

You Have An Aqueous Solution For Which $[H_3O^+] = 2.5 \times 10^{-4} \text{ M}$. What Is The Ph Of This Solution?

You Have An Aqueous Solution For Which $[H_3O^+] = 2.5 \times 10^{-4} \text{ M}$. What Is The pH Of This Solution?

Understanding the pH of a solution is fundamental in chemistry, especially in fields involving acids, bases, and aqueous solutions. The pH value not only indicates the acidity or alkalinity of a solution but also impacts various chemical reactions, biological processes, and industrial applications. This article provides a comprehensive explanation of how to determine the pH of an aqueous solution given the concentration of hydrogen ions, specifically when $[H_3O^+] = 2.5 \times 10^{-4} \text{ M}$, along with related concepts, calculations, and practical considerations.

What Is pH and Why Is It Important?

Definition of pH

pH is a logarithmic measure of the hydrogen ion concentration ($[H_3O^+]$) in a solution. It quantifies the acidity or alkalinity of the solution. The pH scale ranges from 0 to 14, with:

- pH < 7 indicating an acidic solution
- pH = 7 indicating a neutral solution
- pH > 7 indicating a basic (alkaline) solution

The pH scale is logarithmic, meaning each whole number change represents a tenfold change in hydrogen ion concentration.

Significance of pH in Chemistry and Beyond

- In Chemistry: pH affects reaction rates, equilibria, solubility, and the stability of compounds.
- In Biology: Enzyme activity, cell function, and overall health depend heavily on maintaining proper pH levels.
- In Industry: manufacturing processes, water treatment, and food preservation rely on controlling pH.

Understanding the Relationship Between $[H_3O^+]$

and pH

The pH Formula

The pH of a solution is calculated using the formula:

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

Where:

- $\log$ is the base-10 logarithm
- $[\text{H}_3\text{O}^+]$ is the molar concentration of hydronium ions in moles per liter (M)

Why Use Logarithms?

Because $[\text{H}_3\text{O}^+]$ concentrations can vary widely—from 10^{-1} M to 10^{-14} M—the logarithmic scale simplifies the expression of these values, making it easier to interpret and compare acidity levels.

Calculating the pH for a Given $[\text{H}_3\text{O}^+]$

Given Data

In this scenario, the hydronium ion concentration is:

- $[\text{H}_3\text{O}^+] = 2.5 \times 10^{-4}$ M

Step-by-Step Calculation

1. Identify the $[\text{H}_3\text{O}^+]$ value: 2.5×10^{-4} M
2. Apply the pH formula:

$$\text{pH} = -\log(2.5 \times 10^{-4})$$

3. Use logarithmic properties:

Recall that $\log(ab) = \log a + \log b$. Therefore,

$$\text{pH} = -[\log 2.5 + \log 10^{-4}]$$

4. Calculate individual logs:

$$-\log 2.5 \approx 0.3979$$

$$-\log 10^{-4} = -4$$

5. Combine the results:

$$\text{pH} = -(0.3979 - 4) = -0.3979 + 4 = 3.6021$$

6. Final pH value:

$$\text{pH} \approx 3.60$$

Therefore, the pH of the solution is approximately 3.60.

Interpreting the pH Value

A pH of approximately 3.60 indicates that the solution is acidic, but not strongly so. For comparison:

- Strong acids like hydrochloric acid (HCl) in high concentration have pH values close to 1 or 2.
- Weak acids or dilute acids tend to have pH values between 3 and 6.
- A neutral solution like pure water has a pH of 7.

This pH value suggests the solution is mildly acidic, which could be characteristic of certain beverages, biological fluids, or diluted acid solutions.

Related Concepts and Additional Calculations

Calculating $[\text{H}_3\text{O}^+]$ from pH

If you know the pH and want to find $[\text{H}_3\text{O}^+]$, use:

$$[\text{H}_3\text{O}^+] = 10^{(-\text{pH})}$$

Example: If $\text{pH} = 3.60$,

$$[\text{H}_3\text{O}^+] = 10^{(-3.60)} \approx 2.51 \times 10^{-4} \text{ M}$$

which aligns with the original data provided.

Understanding pOH and Its Relation to pH

pOH is another measure of alkalinity, calculated as:

$$\text{pOH} = -\log [\text{OH}^-]$$

Since $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ are related by the water dissociation constant (K_w):

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

and

$$\text{pH} + \text{pOH} = 14$$

Given $\text{pH} \approx 3.60$,

$$\text{pOH} = 14 - 3.60 = 10.40$$

and

$$[\text{OH}^-] = 10^{(-10.40)} \approx 3.98 \times 10^{-11} \text{ M}$$

This confirms the solution's acidity.

Practical Applications and Considerations

Measuring pH in Laboratory Settings

- pH meters: Accurate electronic devices that measure the voltage difference correlated with pH.
- pH indicators: Acid-base dyes that change color depending on the pH range.
- Calibration: Essential to ensure precise pH readings, typically using standard buffer solutions.

Influence of Temperature

The value of K_w varies with temperature, affecting pH calculations. At higher temperatures, K_w increases, which can slightly alter the relationship between pH and $[\text{H}_3\text{O}^+]$.

Limitations and Sources of Error

- Contamination or impurities can skew measurements.
- Calibration errors in pH meters.
- Ionic strength and activity coefficients can influence the accuracy of $[\text{H}_3\text{O}^+]$ estimations.

Conclusion

Calculating the pH of an aqueous solution based on hydrogen ion concentration is a fundamental skill in chemistry. Given $[\text{H}_3\text{O}^+] = 2.5 \times 10^{-4} \text{ M}$, the pH is approximately 3.60, indicating a mildly acidic solution. Understanding how to perform this calculation, interpret the pH value, and relate it to other properties like pOH and hydroxide ion concentration is vital for students, researchers, and professionals working with acid-base chemistry. Proper measurement techniques and awareness of influencing factors ensure accurate pH determination, which is essential for various scientific and industrial applications.

Whether you are analyzing biological fluids, preparing chemical solutions, or conducting environmental assessments, mastering pH calculations provides a critical tool for understanding and controlling chemical phenomena in aqueous environments.

Frequently Asked Questions

What is the formula to calculate pH from the concentration of H_3O^+ ?

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

Given $[\text{H}_3\text{O}^+] = 2.5 \times 10^{-4} \text{ M}$, how do you compute the pH of the solution?

$$\text{pH} = -\log(2.5 \times 10^{-4}) \approx 3.60$$

Why is pH an important measure for aqueous solutions?

pH indicates the acidity or alkalinity of a solution, which affects chemical reactivity and biological processes.

What does a pH of 3.60 tell us about the solution's acidity?

It indicates the solution is acidic since pH is below 7.

How does the concentration of H_3O^+ ions relate to the strength of an acid in solution?

Higher $[\text{H}_3\text{O}^+]$ concentrations correspond to stronger acids, as they produce more H_3O^+ ions.

If the $[\text{H}_3\text{O}^+]$ was $1.0 \times 10^{-7} \text{ M}$, what would be the pH?

$\text{pH} = -\log(1.0 \times 10^{-7}) = 7$, which is neutral.

What is the significance of calculating pH in chemical and biological contexts?

It helps determine the environment's acidity, affecting enzyme activity, solubility, and chemical reactions.

Can you convert the pH value back to the $[\text{H}_3\text{O}^+]$ concentration?

Yes, $[\text{H}_3\text{O}^+] = 10^{(-\text{pH})}$. For $\text{pH} \approx 3.60$, $[\text{H}_3\text{O}^+] \approx 2.5 \times 10^{-4} \text{ M}$.

Additional Resources

Understanding pH Calculation for an Aqueous Solution with $[\text{H}_3\text{O}^+] = 2.5 \times 10^{-4} \text{ M}$

Introduction to pH and Its Significance in Chemistry

In aqueous chemistry, the concept of pH is fundamental to understanding the acidity or alkalinity of a solution. The pH scale provides a quantitative measure of hydrogen ion concentration ($[\text{H}_3\text{O}^+]$) in a solution, which directly influences a solution's chemical behavior, reactivity, and biological relevance. Whether you're studying environmental science, biochemistry, or industrial chemistry, an understanding of pH calculations is essential.

The Basics of pH and Hydrogen Ion Concentration

Definition of pH:

pH is defined as the negative logarithm (base 10) of the hydrogen ion concentration:

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\[ \boxed{\mathrm{pH} = -\log [\mathrm{H}_3\mathrm{O}^+]} \]
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Significance:

- Acidic solutions: $\text{pH} < 7$
- Neutral solutions: $\text{pH} = 7$
- Basic (alkaline) solutions: $\text{pH} > 7$

The pH scale generally ranges from 0 (most acidic) to 14 (most basic), although values outside this range are possible in extreme conditions.

Calculating the pH from Hydrogen Ion Concentration

Given the $[\text{H}_3\text{O}^+]$, calculating pH involves straightforward logarithmic computation:

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

In this specific case, the $[\text{H}_3\text{O}^+]$ is provided as:

$$[\text{H}_3\text{O}^+] = 2.5 \times 10^{-4} \text{ M}$$

Step-by-Step pH Calculation

Step 1: Write down the known $[\text{H}_3\text{O}^+]$ value:

$$[\text{H}_3\text{O}^+] = 2.5 \times 10^{-4} \text{ M}$$

Step 2: Apply the pH formula:

$$\text{pH} = -\log (2.5 \times 10^{-4})$$

Step 3: Use logarithmic properties to simplify:

$$\text{pH} = -(\log 2.5 + \log 10^{-4})$$

Recall that:

$$-\log (a \times 10^b) = -\log a + b$$

$$-\log 10^b = b$$

So,

$$\mathrm{pH} = -\left(\log 2.5 - 4 \right)$$

Step 4: Calculate $(\log 2.5)$:

$$\log 2.5 \approx 0.3979$$

Step 5: Final calculation:

$$\mathrm{pH} = - (0.3979 - 4) = -0.3979 + 4 = 3.6021$$

Rounded to two decimal places:

$$\boxed{\mathrm{pH} \approx 3.60}$$

This indicates the solution is moderately acidic.

Interpreting the pH Value

A pH of approximately 3.60 suggests:

- The solution is acidic but not strongly so.
- It could be similar in acidity to fruit juices like orange juice or vinegar diluted with water.
- The hydrogen ion concentration is higher than neutral water ($[\mathrm{H}_3\mathrm{O}^+] = 1.0 \times 10^{-7} \text{ M}$), confirming acidity.

Understanding the Relationship Between $[\mathrm{H}_3\mathrm{O}^+]$ and pH

The logarithmic nature of pH means small changes in hydrogen ion concentration result in significant pH shifts:

- For example, decreasing $[\mathrm{H}_3\mathrm{O}^+]$ by a factor of 10 increases pH by 1.
- Conversely, increasing $[\mathrm{H}_3\mathrm{O}^+]$ by a factor of 10 decreases pH by 1.

In this case:

$$\begin{aligned} & \backslash \\ & \mathrm{pH} \approx 3.60 \rightarrow [\mathrm{H}_3\mathrm{O}^+] = 2.5 \times 10^{-4} \text{ M} \\ & \backslash \end{aligned}$$

which is higher than neutral water, confirming acidity.

Implications of pH in Various Contexts

1. Biological Systems:

- Many biological processes are pH-sensitive.
- Human blood maintains a pH around 7.4; solutions with pH 3.6 are far more acidic and can be harmful or require buffering.

2. Environmental Chemistry:

- Acidic solutions ($\text{pH} < 5.6$) can lead to soil acidification, affecting plant life.
- Acid rain typically has a pH below 5, which can be compared to the calculated pH to understand environmental impact.

3. Industrial Applications:

- pH control is critical in manufacturing, pharmaceuticals, and water treatment.
- Knowing the precise pH helps in designing processes and ensuring safety.

Additional Considerations in pH Calculations

1. Logarithmic Calculations:

- Always verify your logarithm calculations, especially when dealing with scientific notation.
- Use a calculator with a log function for accuracy.

2. Significance of Significant Figures:

- The $[\text{H}_3\text{O}^+]$ value has two significant figures; thus, the pH should be reported with two decimal places for consistency and precision.

3. pH of Diluted Solutions:

- When diluting acids or bases, the pH changes predictably based on dilution factors, following the principles of molarity and logarithms.

4. pH of Strong vs. Weak Acids:

- For strong acids like HCl, the $[\text{H}_3\text{O}^+]$ is directly related to the initial concentration.
- For weak acids, equilibrium calculations are necessary, but in this problem, direct calculation suffices.

Additional Related Calculations and Concepts

1. pOH Calculation:

Since:

$$\mathrm{pH} + \mathrm{pOH} = 14$$

we can find:

$$\mathrm{pOH} = 14 - 3.60 = 10.40$$

and

$$[\mathrm{OH}^-] = 10^{-\mathrm{pOH}} = 10^{-10.40} \approx 3.98 \times 10^{-11} \text{ M}$$

This low hydroxide concentration confirms the acidic nature.

2. Acid Dissociation Constant (K_a):

If this solution results from a weak acid, understanding its K_a could involve further calculations, but with the given data, it's beyond scope.

Summary and Final Remarks

In conclusion, calculating the pH of a solution with a known $[\text{H}_3\text{O}^+]$ concentration is a fundamental skill in chemistry. For the given $[\text{H}_3\text{O}^+] = 2.5 \times 10^{-4} \text{ M}$, the pH is approximately 3.60, indicating an acidic solution. This calculation underscores the importance of logarithmic functions in chemistry and highlights how subtle changes in hydrogen ion concentration significantly impact the solution's acidity.

Understanding these principles enables chemists and students alike to analyze solutions accurately, predict their behavior, and apply this knowledge across various scientific disciplines. Whether designing pharmaceuticals, assessing environmental impact, or studying biological systems, mastering pH calculations remains an essential component of chemical literacy.

In essence, the pH of the solution with $[H_3O^+] = 2.5 \times 10^{-4} \text{ M}$ is approximately 3.60, reflecting a moderately acidic environment with broad implications across scientific and practical applications.

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