What Is The PH Of A 40.0 ML Solution That Is 0.13 M In CN And 0.27 M In HCN? The Ka For HCN Is 4.9109.

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Understanding the pH of a solution containing cyanide (CN^-) and hydrocyanic acid (HCN) requires a grasp of acid-base chemistry, equilibrium principles, and how conjugate acid-base pairs influence solution pH. In this article, we will explore the step-by-step process to determine the pH of such a solution, including relevant concepts, calculations, and practical considerations.

Introduction to Acid-Base Equilibria

Before delving into the specific problem, it's essential to understand the fundamentals of acid-base chemistry:

- Acids and Bases: Substances that donate or accept protons (H^+) respectively.
- Conjugate Acid-Base Pairs: An acid and its conjugate base differ by one proton.
- pH: A measure of hydrogen ion concentration, calculated as $pH = -\log[H^+]$.

In aqueous solutions, weak acids and their conjugate bases establish equilibrium systems affecting pH. For HCN, a weak acid, and CN^-, its conjugate base, their interactions determine the solution's acidity or basicity.

Understanding the Components of the Solution

The solution contains:

- Hydrocyanic acid (HCN): 0.27 M
- Cyanide ion (CN^-): 0.13 M

The presence of both indicates a buffer system, where the weak acid (HCN) and its conjugate base (CN^-) coexist, resisting changes in pH upon addition of small amounts of acid or base.

Relevant Equilibrium and Constants

The key equilibrium for hydrocyanic acid is:

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

The acid dissociation constant (Ka) quantifies the strength of the acid:

$$ K_a = 4.9109 \times 10^{-10} $$

The conjugate base (CN^-) can also react with water (hydrolysis):

$$ \text{CN}^- + \text{H}_2O \rightleftharpoons \text{HCN} + \text{OH}^- $$

The base dissociation constant (Kb) for CN^- can be derived from Ka and Kw:

$$ K_b = \frac{K_w}{K_a} $$

where Kw is the ion product of water at 25°C:

$$ K_w = 1.0 \times 10^{-14} $$

Calculating Kb for CN^-

Given:

$$ K_a = 4.9109 \times 10^{-10} $$

Calculate:

$$ K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.9109 \times 10^{-10}} $$

$$ K_b \approx 2.037 \times 10^{-5} $$

This Kb indicates the strength of cyanide as a weak base.

Establishing the pH Calculation Framework

Since the solution contains both HCN and CN^-, it forms a buffer solution. The pH of a buffer can be estimated using the Henderson-Hasselbalch equation:

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

Where:

- $[A^-]$ = concentration of the conjugate base (CN^-)
- [HA] = concentration of the weak acid (HCN)
- pKa = -log(Ka)

Calculate pKa:

$$ pK_a = -\log(4.9109 \times 10^{-10}) $$

$$ pK_a \approx 9.308 $$

Applying the Henderson-Hasselbalch Equation

Insert known concentrations:

$$ \text{pH} = 9.308 + \log \left(\frac{0.13}{0.27} \right) $$

Calculate the ratio:

$$ \frac{0.13}{0.27} \approx 0.4815 $$

Then:

$$ \text{pH} = 9.308 + \log(0.4815) $$

Calculate the logarithm:

$$ \log(0.4815) \approx -0.317 $$

Finally:

$$ \text{pH} = 9.308 - 0.317 = 8.991 $$

Thus, the pH of the solution is approximately 8.99.

Additional Considerations and Validation

While the above calculation assumes ideal buffer behavior, it's essential to consider:

- Dilution effects: The total volume of 40.0 mL is relatively small, but since concentrations are given, the calculation remains valid.
- Activity corrections: In dilute solutions, activity coefficients could slightly alter the pH, but for this context, they are negligible.
- Presence of other ions: No other significant ions are specified, so their influence on pH is minimal.

Summary of the Calculation Process

To determine the pH of the solution:

1. Identify the components: HCN and CN^-.
2. Find the pKa of HCN: $pK_a = -\log(Ka)$.
3. Use the Henderson-Hasselbalch equation: $pH = pK_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]} \right)$.
4. Insert concentrations: $[A^-] = 0.13$ M, $[HA] = 0.27$ M.
5. Calculate the pH: Approximately 8.99.

This pH indicates a slightly basic solution, consistent with the buffer system of HCN and CN^-.

Practical Implications and Applications

Understanding the pH of such solutions is vital in various chemical and industrial processes:

- Analytical Chemistry: Buffer systems are used to maintain pH during titrations or reactions.
- Environmental Chemistry: Cyanide compounds' behavior in water bodies can influence toxicity and remediation strategies.
- Pharmaceuticals: Buffer solutions are critical in drug formulation to ensure stability and efficacy.

Knowing how to compute pH in complex systems helps chemists design appropriate solutions, predict chemical behavior, and troubleshoot issues effectively.

Conclusion

In conclusion, the pH of a 40.0 mL solution containing 0.13 M CN^- and 0.27 M HCN, with a Ka of 4.9109, is approximately 8.99. This calculation hinges on understanding acid-base equilibria, the relationship between Ka, Kb, and pKa, and the application of the Henderson-Hasselbalch equation. Mastery of these concepts allows chemists to analyze buffer systems accurately and predict solution behavior in various scientific and industrial contexts.

Further Reading and Resources

- "Chemistry: The Central Science" by Brown, LeMay, Bursten – Chapters on acids and bases
- Educational websites on buffer systems and pH calculations
- Online pH calculators for quick estimations

By integrating theoretical knowledge with practical calculation techniques, understanding the pH of complex solutions becomes an accessible and essential skill for chemistry professionals and students alike.

Frequently Asked Questions

What information is needed to determine the pH of a solution containing HCN and CN-?

You need the concentrations of HCN and CN-, the acid dissociation constant (Ka) for HCN, and the initial concentrations to set up an equilibrium expression and calculate the pH.

How does the relationship between HCN and CN- concentrations influence the pH calculation?

Since HCN is a weak acid and CN- is its conjugate base, their concentrations determine the position of equilibrium, which can be used in an ICE table to find pH using the Henderson-Hasselbalch equation.

What is the first step in calculating the pH of the given solution?

Identify the initial concentrations of HCN and CN-, then set up the equilibrium expression for HCN dissociation to find the concentration of H+ ions.

How is the Henderson-Hasselbalch equation applied in this scenario?

The equation pH = pKa + log([A-]/[HA]) is used, where [A-] is the concentration of CN- and [HA] is the concentration of HCN, to estimate the pH.

What is the pKa of HCN given its Ka value?

pKa = -log(Ka) = -log(4.9109 x 10^-10) ≈ 9.31.

How does the initial concentration of HCN (0.27 M) compare to the concentration of CN- (0.13 M) in determining pH?

Since [A-] (CN-) is less than [HA] (HCN), the solution is slightly acidic, and the pH will be below 7, but the exact value depends on the equilibrium calculations.

Can the pH be accurately calculated using the Henderson-Hasselbalch equation in this case?

Yes, because the solution contains a weak acid and its conjugate base, and the concentrations are known, making it suitable for this approximation.

What is the approximate pH of the solution based on the given data?

Using Henderson-Hasselbalch: pH ≈ 9.31 + log(0.13/0.27) ≈ 9.31 + log(0.48) ≈ 9.31 - 0.32 ≈ 8.99.

What assumptions are made in calculating the pH for this buffer solution?

It is assumed that the system reaches equilibrium, the concentrations of HCN and CN- remain relatively unchanged during the calculation, and activity coefficients are approximately 1.

Additional Resources

pH Calculation of a 40.0 mL Solution Containing 0.13 M CN^- and 0.27 M HCN: An Expert Analysis

Understanding the pH of a solution that contains both a weak acid and its conjugate base is fundamental in the realm of chemistry, especially in the context of buffer systems. The solution in question involves hydrocyanic acid (HCN), a weak acid, and its conjugate base, cyanide ion (CN^-). This scenario offers a classic example of the principles underlying buffer chemistry, the application of the Henderson-Hasselbalch equation, and the importance of equilibrium constants.

In this comprehensive exploration, we will meticulously analyze the pH of a 40.0 mL solution with concentrations of 0.13 M CN^- and 0.27 M HCN, utilizing the provided acid dissociation constant (Ka = 4.9109) for HCN. This discussion aims to serve both as an educational guide and an expert review for those interested in aqueous equilibria and buffer systems.

Fundamental Concepts in Acid-Base Chemistry

Before diving into calculations, it is essential to revisit some core concepts:

- Weak acids and their conjugates: HCN is a weak acid that partially dissociates in aqueous solution:

$HCN \rightleftharpoons H^+ + CN^-$

- Equilibrium constant (Ka): Quantifies the strength of the acid. The smaller the Ka, the weaker the acid.

- Buffer systems: Mixtures of a weak acid and its conjugate base that resist changes in pH upon addition of small amounts of acids or bases.

- pH and pKa relationship: The pH of a buffer solution can be effectively calculated using the Henderson-Hasselbalch equation when both the acid and its conjugate base are present.

Given Data and Initial Considerations

Let's organize the information provided:

Parameter	Value
Volume of solution	40.0 mL (0.040 L)
Concentration of CN^-	0.13 M
Concentration of HCN	0.27 M
Ka for HCN	4.9109

Note: The values appear to be the initial concentrations before any reactions or dilutions, but since they are given explicitly, we will assume these are the concentrations in the solution after mixing, at equilibrium.

Step 1: Validating the Data and Assumptions

Before proceeding, confirm the following:

- The solution contains both HCN and CN^- at these concentrations, representing a buffer system.
- The solution volume (40.0 mL) is consistent with the concentrations, or whether the concentrations are initial or adjusted after mixing. Since the problem states these concentrations explicitly, we will treat them as the effective concentrations at equilibrium.

Step 2: Calculating pKa from Ka

The pKa is a logarithmic measure of the acid strength:

$$
pK_a = -\log_{10} (K_a)
$$

Given:

$$
K_a = 4.9109
$$

Calculating:

$$
pK_a = -\log_{10}(4.9109) \approx -0.691
$$

Interpretation:

- Since pKa is approximately -0.69, HCN is a relatively weak acid, but notably, its pKa is negative, indicating it is stronger than many other weak acids, which typically have pKa values above 0.

Step 3: Applying the Henderson-Hasselbalch Equation

The most straightforward way to estimate the pH of this buffer solution is via the Henderson-Hasselbalch equation:

$$
pH = pK_a + \log_{10} \left(\frac{[\text{A}^-]}{[\text{HA}]} \right)
$$

Where:

- $[\text{A}^-]$ is the concentration of the conjugate base (CN^-).
- $[\text{HA}]$ is the concentration of the weak acid (HCN).

Plugging in the values:

$$
pH = -0.69 + \log_{10} \left(\frac{0.13}{0.27} \right)
$$

Calculate the ratio:

$$
\frac{0.13}{0.27} \approx 0.4815
$$

Now, compute the logarithm:

$$
\log_{10}(0.4815) \approx -0.317
$$

Finally:

$$
pH = -0.69 + (-0.317) = -1.007
$$

Result:

$$
\boxed{
\text{pH} \approx -1.01
}
$$

Interpreting the Result: Is a Negative pH Plausible?

A pH of roughly -1.01 is physically possible in extremely acidic solutions, indicating a very high concentration of free H^+ ions. However, in practical aqueous chemistry, such a negative pH suggests a highly acidic environment or, more likely, that the assumptions about the species concentrations need to be refined.

Key points:

- The pKa being negative indicates the acid is quite strong compared to typical weak acids.
- The initial concentrations suggest a buffer with a higher ratio of HCN to CN^-, which would tend to produce a mildly acidic pH, but our calculation indicates an even more acidic environment.

Step 4: Considering the Acid-Base Equilibrium in Depth

Given the apparent contradiction, it’s prudent to analyze the actual equilibrium process and whether the initial concentrations reflect the true equilibrium state.

The dissociation of HCN:

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

\]

- The equilibrium expression:

\[
K_a = \frac{[\text{H}^+][\text{CN}^-]}{[\text{HCN}]}
\]

Suppose the initial concentrations are as given, and we define:

- \(x\) as the concentration of \(\text{H}^+\) produced at equilibrium.

Then, at equilibrium:

Species	Initial (M)	Change	Equilibrium concentration (M)
HCN	0.27	\(-x\)	\(0.27 - x\)
CN^-	0.13	\(+x\)	\(0.13 + x\)

Applying the equilibrium expression:

\[
K_a = \frac{x \times (0.13 + x)}{0.27 - x}
\]

Given the small value of \(K_a\), we expect \(x\) to be small relative to initial concentrations, so approximate:

\[
K_a \approx \frac{x \times 0.13}{0.27}
\]

Solving for \(x\):

\[
x \approx \frac{K_a \times 0.27}{0.13} = \frac{4.9109 \times 0.27}{0.13}
\]

Calculate numerator:

\[
4.9109 \times 0.27 \approx 1.325
\]

Then:

\[
x \approx \frac{1.325}{0.13} \approx 10.19\, \text{M}
\]

This is nonsensical because the initial concentrations are only 0.27 M and 0.13 M, and \(x\) cannot be

larger than the initial concentrations.

Conclusion:

The approximation breaks down because the initial K_a value is large, implying the acid dissociates extensively. The calculated K_a appears inconsistent with typical values, suggesting that perhaps the given K_a is not correct or is expressed in a different context.

Step 5: Reconsidering the Ka Value and Its Implications

The value $K_a = 4.9109$ is unusual because:

- K_a values are typically less than 1 for weak acids.
- A K_a greater than 1 indicates a strong acid, but HCN is a weak acid with a K_a around 6.2×10^{-10}.

Potential Misinterpretation:

It’s plausible that the provided K_a value is not in the standard form or units, or perhaps it's a typo or misprint.

Standard value for HCN's K_a:

$$
K_a \approx 6.2 \times 10^{-10}
$$

Corresponding pK_a:

$$
pK_a \approx -\log_{10}(6.2 \times 10^{-10}) \approx 9.21
$$

Final, Refined Approach: Using Standard K_a for HCN

Given the inconsistency, assume the typical $K_a \approx 6.2 \times 10^{-10}$ for HCN.

Calculate pK_a:

$$
pK_a \approx 9.21
$$

\]

Now, applying Henderson-Hasselbalch

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