What Is The Ka Of The Acid Ha Given That A 1.20 M Solution Of The Acid Has A Ph Of 0.20? The Equation

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Understanding the acidity of a solution involves calculating the acid dissociation constant, known as Ka. When given specific data such as concentration and pH, we can determine Ka by applying foundational principles of acid-base chemistry. In this article, we will explore the step-by-step methodology to find the Ka of an acid named Ha, given that a 1.20 M solution exhibits a pH of 0.20. We will delve into the relevant equations, assumptions, and calculations that lead to the precise value of Ka.

Fundamental Concepts in Acid-Base Chemistry

What Is pH and How Is It Related to Hydrogen Ion Concentration?

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

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

- Hydrogen Ion Concentration: From pH, we can derive:

$$
[\mathrm{H}^+] = 10^{-\text{pH}}
$$

- Implication: A pH of 0.20 indicates a highly acidic solution with a high concentration of hydrogen ions.

Understanding Acid Dissociation and Ka

- Ionization of a Monoprotic Acid (Ha):

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

- Expression for Ka:

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

- Assumption: The initial concentration of Ha is known, and some fraction dissociates to produce $[\mathrm{H}^+]$ and $[\mathrm{A}^-]$.

Calculating the Hydrogen Ion Concentration from pH

Deriving [H+] from pH

Given the pH:

$$
\text{pH} = 0.20
$$

Calculate:

$$
[\mathrm{H}^+] = 10^{-\text{pH}} = 10^{-0.20}
$$

Using a calculator:

$$
[\mathrm{H}^+] \approx 0.631 \, \text{M}

\]

This high concentration confirms the acid is strongly dissociating.

Setting Up the Acid Dissociation Equilibrium

Initial Concentration and Dissociation

- Initial concentration of Ha:

$$
[\mathrm{Ha}]_{initial} = 1.20\, \text{M}
$$

- Change due to dissociation:

$$
\mathrm{Ha} \rightarrow \mathrm{H}^+ + \mathrm{A}^-
$$

- The amount dissociated is approximately equal to $[\mathrm{H}^+]$ because of the 1:1 dissociation ratio.

- Equilibrium concentrations:

$$
[\mathrm{H}^+]_{eq} = x = 0.631\, \text{M}
$$

$$
[\mathrm{A}^-]_{eq} = x = 0.631\, \text{M}
$$

$$
[\mathrm{Ha}]_{eq} = 1.20 - x = 1.20 - 0.631 = 0.569\, \text{M}
$$

Assumptions in the Calculation

- The dissociation is significant, but not complete, suitable for strong acid approximation.

- The change in concentration due to dissociation is large compared to initial concentrations, so the approximation holds.

- The system reaches equilibrium rapidly.

Calculating the Acid Dissociation Constant (Ka)

Plugging Values into the Ka Expression

$$
K_a = \frac{[\mathrm{H}^+][\mathrm{A}^-]}{[\mathrm{Ha}]} = \frac{(0.631)(0.631)}{0.569}
$$

Calculate numerator:

$$
(0.631)^2 \approx 0.398
$$

Divide by the remaining concentration:

$$
K_a \approx \frac{0.398}{0.569} \approx 0.699
$$

Thus,

$$
\boxed{
K_a \approx 0.70
}
$$

This indicates that the acid Ha is strongly dissociating, given the high $[\mathrm{H}^+]$ and large Ka value.

Discussion of Results and Implications

Understanding the Significance of a High Ka

- The calculated Ka value (~0.70) suggests that Ha is a strong acid because typical weak acids have Ka values much less than 1.

- Comparison with Known Acids:

- Strong acids like HCl or H_2SO_4 have very high Ka or dissociation constants, often approaching or exceeding 1.

- Implication for Solution Behavior:

- The solution will be highly conductive due to abundant free hydrogen ions.

- The pH is very low, consistent with the observed pH of 0.20.

Limitations and Considerations

- The calculation assumes complete dissociation due to the high $[\mathrm{H}^+]$, but for very strong acids, the concept of Ka becomes less meaningful because they dissociate completely.

- The initial assumptions are valid since the pH indicates significant dissociation.

- For acids with lower pH or weaker dissociation, a more nuanced approach considering activity coefficients may be necessary.

Conclusion

- Given the concentration of the acid (1.20 M) and the pH (0.20), the hydrogen ion concentration is approximately 0.631 M.

- Using the dissociation equilibrium and the initial concentration, the Ka of the acid Ha is approximately 0.70.

- This high Ka value classifies Ha as a strong acid, with nearly complete dissociation in aqueous solution.

Understanding how to determine Ka from pH and concentration enables chemists to classify acids and predict their behavior in various chemical contexts. Accurate calculations depend on meticulous application of equilibrium principles and recognition of the assumptions involved.

Key Points Summary:

- pH directly yields $[\mathrm{H}^+]$.

- The equilibrium concentrations are derived assuming a 1:1 dissociation.

- The Ka value is calculated using the equilibrium concentrations.

- The resulting high Ka indicates a strong acid behavior for Ha.

If further refinement is needed, experimental data or advanced models considering activity coefficients can be employed, but for most practical purposes, this approach provides a reliable estimate of the acid's dissociation constant.

Frequently Asked Questions

What is the definition of the acid dissociation constant (Ka)?

Ka is the equilibrium constant that measures the strength of an acid in solution, representing the ratio of the concentrations of dissociated ions to the undissociated acid at equilibrium.

How do you calculate the pKa from the pH and the concentration of the

acid?

pKa can be calculated using the relationship $pKa = pH + \log([A^-]/[HA])$. If $[A^-]$ is negligible initially, it can be approximated based on the degree of dissociation inferred from pH.

Given a 1.20 M solution with a pH of 0.20, how do you find the concentration of H+ ions?

The concentration of H+ ions is $[H^+] = 10^{-pH} = 10^{-0.20} \approx 0.63$ M.

What is the dissociation equation for the acid HA?

The dissociation equation is $HA \rightleftharpoons H^+ + A^-$.

How is the Ka value calculated from the given data?

$Ka = [H^+][A^-]/[HA]$. Assuming complete dissociation at the initial stage, $[A^-] \approx [H^+] \approx 0.63$ M, and $[HA] \approx$ initial concentration minus dissociated amount, which can be approximated for weak acids.

What assumptions are made when calculating Ka from pH and initial concentration?

It is assumed that the dissociation is small (weak acid), and the change in concentration due to dissociation is negligible compared to the initial concentration, or that dissociation is significant enough to approximate $[A^-]$ as $[H^+]$.

What is the significance of the pH value being 0.20 in this context?

A pH of 0.20 indicates a very high concentration of H+ ions (~0.63 M), suggesting the acid is strong or highly dissociative; however, since the initial concentration is only 1.20 M, it may be a strong acid or a very dissociable weak acid.

Can you determine if HA is a strong or weak acid based on the pH and concentration provided?

Given the low pH and high H+ concentration relative to initial concentration, HA is likely a strong acid or a very strong weak acid with near-complete dissociation.

What is the formula for calculating Ka given pH and initial concentration

of the acid?

Calculate [H+] from pH, then approximate [A−] ≈ [H+], and use Ka = [H+][A−]/[HA], where [HA] ≈ initial concentration minus dissociated amount.

What is the approximate value of Ka for the acid Ha under these conditions?

Using [H+] ≈ 0.63 M, [A−] ≈ 0.63 M, and [HA] ≈ 1.20 - 0.63 = 0.57 M, then Ka ≈ (0.63)(0.63)/0.57 ≈ 0.70.

Additional Resources

Acid Dissociation Constant (Ka): Unlocking the Secrets of Acid Strength

In the fascinating world of chemistry, understanding the strength of acids is fundamental. Whether you're a student delving into foundational concepts or a professional chemist analyzing complex reactions, grasping the concept of the acid dissociation constant, Ka, is essential. Today, we explore this critical parameter through a practical example: determining the Ka of hydrocyanic acid (Ha) given specific solution conditions.

Our journey begins with an overview of the key concepts, progresses through detailed calculations, and culminates in interpreting what the Ka value reveals about hydrocyanic acid's behavior in aqueous solutions. Think of this as an expert review—comprehensive, insightful, and designed to deepen your understanding of acid strength assessment.

Understanding the Basics: What Is Ka?

Defining Ka

The acid dissociation constant, denoted as Ka, quantifies the strength of an acid in aqueous solution. It is a specific equilibrium constant that represents the extent to which an acid dissociates into its ions:

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

The equilibrium expression for this dissociation is:

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

Here,

- $[\text{H}^+]$ is the concentration of hydrogen ions,
- $[\text{A}^-]$ is the concentration of the conjugate base,
- $[\text{HA}]$ is the concentration of undissociated acid.

A higher Ka indicates a stronger acid, which dissociates more extensively in water, whereas a lower Ka signifies a weaker acid.

Why Is Ka Important?

Understanding Ka helps predict:

- The pH of an acid solution,
- The degree of dissociation,
- How an acid behaves in various chemical environments,
- Its reactivity and safety profile.

For hydrocyanic acid (HCN), knowing its Ka is crucial because it is a weak but highly toxic acid, with applications and implications spanning from industrial processes to toxicology.

Setting the Stage: The Given Data and Its Significance

Before diving into calculations, let's clarify the information provided:

- Concentration of the acid solution, [Ha]: 1.20 M
- pH of the solution: 0.20

This data offers a snapshot of the solution's acidity and concentration, forming the basis for calculating Ka.

Why is this data sufficient? Because the pH provides the concentration of hydrogen ions directly, which is the key to determining Ka, given the initial acid concentration.

Step-by-Step Calculation of Ka for Hydrocyanic Acid

Let's examine the process comprehensively, from extracting ion concentrations to applying equilibrium expressions.

1. Calculating $[\text{H}^+]$ from pH

pH is defined as:

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

Given $ pH = 0.20 $:

$$ [\text{H}^+] = 10^{-pH} = 10^{-0.20} $$

Calculating:

$$ [\text{H}^+] \approx 10^{-0.20} \approx 0.631 \text{ M} $$

This high concentration indicates a strongly acidic solution.

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2. Determining the Concentrations of Acid and Conjugate Base at Equilibrium

- Initial concentration of HA: 1.20 M
- Concentration of $[\text{H}^+]$ at equilibrium: 0.631 M

Since hydrocyanic acid dissociates as:

$$ \text{HCN} \rightleftharpoons \text{H}^+ + \text{CN}^- $$

and assuming negligible auto-ionization of water and other side reactions, the dissociation will produce approximately 0.631 M of HCN's conjugate base (CN^-) and 0.631 M of H^+.

The remaining concentration of undissociated HCN:

$$ [\text{HA}]_{eq} = 1.20 - 0.631 = 0.569 \text{ M} $$

Note: Because the initial concentration is high, and dissociation is limited, this approximation is valid.

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3. Calculating Ka Using the Equilibrium Expression

The equilibrium expression:

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

Substituting the known values:

$$K_a = \frac{(0.631)(0.631)}{0.569}$$

Calculating numerator:

$$0.631 \times 0.631 \approx 0.398$$

Then,

$$K_a \approx \frac{0.398}{0.569} \approx 0.699$$

This suggests:

$$\boxed{
K_a \approx 0.70
}$$

Interpreting the Result: What Does the Ka Tell Us About Hydrocyanic Acid?

The calculated $K_a \approx 0.70$ indicates a relatively weak acid compared to strong acids like HCl or H_2SO_4, which have Ka values close to or exceeding 1.0. Hydrocyanic acid's acidity is moderate, but the high concentration of $[\text{H}^+]$ (0.631 M) in this particular solution underscores its potency in specific contexts.

Comparison with Known Values:

The literature reports hydrocyanic acid's K_a as approximately 6.2×10^{-10}. Our calculated value appears significantly higher, which suggests that the initial assumptions or approximations in this scenario, such as the high concentration and pH, may not reflect typical dilute solutions. Instead, this example illustrates the importance of context and the influence of concentration and environmental conditions on acid strength assessments.

Why the discrepancy?

- The pH value provided (0.20) corresponds to an extremely acidic solution, unlikely for hydrocyanic acid at 1.20 M unless it behaves differently under specific conditions.
- The calculation assumes complete dissociation proportional to the hydrogen ion concentration, which may not be accurate for weak acids with very low K_a.

In real-world applications, the calculation of K_a for weak acids like hydrocyanic acid typically involves more nuanced measurements and considerations, including activity coefficients and temperature effects.

Additional Factors to Consider in Acid Strength Analysis

- Temperature Dependence: K_a varies with temperature; most values are standardized at 25°C.
- Ionization Degree: Weak acids have low degrees of ionization; pH should align with their known dissociation constants.
- Concentration Effects: At high concentrations, activity coefficients influence the effective dissociation.

Conclusion: The Significance of Determining Ka

Understanding the acid dissociation constant K_a is more than an academic exercise—it provides a window into the behavior, reactivity, and safety considerations of acids like hydrocyanic acid.

From the calculation above, we've demonstrated how to derive K_a using pH and concentration data, revealing that hydrocyanic acid's acidity in solution is moderate but highly dependent on solution conditions. Accurate determination of K_a informs chemical handling, environmental impact assessments, and applications in various industries.

In essence, whether you're a student, researcher, or industry professional, mastering the calculation and interpretation of K_a empowers you to predict and manipulate acid-base reactions effectively. This example underscores the importance of precise measurements and a nuanced understanding of equilibrium chemistry in real-world contexts.

In summary:

- The pH directly informs $[\text{H}^+]$.
- Calculating $[\text{H}^+]$ and equilibrium concentrations of other species allows for K_a determination.
- Hydrocyanic acid's K_a reflects its weak but significant acidity.
- Real-world values may differ due to experimental conditions, reinforcing the importance of context.

This comprehensive exploration underscores the power and utility of basic equilibrium principles in elucidating acid strengths, making it an indispensable tool in the chemist's toolkit.

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