28. What Is The Ph Of 0.10 M Kbro? Ka(hbro) = 2.0x109

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Understanding the pH of solutions, especially those involving weak acids and bases, is fundamental in chemistry. In this article, we will explore the calculation of the pH of a 0.10 M potassium bromide (KBrO), considering the acid dissociation constant (Ka) of hypobromous acid (HBrO), which is given as 2.0×109. We will analyze whether KBrO acts as a weak acid or base, determine the relevant equilibrium expressions, and ultimately calculate the pH of the solution.

Introduction to pH and Acid-Base Chemistry

The pH of a solution measures its acidity or alkalinity, defined as the negative logarithm of hydrogen ion concentration:

$pH = -\log [H^+]$

Understanding pH requires knowledge of the nature of solutes—whether they are acids, bases, or salts—and how they influence hydrogen ion concentration in aqueous solutions.

Acids, Bases, and Salts in Solution

- **Strong acids and bases**: Completely dissociate in water, providing a straightforward calculation of pH.
- **Weak acids and bases**: Partially dissociate, requiring equilibrium calculations involving Ka or Kb.
- **Salts**: Can hydrolyze in water affecting pH depending on the nature of the ions.

Understanding Potassium Bromide (KBr) and Its

Behavior in Solution

KBr is a salt composed of potassium ions (K^+) and bromide ions (Br^-). When dissolved in water:

- K^+ is a cation derived from a strong base (KOH) and does not affect pH significantly.
- Br^- is the conjugate base of a strong acid (HBr) and also does not affect pH significantly.

Therefore, KBr is generally considered a neutral salt, which means its aqueous solution typically has a pH near 7. However, the mention of hypobromous acid's Ka (2.0×10^{9}) suggests the problem involves acid-base considerations, perhaps related to the properties of hypobromous acid (HBrO) or bromide ions in equilibrium with it.

Clarifying the Role of HBrO and Its Ka

Hypobromous Acid (HBrO)

HBrO is a weak acid with a given acid dissociation constant (Ka) of 2.0×10^{9}. However, this value seems anomalous because Ka values are typically less than 1 for weak acids. A Ka greater than 1 indicates a very strong acid, which would essentially dissociate completely, behaving like a strong acid.

Note: It's likely that the given Ka is a typo or a misinterpretation. Usually, Ka values for weak acids are much less than 1 (e.g., 10^{-9}), and for strong acids, Ka is very large (approaching infinity). For the purposes of this problem, we will assume the given Ka is accurate and indicates a very strong acid, possibly meaning that HBrO is a strong acid in this context.

Potential Confusion and Clarification

- If HBrO is a strong acid (Ka > 1), then it dissociates completely, and the pH is determined directly by its initial concentration.
- If HBrO is a weak acid (Ka < 1), then we need to perform equilibrium calculations to find $[H^+]$.

Given the Ka = 2.0×10^9 (which is extremely high), the most reasonable assumption is that HBrO is strongly dissociating, effectively behaving as a strong acid.

Determining the pH of the Solution

Scenario 1: HBrO is a Strong Acid

- Since HBrO dissociates completely, the concentration of H+ ions in solution equals the initial concentration of HBrO, which is 0.10 M.

- Therefore, [H+] = 0.10 M.

- Calculating pH:

pH = -log(0.10) = 1.00

This indicates a strongly acidic solution with a pH of approximately 1.

Scenario 2: HBrO is a Weak Acid (for completeness)

- If we suppose that HBrO is weak, with Ka = 2.0×10-9 (assuming a typo in the problem), then the dissociation of HBrO can be represented as:

HBrO ⇌ H+ + BrO-

- Initial concentration of HBrO = 0.10 M.

- Equilibrium expression:

Ka = [H+][BrO-] / [HBrO]

- Let x = [H+] at equilibrium, assuming x is small compared to 0.10 M:

[H+] ≈ x

[BrO-] ≈ x

[HBrO] ≈ 0.10 - x ≈ 0.10 (since x is small)

- Therefore,

Ka ≈ x2 / 0.10

x2 = Ka × 0.10

x2 = 2.0×10-9 × 0.10 = 2.0×10-10

x = √(2.0×10-10) ≈ 1.41×10-5 M

- pH = -log(1.41×10-5) ≈ 4.85

This pH indicates a weakly acidic solution.

Interpreting the Results and Final Answer

Given the original data (Ka = 2.0×109), it strongly suggests that HBrO is acting as a strong acid in this scenario. Therefore, the pH of the 0.10 M solution is approximately 1.00.

Final conclusion:

The pH of 0.10 M KBrO, given the Ka of hypobromous acid as 2.0×10^{9}, is approximately 1.00, indicating a highly acidic solution.

Additional Considerations and Related Topics

1. Understanding Acid Dissociation Constants (Ka)

- Ka indicates the strength of an acid; larger values mean stronger acids.
- Typical Ka for weak acids ranges from 10^{-2} to 10^{-10}.
- Values greater than 1 imply complete dissociation, akin to strong acids.

2. The Role of Salts in pH Regulation

- Salts like KBrO are generally neutral, but their constituent ions can influence pH if they originate from weak acids or bases.
- Hydrolysis of ions can lead to pH shifts, but in the case of K^+ and Br^-, the solution remains neutral unless other factors are involved.

3. Common Mistakes in pH Calculations

- Misinterpreting the Ka value—ensuring it's within a reasonable range.
- Forgetting to consider the dissociation extent or assuming complete dissociation when inappropriate.
- Neglecting the initial concentration of the acid or base.

Summary

This comprehensive explanation has illustrated how to approach calculating the pH of a solution involving hypobromous acid with a given Ka, considering scenarios where it acts as a strong or weak acid. Given the unusual Ka value, the solution behaves as strongly acidic, resulting in a pH of approximately 1.00 for the 0.10 M KBrO solution.

Understanding these concepts not only aids in solving similar problems but also deepens your overall grasp of acid-base chemistry, equilibrium calculations, and how to interpret dissociation constants effectively.

Frequently Asked Questions

What is the pH of a 0.10 M KBr solution?

Since KBr is a neutral salt derived from a strong acid (HBr) and a strong base (KOH), its aqueous solution has a pH close to 7.

Does KBr affect the pH of its solution?

No, KBr does not affect the pH significantly because it is a neutral salt formed from strong acid and base, resulting in a neutral solution.

What role does the Ka of HBr play in determining the pH of KBr solution?

The Ka of HBr indicates its strength as a strong acid; since KBr is formed from HBr and KOH, the strong acid and base neutralize each other, so Ka of HBr does not directly influence the pH of KBr solution.

Is the pH of 0.10 M KBr solution affected by the Ka of HBr?

No, because HBr is a strong acid with a very high Ka, it dissociates completely, and the resulting KBr solution remains neutral, with a pH near 7.

How do you calculate the pH of a neutral salt solution like KBr?

For a neutral salt like KBr, the pH is approximately 7 because it originates from a strong acid and strong base, which fully dissociate without affecting the pH.

What is the significance of the given Ka value of 2.0x10^9 for HBr?

The Ka value of 2.0x10^9 indicates that HBr is a very strong acid, dissociating completely in solution, which is relevant for understanding its salts but does not alter the neutrality of KBr solutions.

Can the Ka value of HBr influence the pH of KBr solution directly?

No, because KBr is a neutral salt, and the strong dissociation of HBr means the solution's pH is unaffected by the Ka in terms of acidity or alkalinity.

What is the typical pH of a 0.10 M KBr solution?

The pH of a 0.10 M KBr solution is approximately 7, indicating a neutral solution.

Why is KBr considered a neutral salt in aqueous solutions?

KBr is formed from a strong acid (HBr) and a strong base (KOH), which fully dissociate and do not affect the pH, making it neutral in solution.

How does the concentration of KBr affect its pH?

Since KBr is a neutral salt, its pH remains close to 7 regardless of concentration, assuming no other interfering substances are present.

Additional Resources

What Is The pH Of 0.10 M KBro? Ka(HBro) = 2.0×10^9

Understanding the pH of a solution involves delving into the fundamental principles of acid-base chemistry, specifically the behavior of weak acids and their dissociation in aqueous solutions. When analyzing a solution such as 0.10 M potassium hydrobromide (KBro), which is the salt of a weak acid (HBro) and a strong base (KOH), it becomes essential to examine the acid dissociation constant (Ka) and how it influences the solution's acidity. In this comprehensive review, we will explore the chemical nature of KBro, interpret the significance of its Ka value, and methodically calculate its pH, providing insights into the underlying chemical equilibria.

Understanding the Chemical Context: KBro and Its Acid-Base Nature

The Composition of KBro

Potassium hydrobromide, chemically denoted as KBro, is an ionic compound formed from potassium ions (K^+) and hydrobromide ions ($HBro^-$). When dissolved in water, KBro dissociates completely into these ions:

KBro (s) → K^+ (aq) + $HBro^-$ (aq)

Since potassium salts are generally considered neutral in aqueous solutions—due to the weak conjugate base nature of $HBro^-$—their influence on pH depends primarily on the properties of the associated acid or base.

The Acidic Character of HBro

The anion $HBro^-$ is the conjugate base of the weak acid HBrO (hypobromous acid). Given the provided dissociation constant (Ka) for HBro:

Ka(HBro) = 2.0×10^9

which appears exceptionally high, indicating the acid is extremely strong. But this raises an immediate point of clarification, as typical Ka values for weak acids are less than 1, and strong acids have dissociation constants effectively approaching infinity (or very large).

The value 2.0×10^{9} suggests a potential typographical error or misinterpretation.

Likely Correction or Clarification:
- If the value was meant to be $Ka = 2.0\times10^{-9}$, it would correspond to a weak acid.
- If it is truly $Ka = 2.0\times10^{9}$, it would imply an acid stronger than 100% dissociation, which is chemically impossible under normal conditions.

Assuming the typical scenario and the context of the question, the more plausible and scientifically consistent value is:

$Ka(HBro) = 2.0\times10^{-9}$

This value indicates a weak acid, which is consistent with the properties of hypobromous acid (HBrO). Therefore, for the purposes of this analysis, we'll proceed with $Ka = 2.0\times10^{-9}$.

Significance of the Acid Dissociation Constant (Ka)

What Does Ka Represent?

Ka quantifies the strength of an acid in aqueous solution. It is derived from the equilibrium expression:

$HBro\ (aq) + H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + Bro^-\ (aq)$

and is expressed as:

$Ka = [H_3O^+][Bro^-] / [HBro]$

A small Ka indicates a weak acid that does not dissociate significantly, whereas a large Ka (close to or exceeding 1) would indicate a very strong acid.

Relation to pKa and pH

The pKa is simply:

$pKa = -\log(Ka)$

Using the assumed Ka:

$pKa = -\log(2.0\times10^{-9}) \approx 8.70$

A pKa of approximately 8.70 suggests the acid is quite weak, and the solution's pH will predominantly depend on the extent of dissociation of HBro in water.

Calculating the pH of 0.10 M KBro Solution

Understanding the Contribution to pH

Since KBro is a salt derived from a weak acid (HBro), its aqueous solution can exhibit basic properties due to the hydrolysis of the conjugate base, Bro^-. The hydrolysis reaction:

$Bro^- + H_2O \rightleftharpoons HBro + OH^-$

determines whether the solution is acidic, neutral, or basic.

Key points:
- The weak acid (HBro) has a small Ka, indicating its conjugate base (Bro^-) will likely produce OH^- ions, making the solution basic.
- Potassium salts generally do not hydrolyze or influence pH significantly, unless the conjugate base is weakly basic or strongly basic.

In this scenario:
- Since the acid is weak ($K_a \approx 2.0\times10^{-9}$), the conjugate base Bro^- will be somewhat basic, leading to an increase in pH.

Step-by-Step Calculation of pH

1. Determine the initial concentration of Bro^-:

Given the salt KBro is 0.10 M, and it dissociates completely:

$[Bro^-]_0 = 0.10$ M

2. Write the hydrolysis equilibrium:

$Bro^- + H_2O \rightleftharpoons HBro + OH^-$

Let x = concentration of OH^- produced at equilibrium.

3. Establish the equilibrium expression:

The base hydrolysis constant (Kb) relates to Ka via:

$K_b = K_w / K_a$

where $K_w = 1.0\times10^{-14}$ (at 25°C).

Calculating Kb:

$$K_b = 1.0\times10^{-14} / 2.0\times10^{-9} = 5.0\times10^{-6}$$

4. Set up the ICE table for hydrolysis:

Species	Initial (M)	Change (M)	Equilibrium (M)
Bro^-	0.10	$-x$	$0.10 - x$
OH^-	0	$+x$	x

The equilibrium expression:

$$K_b = [OH^-]^2 / [Bro^-] \approx x^2 / (0.10 - x)$$

Since Kb is small, x will be small, and the approximation:

$$0.10 - x \approx 0.10$$

5. Solve for x:

$$x^2 \approx K_b \times 0.10$$

$$x^2 \approx 5.0\times10^{-6} \times 0.10 = 5.0\times10^{-7}$$

$$x \approx \sqrt{(5.0\times10^{-7})} \approx 7.07\times10^{-4}\ M$$

6. Calculate pOH and pH:

$$pOH = -\log[OH^-] \approx -\log(7.07\times10^{-4}) \approx 3.15$$

$$pH = 14 - pOH \approx 14 - 3.15 \approx 10.85$$

Interpreting the Results: Is the Solution Acidic, Neutral, or Basic?

Based on the calculations, the pH of the 0.10 M KBro solution is approximately 10.85, indicating a basic solution. This is consistent with the expectation that the conjugate base of a weak acid hydrolyzes to produce hydroxide ions, increasing the pH.

Implications:
- The solution is mildly basic.
- The degree of hydrolysis is limited but sufficient to shift the pH above 7.
- The pH is not strongly alkaline, reflecting the weak basic nature of Bro^-.

Conclusion and Broader Implications

The calculation and analysis of the pH for a 0.10 M KBro solution demonstrate the importance of understanding acid-base equilibria, conjugate relationships, and hydrolysis reactions. Despite the initial confusion over the Ka value, assuming it corresponds to a typical weak acid ($K_a = 2.0 \times 10^{-9}$) aligns with known chemical properties of hypobromous acid and its conjugate base.

This case underscores several critical points:

- The importance of verifying data accuracy, especially when encountering anomalously high Ka values.
- The relationship between Ka and Kb, and how conjugate bases influence solution pH.
- The methodology for calculating pH in salt solutions derived from weak acids, involving hydrolysis equilibria.

In broader terms, understanding how various salts influence solution pH is vital in fields ranging from analytical chemistry to environmental science, where salt composition impacts water chemistry, biological systems, and industrial processes. The nuanced interplay between dissociation constants, hydrolysis, and ionic strength defines the behavior of solutions and informs practical applications.

Final Note: Always ensure the correct interpretation of data and recognize the chemistry principles that underpin pH calculations. As demonstrated, a careful, stepwise approach can yield accurate insights into solution behavior, essential for scientists, educators, and students alike.

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