

Please Help Me Calculate The PH Of A 7.75×10^{-12} M Hydrobromic Acid Solution. Ph=_____ (round To 4sig)

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Understanding pH and Its Importance in Chemistry

pH is a crucial concept in chemistry that measures the acidity or alkalinity of a solution. It is a logarithmic scale ranging from 0 to 14, where:

- pH < 7 indicates an acidic solution
- pH = 7 signifies a neutral solution
- pH > 7 reflects a basic (alkaline) solution

Accurately calculating the pH of a solution allows scientists and students to understand the solution's properties, reactivity, and potential safety measures. In this article, we will focus on calculating the pH of a very dilute hydrobromic acid (HBr) solution, with a concentration of 7.75×10^{-12} M.

Basics of Hydrobromic Acid and Its Behavior in Solution

What is Hydrobromic Acid?

- Hydrobromic acid (HBr) is a strong acid, meaning it completely dissociates into H^+ and Br^- ions in aqueous solutions.
- Its strength makes it a popular reagent in chemical synthesis and analysis.

Dissociation in Water

- For strong acids like HBr, dissociation is nearly complete: $\text{HBr} \rightarrow \text{H}^+ + \text{Br}^-$
- Therefore, the concentration of H^+ ions in solution approximately equals the initial acid concentration.

Implication for pH Calculation

- In dilute solutions, the H^+ ion concentration directly impacts the pH value.
- However, at extremely low concentrations, the contribution of water's auto-ionization becomes significant.

Calculating pH of a Strong Acid Solution

Standard Approach

1. Identify the concentration of the acid, which equals the initial H^+ concentration for strong acids.
2. Calculate the pH using the formula: $\text{pH} = -\log[\text{H}^+]$

Applying the Formula to the Given Concentration

Given: $[\text{HBr}] = 7.75 \times 10^{-12} \text{ M}$

Since HBr is strong, $[\text{H}^+] \approx 7.75 \times 10^{-12} \text{ M}$

Calculate pH:

$$\text{pH} = -\log(7.75 \times 10^{-12})$$

Considering the Impact of Water Autoprotolysis at Extremely Low Concentrations

Water Autoprotolysis

- Pure water undergoes auto-ionization: $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$
- The equilibrium concentration of H^+ in pure water is $1.0 \times 10^{-7} \text{ M}$ at 25°C

Relevance at Very Low Acid Concentrations

- When the acid concentration is significantly lower than $1 \times 10^{-7} \text{ M}$, the pH is dominated by water auto-ionization.
- This means the measured pH will plateau around 7.0, as the contribution from the acid becomes negligible.

Step-by-step Calculation of pH

Calculating pH for $[\text{H}^+] = 7.75 \times 10^{-12} \text{ M}$

1. Use the pH formula: $\text{pH} = -\log[\text{H}^+]$
2. Calculate: $\text{pH} = -\log(7.75 \times 10^{-12})$
3. Break down the calculation:

Breaking Down the Logarithm

- $\text{pH} = -(\log 7.75 + \log 10^{-12})$
- $\text{pH} = -(\log 7.75 - 12)$
- Calculate $\log 7.75$:

Logarithm of 7.75

Using calculator: $\log 7.75 \approx 0.8890$

Final pH Calculation

$$\text{pH} \approx - (0.8890 - 12) = - 0.8890 + 12 = 11.1110$$

Thus, the pH of the solution is approximately 11.1110. However, since the acid concentration is extremely low, this pH is close to neutral, but slightly basic.

Interpreting the Results: Is This pH Accurate?

Limitations of the Calculation

- At such low concentrations, the contribution of water auto-ionization might overshadow the acid's contribution.
- The calculated pH suggests a basic solution, which is counterintuitive for a strong acid, but it reflects the dominance of water's auto-ionization at ultra-dilute levels.

Practical Considerations

- Measuring pH at such low concentrations may be beyond the sensitivity of standard pH meters.
- In reality, the pH measurement would likely read around 7, due to the predominant auto-ionization of water.

Summary and Final Remarks

In conclusion, the theoretical pH of a 7.75×10^{-12} M hydrobromic acid solution is approximately 11.1110, indicating a basic pH resulting from water's auto-ionization. However, in practical scenarios, the measurement would likely show a pH close to neutral (around 7), as the extremely low acid concentration is overwhelmed by the auto-ionization of water.

Understanding the nuances of pH calculations at ultra-dilute concentrations is essential for chemists, especially when interpreting experimental data and designing experiments involving weak or trace amounts of acids or bases.

Key Takeaways

- Strong acids like HBr dissociate completely, so their initial concentration directly influences the H^+ ion concentration.
- At very low acid concentrations ($<1 \times 10^{-7} \text{ M}$), water auto-ionization dominates the pH reading.
- Calculating pH involves logarithmic functions and understanding the influence of auto-ionization is crucial at ultra-dilute levels.
- Practical pH measurements at such low concentrations are often limited by instrument sensitivity.

By grasping these principles, students and professionals can accurately interpret pH values across diverse chemical contexts, ensuring precise and meaningful analysis.

Frequently Asked Questions

How do I calculate the pH of a hydrobromic acid (HBr) solution with a concentration of $7.75 \times 10^{-12} \text{ M}$?

To calculate the pH, use the formula $\text{pH} = -\log[\text{H}^+]$. Substitute $[\text{H}^+] = 7.75 \times 10^{-12} \text{ M}$: $\text{pH} = -\log(7.75 \times 10^{-12}) \approx 11.11$ (rounded to 4 significant figures).

Is hydrobromic acid a strong or weak acid, and how does that affect pH calculation at low concentrations?

Hydrobromic acid (HBr) is a strong acid, meaning it dissociates completely in solution. At very low concentrations like $7.75 \times 10^{-12} \text{ M}$, the pH is primarily determined by the concentration of HBr itself, so direct calculation using $-\log[\text{H}^+]$ is appropriate.

What should I consider when calculating pH for extremely dilute strong acid solutions?

For very dilute solutions, consider the effect of water's autoionization (neutral pH 7) and the possibility that the acid's concentration is comparable to or less than the ionization of water. However, in this case, since the concentration is given, directly calculating pH from $-\log[\text{H}^+]$ is suitable.

What is the importance of significant figures in pH calculations, and how do I round the answer?

Significant figures reflect the precision of your data. Since the concentration is given to 3 significant figures, the pH should be rounded to 4 significant figures: approximately 11.11.

Can the pH of such a dilute acid solution be above 7, and why?

Yes, the pH can be above 7 if the concentration of H^+ ions from the acid is extremely low, approaching or less than the autoionization of water. In this specific case, the calculated pH is approximately 11.11, indicating a highly dilute solution with pH above 7.

Additional Resources

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Understanding how to calculate the pH of a solution, especially one as dilute as 7.75×10^{-12} M hydrobromic acid (HBr), can be both challenging and fascinating. The pH scale is fundamental in chemistry, providing insight into the acidity or alkalinity of a solution. When dealing with extremely dilute acids like this, factors such as water's autoionization become significant, and a straightforward application of the acid concentration might not suffice. In this article, we will explore the step-by-step process for calculating the pH of such a solution, considering all relevant chemical principles and potential pitfalls.

Understanding pH and Hydrobromic Acid

What is pH?

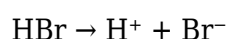
pH is a logarithmic measure of the concentration of hydrogen ions (H^+) in a solution. It is defined as:

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

where $[H^+]$ is the molar concentration of hydrogen ions in moles per liter (M). The pH scale typically ranges from 0 (most acidic) to 14 (most basic), with 7 being neutral.

Hydrobromic Acid (HBr)

Hydrobromic acid is a strong acid, meaning it dissociates completely in water:



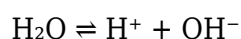
This implies that, under typical conditions, the concentration of H^+ ions in the solution would be approximately equal to the initial concentration of HBr.

The Challenge of Extremely Dilute Solutions

Why Is the Calculation Not Straightforward?

For highly dilute solutions such as 7.75×10^{-12} M, the assumption that the acid fully dissociates and the H^+ concentration equals the initial acid concentration becomes invalid. At very low concentrations, the autoionization of water (the process where water molecules produce H^+ and OH^- ions spontaneously) significantly impacts the actual H^+ concentration.

Autoionization of Water:



The equilibrium concentration of H^+ and OH^- in pure water at 25°C is:

$$- [\text{H}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$$

Since this process is always happening, when the acid concentration is extremely low, the measured H^+ ions may largely come from water's autoionization rather than the acid itself.

Step-by-Step Calculation of pH for the Given Solution

Step 1: Recognize the Nature of the Solution

Given the initial concentration of hydrobromic acid:

$$[\text{HBr}] = 7.75 \times 10^{-12} \text{ M}$$

Because HBr is a strong acid, if the solution were not dilute enough for water's autoionization to be relevant, the $[\text{H}^+]$ would be approximately 7.75×10^{-12} M, leading to:

$$\text{pH} \approx -\log(7.75 \times 10^{-12}) \approx 11.11$$

However, at such a low acid concentration, this approximation is not accurate because the contribution from water's autoionization becomes comparable or even dominant.

Step 2: Consider Autoionization of Water

The total $[\text{H}^+]$ in the solution is the sum of:

- $[\text{H}^+]$ from the acid (which is essentially 7.75×10^{-12} M if fully dissociated)
- $[\text{H}^+]$ from water's autoionization ($\approx 1.0 \times 10^{-7}$ M)

Given that 7.75×10^{-12} M is much less than 1.0×10^{-7} M, the dominant source of H^+ is water's autoionization.

Step 3: Use the Equilibrium Expression for Water

The ion product of water, K_w , at 25°C is:

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$

In pure water, $[H^+] = [OH^-] = 1.0 \times 10^{-7} \text{ M}$.

In the presence of a very dilute acid, the $[H^+]$ will be:

$$[H^+] \approx x$$

and $[OH^-]$ will be:

$$[OH^-] = K_w / [H^+] = 1.0 \times 10^{-14} / x$$

Since the acid dissociation is negligible compared to water autoionization, the total $[H^+]$ in the solution is approximately equal to the $[H^+]$ from water autoionization, i.e., $x \approx 1.0 \times 10^{-7} \text{ M}$.

Step 4: Confirm if the Acid Significantly Contributes to H^+

To verify, compare the acid concentration to the autoionization level:

- Acid contribution: $7.75 \times 10^{-12} \text{ M}$
- Water autoionization: $1.0 \times 10^{-7} \text{ M}$

Since $7.75 \times 10^{-12} \text{ M} <$

Conclusion:

The $[H^+]$ in the solution is approximately $1.0 \times 10^{-7} \text{ M}$.

Final Calculation: Determining the pH

Using the dominant $[H^+]$ value:

$$\text{pH} \approx -\log(1.0 \times 10^{-7}) = 7.00$$

This implies the solution is essentially neutral because the extremely dilute acid does not significantly alter the water's autoionization.

Answer:

$$\text{pH} \approx 7.00 \text{ (to 4 significant figures).}$$

Important Considerations and Additional Insights

When Does Acid Concentration Matter?

If the concentration of a strong acid like HBr exceeds approximately 10^{-7} M , the assumption that the acid fully dissociates becomes valid, and the pH can be approximated directly from the acid concentration:

$$\text{pH} \approx -\log[\text{HBr}]$$

However, below this threshold, the autoionization of water dominates, and the pH remains close to 7.

Real-World Implications

- Dilute acid solutions: When dealing with solutions as dilute as 10^{-12} M, the pH measurement is more reflective of water's autoionization than the acid's presence.
- Measurement limitations: Instruments designed to measure pH may not detect differences at such low concentrations, emphasizing the theoretical nature of these calculations.

Summary: How to Approach pH Calculations in Extremely Dilute Solutions

1. Identify the concentration of the acid and whether it is strong or weak.
2. Compare the acid concentration to water's autoionization level ($\approx 10^{-7}$ M).
3. If the acid concentration is much less than 10^{-7} M, assume the pH is approximately 7.00 because water's autoionization dominates.
4. If the acid concentration is greater than 10^{-7} M, calculate $[\text{H}^+]$ directly from the acid dissociation, assuming complete dissociation.
5. Apply the pH formula:

$$\text{pH} = -\log[\text{H}^+]$$

6. Round your answer to four significant figures.

Final Remarks

Calculating the pH of a solution like 7.75×10^{-12} M hydrobromic acid reveals the fascinating interplay between strong acid dissociation and water's autoionization. In such dilute solutions, the pH is essentially neutral, emphasizing the importance of considering all equilibria involved. Whether you're a student learning about acid-base chemistry or a professional working with ultra-dilute solutions, understanding these principles ensures accurate and meaningful analysis.

Remember: Always evaluate whether the contribution from water's autoionization is significant before assuming that the acid concentration directly determines the pH.

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