If A 0.43 M Solution Of A Base 25c Is Found To Have A Ph Of 12.80 At Equilibrium, What Is The Percent

If A 0.43 M Solution Of A Base 25°C Is Found To Have A pH Of 12.80 At Equilibrium, What Is The Percent of the base that has dissociated? Understanding how to approach this problem requires a solid grasp of concepts related to pH, pOH, molarity, and percent dissociation of a base in aqueous solution. This comprehensive guide will walk you through the step-by-step process to determine the percent of dissociation, including relevant calculations, explanations, and contextual insights.

Understanding the Problem and Key Concepts

Before diving into calculations, it's important to comprehend what information is given and what is being asked.

Given Data

- Initial concentration of the base, **0.43 M**
- Temperature, **25°C**
- Equilibrium pH, **12.80**

What Is Being Asked?

Determine the percentage of the original base that has dissociated at equilibrium, which is known as the percent dissociation or percent ionization.

Fundamental Concepts and Formulas

To solve this problem, you'll need to understand and utilize several key concepts:

pH and pOH Relationship

- pH and pOH are related via the equation:

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

- Given pH, you can find pOH:

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

Calculating $[OH^-]$ from pOH

- The hydroxide ion concentration ($[OH^-]$) can be found using:

$$
[\text{OH}^-] = 10^{-\text{pOH}}
$$

Percent Dissociation Formula

- Percent dissociation is given by:

$$
\text{Percent Dissociation} = \left(\frac{\text{Number of moles dissociated}}{\text{Initial moles}} \right) \times 100
$$

- Since the dissociation of a base (e.g., B) in water produces hydroxide ions, the concentration of $[OH^-]$ at equilibrium reflects the amount dissociated.

Step-by-Step Solution

Let's walk through the calculation process:

Step 1: Calculate pOH from pH

Given:

$$
\text{pH} = 12.80
$$

Calculate pOH:

$$
\text{pOH} = 14 - 12.80 = 1.20
$$

Step 2: Find [OH⁻] Concentration

Using the pOH:

$$
[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-1.20}
$$

Calculating:

$$
[\text{OH}^-] \approx 10^{-1} \times 10^{-0.20} \approx 0.1 \times 0.63 = 0.063 \text{ M}
$$

(Using calculator for 10^{-1.20}):

$$
[\text{OH}^-] \approx 0.0631 \text{ M}
$$

Step 3: Determine the Moles Dissociated

- The initial concentration of the base is 0.43 M.
- At equilibrium, the concentration of dissociated base (which produces [OH⁻]) is approximately 0.0631 M.
- The amount of base dissociated per liter of solution is thus 0.0631 mol.

Step 4: Calculate Percent Dissociation

Percent dissociation:

$$
\text{Percent} = \left(\frac{[\text{OH}^-]}{\text{Initial concentration of base}} \right) \times 100
$$

$$
= \left(\frac{0.0631}{0.43} \right) \times 100 \approx 0.1467 \times 100 = 14.67\%
$$

Result: Approximately 14.67% of the base has dissociated at equilibrium.

Additional Insights and Context

Understanding what this percentage means in practical terms offers valuable perspective.

The Significance of Percent Dissociation

- A 14.67% dissociation indicates that a significant portion of the base remains undissociated.
- Strong bases typically dissociate nearly 100%, whereas weak bases dissociate less.

Relating to Acid-Base Strength

- The degree of dissociation at equilibrium reflects the strength of the base.
- Since only ~14.7% dissociates, this suggests a weak base.

Implications for Solution pH and Conductivity

- Higher $[OH^-]$ concentration correlates with higher pH.
- Conductivity of the solution depends on the ion concentration; more dissociation results in higher conductivity.

Factors Affecting Dissociation and pH

Various factors influence the degree of dissociation:

Temperature

- Temperature can shift the equilibrium position.
- The problem specifies 25°C, standard room temperature.

Concentration of the Base

- Higher initial concentrations may lead to different degrees of dissociation due to Le Chatelier's principle.

Nature of the Base

- Strong bases like NaOH dissociate completely.
- Weak bases like ammonia dissociate partially, as in this problem.

Additional Calculations and Considerations

To deepen understanding, consider related calculations:

Calculating the Equilibrium Concentration of the Undissociated Base

- Initial concentration: 0.43 M
- Dissociated: 0.0631 M
- Remaining undissociated base:
$$
0.43 - 0.0631 \approx 0.3669 \text{ M}
$$

Impact on pOH and pH

- Confirm the pH with the calculated [OH⁻]:
$$
\text{pOH} = -\log(0.0631) \approx 1.20
$$
- pH:
$$
14 - 1.20 = 12.80
$$
- Consistent with the problem statement, verifying calculations.

Summary and Key Takeaways

- The pH of 12.80 corresponds to a pOH of 1.20, which yields an $[OH^-]$ concentration of approximately 0.0631 M.
- The percent dissociation of the base is approximately 14.67%, indicating partial ionization typical of weak bases.
- Understanding the relationship between pH, pOH, and ion concentrations is critical for solving such problems.
- The calculations demonstrate how equilibrium concentrations inform us about the extent of dissociation and the strength of the base.

Conclusion

In conclusion, when a 0.43 M solution of a base at 25°C has a pH of 12.80 at equilibrium, approximately 14.67% of the base has dissociated. This insight is essential for chemists and students to understand the behavior of weak bases in aqueous solutions, their pH levels, and their degree of ionization. Mastery of these concepts enables accurate analysis of solution equilibria, informs laboratory practices, and supports the development of a deeper understanding of acid-base chemistry.

Additional Resources for Further Learning

- Textbooks on Acid-Base Equilibria
- Online tutorials on pH and pOH calculations
- Practice problems involving weak acids and bases
- Simulation tools for visualizing dissociation and equilibrium processes

This comprehensive overview should equip you with the knowledge and confidence to tackle similar problems involving pH, percent dissociation, and equilibrium in aqueous chemistry.

Frequently Asked Questions

What is the initial concentration of the base in the solution?

The initial concentration of the base is 0.43 M.

How do you determine the concentration of hydroxide ions (OH^-) from the pH value?

Since pH is given as 12.80, first calculate pOH = 14 - 12.80 = 1.20, then find $[OH^-]$ = 10^(-pOH) ≈ 10^(-1.20) ≈ 0.0631 M.

What is the significance of the pH being 12.80 at equilibrium?

A pH of 12.80 indicates a strongly basic solution, meaning the hydroxide ion concentration is high, and the base has dissociated significantly.

How do you find the percent dissociation of the base in the solution?

Percent dissociation = (concentration of dissociated base / initial concentration) × 100%. Using $[OH^-] \approx 0.0631$ M, for a base with molar mass M, calculate the dissociated amount and then the percent.

What assumptions are made when calculating the percent dissociation in this problem?

Assumptions include that the initial concentration is unchanged by dissociation except for the dissociated part, and the solution behaves ideally without other equilibria affecting the pH.

How does the pH relate to the extent of dissociation of the base?

A higher pH corresponds to greater hydroxide ion concentration, indicating a higher degree of dissociation for the base at equilibrium.

What is the approximate percent dissociation of the base based on the given data?

Calculating from $[OH^-] \approx 0.0631$ M and initial concentration of 0.43 M, percent dissociation ≈ (0.0631 / 0.43) × 100% ≈ 14.7%.

Additional Resources

If A 0.43 M Solution Of A Base 25°C Is Found To Have A pH Of 12.80 At Equilibrium, What Is The Percent?: An In-Depth Investigation into Acid-Base Equilibria and Concentration Calculations

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Introduction

Understanding the behavior of bases in aqueous solutions is foundational to chemistry, with applications spanning industrial processes, environmental science, and biochemistry. The problem statement—"If a 0.43 M solution of a base at 25°C is found to have a pH of 12.80 at equilibrium, what is the percent?"—serves as an intriguing case study for exploring the principles of acid-base equilibria, concentration calculations, and the interpretation of pH data. This investigation aims to dissect this problem with scientific rigor, providing clarity on the underlying concepts, methodology, and implications.

Background and Theoretical Framework

The Nature of Bases in Water

A base is a substance capable of accepting hydrogen ions (H^+) or donating hydroxide ions (OH^-). When dissolved in water, bases establish an equilibrium characterized by the dissociation process:

$$ \text{B} + \text{H}_2\text{O} \rightleftharpoons \text{BH}^+ + \text{OH}^- $$

or, for strong bases such as NaOH, complete dissociation occurs. However, many bases are weak and only partially dissociate, necessitating the use of equilibrium expressions to determine concentrations of ions.

pH and pOH Relationship

The pH of a solution measures the hydrogen ion concentration:

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

Conversely, pOH relates to hydroxide ions:

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

and these are connected via:

$$ \text{pH} + \text{pOH} = 14 \quad \text{at 25°C} $$

Given a pH of 12.80, the pOH can be calculated straightforwardly:

$$ \text{pOH} = 14 - 12.80 = 1.20 $$

which allows determination of hydroxide ion concentration.

Calculation of Hydroxide Ion Concentration

Using the pOH:

$$ [\text{OH}^-] = 10^{-\text{pOH}} = 10^{-1.20} $$

Calculating explicitly:

$$ [\text{OH}^-] = 10^{-1} \times 10^{-0.2} \approx 0.1 \times 10^{-0.2} $$

Recall that:

$$ 10^{-0.2} \approx 0.63 $$

Thus,

$$ [\text{OH}^-] \approx 0.1 \times 0.63 = 0.063 \, \text{M} $$

This indicates that at equilibrium, the hydroxide ion concentration is approximately 0.063 mol/L.

Determining the Extent of Dissociation

Given the initial concentration of the base (0.43 M), and the equilibrium hydroxide concentration (0.063 M), we can analyze the degree of dissociation.

The Dissociation Reaction

Assuming the base is weak and dissociates as:

$$ \text{B} \rightleftharpoons \text{OH}^- + \text{B}^+ $$

or a similar reaction, depending on the specific base. For simplicity, and because the problem refers generally to a "base," assume the dissociation directly produces hydroxide ions, and the initial concentration is the concentration of the base before dissociation.

Calculation of Percent Dissociation

The amount of base that dissociates to produce hydroxide ions is approximately equal to the hydroxide concentration at equilibrium:

$$ \text{Dissociated amount} = [\text{OH}^-] = 0.063 \, \text{M} $$

The percent dissociation is then:

$$ \% \text{Dissociation} = \frac{\text{Dissociated amount}}{\text{Initial concentration}} \times 100 $$

$$ \% \text{Dissociation} = \frac{0.063}{0.43} \times 100 \approx 14.65\% $$

This indicates that roughly 14.65% of the initial base molecules have dissociated at equilibrium.

Additional Considerations and Validations

Questioning the Assumption of Complete Dissociation

Given the pH and hydroxide concentration, it is evident that the base is weak, as the hydroxide concentration is significantly less than the initial molarity. If it were a strong base like NaOH, the hydroxide concentration would match the initial molarity closely, with minimal dissociation.

Equilibrium Constant (K_a or K_b)

The equilibrium constant for base dissociation (K_b) can be inferred from the hydroxide concentration:

$$K_b = \frac{[\text{OH}^-]^2}{[\text{B}]_{\text{initial}} - [\text{OH}^-]}$$

Plugging in the numbers:

$$K_b \approx \frac{(0.063)^2}{0.43 - 0.063} = \frac{0.00397}{0.367} \approx 0.0108$$

This value aligns with typical weak bases, indicating partial dissociation.

Implications and Broader Context

Practical Significance

Understanding the degree of dissociation and percent dissociation of bases in solution is critical in numerous applications:

- Buffer solution design: Knowing how much of a base dissociates informs buffer capacity.
- Industrial processes: Accurate pH control is essential in manufacturing and chemical synthesis.
- Environmental chemistry: The dissociation of alkaline substances influences water chemistry and pollutant behavior.

Limitations and Assumptions

- The calculations assume ideal behavior and neglect activity coefficients, which can influence ionic activity in concentrated solutions.
- The temperature is fixed at 25°C; deviations can alter ionization constants and pH values.
- The specific identity of the base affects dissociation behavior; here, a generic weak base assumption is made.

Conclusion

The investigation into "If a 0.43 M solution of a base at 25°C is found to have a pH of 12.80 at equilibrium, what is the percent?" reveals that approximately 14.65% of the initial base molecules dissociate under these conditions. This detailed analysis underscores the importance of pH and pOH calculations, the relationship between ion concentrations and dissociation, and the broader implications for chemical equilibria.

By systematically dissecting the problem, applying fundamental principles, and validating assumptions, we gain a comprehensive understanding of the

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behavior of weak bases in aqueous solutions. This knowledge not only
illuminates specific problem-solving techniques but also enhances our grasp
of the delicate balance governing acid-base chemistry—a core pillar of modern
science.

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