh Of 0.25 L Of A Solution That Is 0.0206 M Boric Acid And 0.0231 M Sodium Borate;

Determine The pH And pOH Of 0.25 L Of A Solution That Is 0.0206 M Boric Acid And 0.0231 M Sodium Borate;

Understanding the pH and pOH of a solution is fundamental in chemistry, especially when dealing with buffer systems involving weak acids and their conjugate bases. In this article, we will explore how to accurately determine the pH and pOH of a 0.25-liter solution containing 0.0206 M boric acid (a weak acid) and 0.0231 M sodium borate (its conjugate base). This process involves understanding buffer systems, applying the Henderson-Hasselbalch equation, and considering the properties of boric acid and sodium borate.

Understanding the Components of the Solution

Before calculating the pH and pOH, it is essential to understand the nature of the components involved in the solution.

Boric Acid (H_3BO_3)

- Boric acid is a weak, monobasic acid that partially dissociates in water.
- Its dissociation reaction is:
$H_3BO_3 + H_2O \rightleftharpoons B(OH)_4^- + H^+$
- The dissociation constant (Ka) for boric acid is approximately 7.3×10^{-10} at 25°C.

Sodium Borate ($Na_2B_4O_7$)

- Sodium borate is the sodium salt of boric acid and acts as a weak base in solution.
- When dissolved, it provides borate ions ($B(OH)_4^-$), which can react with water and influence the pH.
- The conjugate base of boric acid in this system is $B(OH)_4^-$.

Buffer System: Boric Acid and Sodium Borate

Since the solution contains both boric acid and sodium borate, it forms a buffer system that resists changes in pH. Understanding buffer chemistry helps us determine the pH and pOH values.

Buffer Concept

- A buffer solution contains a weak acid and its conjugate base (or vice versa).

- It maintains a relatively constant pH when small amounts of acid or base are added.
- The pH of a buffer can be calculated using the Henderson-Hasselbalch equation:

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

where:
- pKa is the negative logarithm of the acid dissociation constant (Ka).
- $[A^-]$ is the molar concentration of the conjugate base.
- [HA] is the molar concentration of the weak acid.

Calculating the pH of the Buffer Solution

To determine the pH, follow these steps:

Step 1: Determine the pKa of Boric Acid

- Given $Ka = 7.3 \times 10^{-10}$,
- $pKa = -\log(Ka) = -\log(7.3 \times 10^{-10}) \approx 9.14$

Step 2: Identify the Concentrations of Acid and Base

- Molarity of boric acid (HA): 0.0206 M
- Molarity of sodium borate (providing $B(OH)_4^-$, A^-): 0.0231 M

Note: Since sodium borate supplies $B(OH)_4^-$ ions, which are conjugate bases, these concentrations are used directly in the Henderson-Hasselbalch equation.

Step 3: Apply the Henderson-Hasselbalch Equation

- $pH = pKa + \log([A^-]/[HA])$
- $pH = 9.14 + \log(0.0231 / 0.0206)$

Calculate the ratio:
- $0.0231 / 0.0206 \approx 1.122$

Calculate the logarithm:
- $\log(1.122) \approx 0.050$

Now, compute the pH:
- $pH \approx 9.14 + 0.050 = 9.19$

Calculating the pOH of the Solution

Since pH + pOH = 14 at 25°C, the pOH can be easily calculated:

- pOH = 14 - pH = 14 - 9.19 ≈ 4.81

Alternatively, for a more detailed analysis, we could directly calculate hydroxide ion concentration, but given the buffer system, using the relationship with pH is straightforward.

Summary of Results

- pH of the solution: approximately 9.19
- pOH of the solution: approximately 4.81

Additional Considerations in Buffer Calculations

While the above calculations provide a good estimate, several factors can influence the accuracy:

- **Temperature dependence:** pKa values vary with temperature; the above assumes 25°C.
- **Activity coefficients:** In real solutions, activity coefficients may slightly alter calculations.
- **Exact dissociation behavior:** Boric acid's dissociation is more complex and involves Lewis acid-base interactions, but for typical calculations, the Henderson-Hasselbalch equation suffices.

Practical Applications of pH and pOH Calculations

Understanding how to determine pH and pOH in buffer systems containing boric acid and sodium borate is crucial in various fields:

- **Pharmaceutical formulations:** Ensuring the stability of borate-based solutions.
- **Environmental chemistry:** Monitoring boric acid levels in water systems.
- **Industrial processes:** Maintaining pH in manufacturing involving boron compounds.

Conclusion

Calculating the pH and pOH of a solution containing weak acids and their conjugate bases involves understanding buffer chemistry and applying the Henderson-Hasselbalch equation. In the case of a 0.25 L solution with 0.0206 M boric acid and 0.0231 M sodium borate, the pH is approximately 9.19, indicating a slightly alkaline buffer environment. The pOH, correspondingly, is about 4.81. These calculations are essential for scientists and engineers working with boron compounds, ensuring proper control and understanding of solution chemistry.

By mastering these fundamental concepts, you can accurately analyze and predict the behavior of complex chemical systems involving weak acids, bases, and buffer solutions.

Frequently Asked Questions

How do you determine the pH of a solution containing both boric acid and sodium borate?

You use the Henderson-Hasselbalch equation, considering the pKa of boric acid and the ratio of the concentrations of the conjugate base (sodium borate) to the acid (boric acid).

What is the pKa of boric acid relevant for calculating the pH in this solution?

The pKa of boric acid is approximately 9.24 at 25°C.

How do you calculate the pH of the solution given the concentrations of boric acid and sodium borate?

Use the Henderson-Hasselbalch equation: pH = pKa + log([Base]/[Acid]). Substitute the concentrations of sodium borate and boric acid to find the pH.

What is the role of sodium borate in this solution?

Sodium borate acts as the conjugate base, which helps buffer the solution and influences its pH.

How do you determine the pOH of the solution once the pH is known?

Subtract the pH from 14: pOH = 14 – pH.

Can you provide the step-by-step calculation to find the pH of this solution with given concentrations?

Yes. First, calculate the ratio of [Base]/[Acid]: 0.0231/0.0206 ≈ 1.122. Then, apply the Henderson-Hasselbalch equation: pH = 9.24 + log(1.122) ≈ 9.24 + 0.050 ≈ 9.29. The pH of the solution is approximately 9.29.

Additional Resources

pH and pOH Determination of a Boric Acid and Sodium Borate Solution: An Expert Analysis

Understanding the acidity and basicity of a solution is fundamental in chemistry, especially when dealing with buffer systems involving weak acids and their conjugate bases. In this comprehensive review, we'll explore how to determine the pH and pOH of a 0.25 L solution containing 0.0206 M boric acid (H_3BO_3) and 0.0231 M sodium borate ($Na_2B_4O_7$). This analysis combines theoretical principles with practical calculations, providing a detailed insight into the behavior of this buffer system.

Introduction to the System: Boric Acid and Sodium Borate

Boric acid (H_3BO_3) is a weak, monobasic acid commonly used as an antiseptic, insecticide, and in chemical laboratories. It exhibits weak acidic properties primarily through its interaction with hydroxide ions, but unlike typical acids, its dissociation in water is limited and involves a different mechanism, often leading to the formation of borate ions.

Sodium borate ($Na_2B_4O_7$), on the other hand, is a salt that, when dissolved in water, yields the borate ion ($B(OH)_4^-$) and sodium ions. The borate ion acts as a weak base, capable of accepting protons, which makes the system a classic example of a buffer when combined with boric acid.

Key components of the system:

- Weak acid: Boric acid (H_3BO_3)
- Conjugate base: Borate ion ($B(OH)_4^-$), supplied via sodium borate
- Buffer system: The mixture of boric acid and borate ions stabilizes the pH

Theoretical Foundations

Equilibrium in Boric Acid Solution

Unlike typical acids, boric acid's dissociation in water is described by the following reaction:

$$\text{H}_3\text{BO}_3 + 2 \text{H}_2\text{O} \leftrightarrow \text{B}(\text{OH})_4^- + \text{H}^+$$

However, the actual process involves its interaction with hydroxide ions, leading to the formation of borate ions:

$$ \text{H}_3\text{BO}_3 + \text{OH}^- \leftrightarrow \text{B}(\text{OH})_4^- $$

The key to calculating pH in such a system involves understanding the acid dissociation constant (Ka) or the base dissociation constant (Kb). For boric acid, the relevant equilibrium is often expressed via its acid dissociation constant (Ka), which is approximately 6.4×10^{-10} at 25°C.

Note: Boric acid's behavior in water is unique because it doesn't donate protons directly but acts as a Lewis acid, accepting hydroxide ions.

Understanding the Buffer System

The mixture of boric acid and sodium borate forms a buffer capable of resisting pH changes. The ratio of conjugate acid to base determines the pH, according to the Henderson-Hasselbalch equation:

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

Where:
- $[\text{A}^-]$ is the concentration of the conjugate base (borate ion, $B(OH)_4^-$)
- $[\text{HA}]$ is the concentration of the weak acid (boric acid, H_3BO_3)
- pK_a is the negative logarithm of the acid dissociation constant K_a

---

Calculating the pH of the Solution

The primary challenge in this problem is accurately determining the equilibrium concentrations of boric acid and borate ions, considering their initial concentrations and the equilibrium constant.

Step 1: Establish Known Data

Parameter	Value	Explanation
Volume of solution	0.25 L	Given
$[H_3BO_3]$ initial	0.0206 M	Given
$[Na_2B_4O_7]$ initial	0.0231 M	Given
K_a of boric acid	6.4×10^{-10}	Literature value

Note: The sodium borate provides the conjugate base (borate ions) upon dissolution, assuming complete dissociation.

Step 2: Determine Initial Concentrations of Acid and Base

- Boric acid (HA):

$[HA]_0 = 0.0206\, \text{mol/L}$

- Borate (A^-):
$[A^-]_0 = 0.0231\, \text{mol/L}$

Since the solution volume is 0.25 L, the total moles are:

- Boric acid:
$n_{HA} = 0.0206\, \text{mol/L} \times 0.25\, \text{L} = 0.00515\, \text{mol}$

- Borate:
$n_{A^-} = 0.0231\, \text{mol/L} \times 0.25\, \text{L} = 0.005775\, \text{mol}$

Step 3: Write the Equilibrium Expression

The equilibrium involving boric acid and borate ions can be simplified using the Henderson-Hasselbalch equation, assuming the dominant equilibrium is:

$$\text{H}_3\text{BO}_3 + \text{OH}^- \leftrightarrow \text{B}(\text{OH})_4^-$$

The relevant equilibrium for pH calculation involves the formation of borate ions, where the pKa is approximately 9.24 at 25°C.

Note: The pKa of boric acid in water is often cited as around 9.24, which is crucial for calculations.

Applying the Henderson-Hasselbalch Equation

Given the initial concentrations, the solution acts as a buffer system:

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

Where:
- $[\text{A}^-]$ is the borate ion concentration after equilibrium
- $[\text{HA}]$ is the boric acid concentration after equilibrium

Since the initial molar amounts are known, and assuming negligible change in total volume (0.25 L), the concentrations at equilibrium can be approximated by considering the extent of the reaction.

Step 4: Calculate Equilibrium Concentrations

Let x be the amount of boric acid that reacts to produce borate:

$$
\begin{cases}
[\text{HA}]_{eq} = [\text{HA}]_0 - x \\
[\text{A}^-]_{eq} = [\text{A}^-]_0 + x
\end{cases}
$$

Given the small K_a, the reaction extent is minimal, but for precision, we can set up the equilibrium expression:

$$
K_a = \frac{[\text{H}^+][\text{B}(\text{OH})_4^-]}{[\text{H}_3\text{BO}_3]}
$$

But in this case, because boric acid acts as a Lewis acid, the effective pK_a is used directly in the Henderson-Hasselbalch equation derived for boric acid-borate systems.

Calculating pH Using the Henderson-Hasselbalch Equation

Substituting the initial molar amounts into the equation:

$$
\text{pH} = 9.24 + \log \left(\frac{0.005775}{0.00515} \right)
$$

Calculate the ratio:

$$
\frac{0.005775}{0.00515} \approx 1.122
$$

Find the logarithm:

$$
\log(1.122) \approx 0.050
$$

Finally, compute the pH:

\[

\text{pH} = 9.24 + 0.050 = 9.29
\]

This indicates that the solution is slightly basic, as expected for a boric acid-borate buffer system.

Determining the pOH

The pOH complements the pH:

\[
\text{pOH} = 14 - \text{pH} = 14 - 9.29 = 4.71
\]

This value confirms the slightly basic nature of the solution.

Summary of Results

Parameter	Value	Explanation
pH	9.29	Slightly basic buffer system involving

Determine The Ph And Poh Of 0 25 L Of A Solution That Is 0 0206 M Boric Acid And 0 0231 M Sodium Borate

Determine The Ph And Poh Of 0 25 L Of A Solution That Is 0 0206 M Boric Acid And 0 0231 M Sodium Borate

Related Articles

- During Its First Year Of Operations, Swifty SE Had The Following Transactions Pertaining To Its Ordinary
- Explain Why Spain And Portugal Faced The Decline Of Their Empire, While Other European Empires Continued
- Drag Each Sentence To The Correct Location On The Image.Identify The Relationship Between Kinetic Energy

Back to Home