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Understanding how to calculate the pH of a solution created by mixing two acid solutions is essential in chemistry, especially in acid-base chemistry. This process involves several key steps: determining the moles of each component before mixing, understanding the reaction between acids, calculating the resulting concentration after mixing, and finally, computing the pH based on the concentration of hydrogen ions (H^+). In this article, we will thoroughly explore how to calculate the pH of a solution prepared by mixing 50 mL of a 0.10 M hydrofluoric acid (HF) solution with 25 mL of a 0.20 M solution, providing detailed explanations and relevant calculations.

Understanding the Components of the Mixture

Before diving into the calculation, it’s crucial to understand the properties of the solutions involved:

- Hydrofluoric Acid (HF): A weak acid with a known acid dissociation constant (Ka). HF is notable for its ability to partially dissociate in water, establishing an equilibrium between undissociated HF and dissociated ions (H^+ and F^-).

- Concentrations and Volumes:
- Solution 1: 50 mL of 0.10 M HF
- Solution 2: 25 mL of 0.20 M HF

Knowing these allows us to calculate the number of moles of HF in each solution, which is fundamental for the subsequent steps.

Calculating Moles of HF in Each Solution

The first step involves converting the given concentrations and volumes into moles.

Step 1: Moles in Solution 1

$$
\text{Moles of HF}_1 = \text{Concentration} \times \text{Volume}
$$

$$
= 0.10\, \text{mol/L} \times 50\, \text{mL} \times \frac{1\, \text{L}}{1000\, \text{mL}}
$$

$$
= 0.10 \times 0.050\, \text{L} = 0.005\, \text{mol}
$$

Step 2: Moles in Solution 2

$$
\text{Moles of HF}_2 = 0.20\, \text{mol/L} \times 25\, \text{mL} \times \frac{1\, \text{L}}{1000\, \text{mL}}
$$

$$
= 0.20 \times 0.025\, \text{L} = 0.005\, \text{mol}
$$

Observation: The moles of HF in both solutions are equal, each being 0.005 mol.

Determining the Total Moles of HF After Mixing

Since the moles of HF in both solutions are equal, upon mixing, the total moles of HF will be the sum:

$$
\text{Total moles of HF} = 0.005\, \text{mol} + 0.005\, \text{mol} = 0.010\, \text{mol}
$$

Next, we need to find the total volume of the mixture.

Calculating Total Volume of the Mixture

$$
V_{total} = 50\, \text{mL} + 25\, \text{mL} = 75\, \text{mL}
$$

Expressed in liters:

$$
V_{total} = 75\, \text{mL} \times \frac{1\, \text{L}}{1000\, \text{mL}} = 0.075\, \text{L}
$$

Calculating the Concentration of HF After Mixing

The initial moles of HF are known, and the total volume allows us to determine the new concentration:

$$
\text{Concentration of HF} = \frac{\text{Total moles of HF}}{\text{Total volume}}
$$

$$
= \frac{0.010\, \text{mol}}{0.075\, \text{L}} \approx 0.133\, \text{M}
$$

This is the initial concentration of HF in the mixture before considering dissociation.

Understanding HF Dissociation and Its Effect on pH

HF is a weak acid, meaning it does not fully dissociate in water. Its dissociation equilibrium can be written as:

$$
\mathrm{HF} \leftrightarrow \mathrm{H^{+}} + \mathrm{F^{-}}
$$

The acid dissociation constant (Ka) for HF is approximately:

$$
K_a \approx 6.6 \times 10^{-4}
$$

This value is critical for calculating the equilibrium concentration of hydrogen ions ($[H^+]$) in the solution, which directly determines the pH.

Setting Up the Equilibrium Expression

Let's denote:

- [HF] as the concentration of undissociated hydrofluoric acid at equilibrium
- $[H^+]$ as the concentration of hydrogen ions at equilibrium
- $[F^-]$ as the fluoride ion concentration

Assuming initial concentration of HF as $C_{initial} = 0.133\, \text{M}$, and the degree of dissociation is represented by x:

$$
\mathrm{HF} \leftrightarrow \mathrm{H^{+}} + \mathrm{F^{-}}
$$

- Initial: [HF] = 0.133 M; $[H^+]$ = 0; $[F^-]$ = 0
- Change: [HF] decreases by x; $[H^+]$ and $[F^-]$ increase by x
- Equilibrium: [HF] = $0.133 - x$, $[H^+]$ = x, $[F^-]$ = x

The equilibrium expression:

$$
K_a = \frac{[H^{+}][F^{-}]}{[HF]} = \frac{x \times x}{0.133 - x}
$$

Given that HF is a weak acid, x is small relative to initial concentration, allowing us to approximate:

$$
0.133 - x \approx 0.133
$$

Thus:

$$
K_a \approx \frac{x^2}{0.133}
$$

Solving for x:

$$
x^2 = K_a \times 0.133
$$

$$
x = \sqrt{K_a \times 0.133}
$$

Plugging in the values:

$$
x = \sqrt{6.6 \times 10^{-4} \times 0.133}
$$

$$
x = \sqrt{8.778 \times 10^{-5}} \approx 0.00937\, \text{M}
$$

This x value is the concentration of H^+ ions, which is also the pH:

$$
\text{pH} = -\log [H^{+}] = -\log 0.00937
$$

Calculating:

$$
\text{pH} \approx 2.03
$$

Final pH of the mixture: approximately 2.03.

Additional Considerations and Accuracy

While the above calculation provides a good estimate, some factors can influence the accuracy:

- Activity coefficients: In dilute solutions, activity coefficients slightly alter the effective concentration.
- Temperature: The Ka value is temperature-dependent; standard calculations assume room temperature (~25°C).
- Complete dissociation assumptions: For weak acids like HF, the approximation $x \ll C_{initial}$ is valid here. For higher concentrations, more precise quadratic solutions might be necessary.

Summary of the Calculation Steps

To summarize, the process involves:

1. Calculating the initial moles of HF in each solution.
2. Combining the moles and volumes to find the total concentration after mixing.

3. Recognizing HF as a weak acid and setting up the dissociation equilibrium.

4. Using the Ka value to approximate the degree of dissociation.

5. Calculating the hydrogen ion concentration.

6. Deriving the pH from the hydrogen ion concentration.

Conclusion

Calculating the pH of a solution formed by mixing two HF solutions involves understanding acid-base equilibria, molarity, and dissociation constants. In this scenario, mixing 50 mL of 0.10 M HF with 25 mL of 0.20 M HF results in a combined concentration of approximately 0.133 M HF, which, upon dissociation, yields a pH close to 2.03. This process exemplifies fundamental principles in acid chemistry and provides a practical approach to pH calculations in mixed solutions.

Additional Resources for Further Learning

- Understanding Acid Dissociation Constants (Ka)
- Calculating pH of Weak Acid Solutions
- Dilution and Molarity in Acid-Base Chemistry
- Common Weak Acids and Their Ka Values
- Practical Applications of pH Calculations in Industry and

Frequently Asked Questions

What is the process to calculate the pH of a solution formed by mixing

HF and another acid or base?

The process involves calculating the moles of each component, determining the resulting concentration after mixing, and then using the acid dissociation constant (Ka) to find the hydrogen ion concentration, which is used to compute the pH.

How do you determine the moles of HF in 50 mL of 0.10 M solution?

Moles of HF = concentration × volume = 0.10 mol/L × 0.050 L = 0.005 mol.

What is the total volume of the mixture when combining 50 mL of 0.10 M HF with 25 mL of another solution?

Total volume = 50 mL + 25 mL = 75 mL or 0.075 L.

How do you calculate the new concentration of HF after mixing?

New concentration = total moles of HF / total volume in liters = 0.005 mol / 0.075 L □ 0.0667 M.

If the second solution is NaOH instead of another acid, how does that affect the pH calculation?

NaOH is a strong base, so it will neutralize some of the HF, requiring a neutralization calculation to find the remaining acid or base, then calculating pH based on the resulting concentration.

Given that HF is a weak acid, how do you find the concentration of H+ ions in the mixture?

Use the acid dissociation expression for HF: Ka = [H+][F-]/[HF], and solve for [H+], considering the initial concentrations and the extent of dissociation.

What is the Ka value for HF, and how does it influence the pH calculation?

The Ka for HF is approximately 6.6 × 10□□, which indicates it's a weak acid; this value is used in equilibrium calculations to find [H+].

How do you perform an ICE table to determine the pH of the mixture?

Set up an ICE (Initial, Change, Equilibrium) table with initial concentrations, define the change due to dissociation, and solve for [H+], then calculate pH as -log[H+].

What assumptions are commonly made in calculating the pH of dilute weak acid solutions?

Assumptions often include that the degree of dissociation is small, so initial concentrations are approximately equal to equilibrium concentrations, simplifying calculations.

Additional Resources

Calculate The pH Of A Solution Prepared By Mixing 50 mL Of A 0.10 M Solution Of HF With 25 mL Of A 0.20 M Solution Of NaOH

Understanding the pH of a solution resulting from mixing acids and bases is fundamental in chemistry, especially in titration, buffer preparation, and various industrial processes. The specific problem of calculating the pH after combining a weak acid, hydrofluoric acid (HF), with a strong base, sodium hydroxide (NaOH), offers a comprehensive example of acid-base neutralization, equilibrium calculations, and solution chemistry. This article aims to provide a detailed, step-by-step analysis of the process, including theoretical foundations, calculation methods, and practical considerations.

Understanding the Components of the Mixture

Before delving into calculations, it's essential to understand the properties of the individual solutions involved:

- Hydrofluoric Acid (HF): A weak acid with a known acid dissociation constant (Ka). HF is characterized by its partial dissociation in water, and its behavior is governed by its Ka value. The given concentration is 0.10 M, with a volume of 50 mL.

- Sodium Hydroxide (NaOH): A strong base that dissociates completely in aqueous solution. The given concentration is 0.20 M, with a volume of 25 mL.

The goal is to determine the pH of the mixture after combining these two solutions.

Step 1: Determine the Moles of Each Reactant

The first step in any solution chemistry problem involving mixing is to convert concentrations and volumes into moles:

Moles of HF:

$$
\text{Moles of HF} = \text{Concentration} \times \text{Volume}
$$

$$
= 0.10\, \text{mol/L} \times 0.050\, \text{L} = 0.005\, \text{mol}
$$

Moles of NaOH:

$$
= 0.20\, \text{mol/L} \times 0.025\, \text{L} = 0.005\, \text{mol}
$$

Notice that both reactants are present in equal molar amounts, each with 0.005 mol.

Step 2: Analyze the Reaction Between HF and NaOH

The neutralization reaction between HF (weak acid) and NaOH (strong base) proceeds as:

$$
\text{HF} + \text{OH}^- \rightarrow \text{F}^- + \text{H}_2\text{O}
$$

Since NaOH is a strong base, it will react completely with HF present, consuming an equivalent amount of HF and producing fluoride ions (F^-).

Given that both reactants have equal molar amounts (0.005 mol), the reaction will:

- Completely consume all of the HF
- Consume an equivalent amount of NaOH (also 0.005 mol)
- Result in no remaining HF in solution
- Generate 0.005 mol of F^- ions

Remaining NaOH after reaction:

$$
0.005\, \text{mol} - 0.005\, \text{mol} = 0\, \text{mol}
$$

Thus, the NaOH is entirely consumed, and the solution now contains fluoride ions and water.

Step 3: Determine the Final Composition of the Solution

Total volume after mixing:

$$
V_{total} = 50\, \text{mL} + 25\, \text{mL} = 75\, \text{mL} = 0.075\, \text{L}
$$

Concentration of fluoride ions (F^-):

$$
[\text{F}^-] = \frac{\text{moles of F}^-}{V_{total}} = \frac{0.005\, \text{mol}}{0.075\, \text{L}} \approx 0.0667\, \text{M}
$$

Since there’s no excess acid or base remaining, the solution contains just fluoride ions and water.

Step 4: Understanding the Equilibrium of Fluoride in Water

The pH of the solution depends on the hydrolysis of fluoride ions because F^- is a weak base:

$$
\text{F}^- + \text{H}_2\text{O} \rightleftharpoons \text{HF} + \text{OH}^-
$$

The equilibrium constant for this process is the base hydrolysis constant, K_b, which relates to the Ka of HF:

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

where:

- $K_w = 1.0 \times 10^{-14}$ (ionization constant of water at 25°C)
- K_a for HF is approximately 6.6×10^{-4}

Calculating K_b:

$$
K_b = \frac{1.0 \times 10^{-14}}{6.6 \times 10^{-4}} \approx 1.52 \times 10^{-11}
$$

This small value indicates that F^- is a weak base, but it still affects the solution's pH.

Step 5: Calculating the pOH and pH

Using the hydrolysis equilibrium expression:

$$
K_b = \frac{[\text{OH}^-]^2}{[\text{F}^-]}
$$

Assuming $x = [\text{OH}^-]$ at equilibrium:

$$
K_b = \frac{x^2}{[\text{F}^-]}
$$

$$
x^2 = K_b \times [\text{F}^-]
$$

$$
x^2 = (1.52 \times 10^{-11}) \times (0.0667)
$$

$$
x^2 \approx 1.015 \times 10^{-12}
$$

$$
x \approx \sqrt{1.015 \times 10^{-12}} \approx 1.01 \times 10^{-6}\, \text{M}
$$

This is the concentration of hydroxide ions:

$$
[\text{OH}^-] \approx 1.01 \times 10^{-6}\, \text{M}
$$

Now, find the pOH:

$$
\text{pOH} = -\log[\text{OH}^-] \approx -\log(1.01 \times 10^{-6}) \approx 5.99
$$

And the pH:

$$
\text{pH} = 14 - \text{pOH} \approx 14 - 5.99 \approx 8.01
$$

Therefore, the pH of the mixture is approximately 8.01, indicating a slightly basic solution.

Additional Considerations and Practical Implications

Calculating the pH of mixed solutions involves understanding acid-base equilibria, molar calculations, and equilibrium constants. Several factors influence the accuracy and relevance of these calculations:

- Assumption of complete reaction: The calculations assume that NaOH fully reacts with HF, which is valid given the molar equivalence.
- Equilibrium constants: Use of accurate K_a and K_b values is essential; these can vary slightly depending on source and temperature.
- Dilution effects: Final concentrations depend on precise volume measurements; minor errors can influence pH calculations.
- Hydrolysis of F^- ions: The weak base behavior of fluoride ions significantly influences the pH, especially in dilute solutions.

Pros and Cons of This Calculation Approach

Pros:

- Provides a clear understanding of acid-base neutralization and equilibrium principles.
- Demonstrates the step-by-step process for calculating pH after mixing solutions.
- Highlights the importance of molar ratios and solution volumes in chemical calculations.
- Uses fundamental constants and equilibrium expressions applicable to similar problems.

Cons:

- Assumes ideal behavior; real-world solutions may have impurities or activity effects.
- Relies on approximate K_a values, which can vary with temperature.
- Does not account for complex equilibria that may occur in real systems.
- Simplifies the hydrolysis reaction as a single-step process, ignoring potential secondary equilibria.

Features and Key Takeaways

- Understanding neutralization: Complete reaction between a weak acid and strong base results in a solution dominated by the conjugate base.
- Weak base hydrolysis: Fluoride ions hydrolyze in water, influencing the pH towards basicity.
- Volume and molarity calculations: Critical for accurate determination of concentrations after mixing.
- Equilibrium constants: Fundamental in predicting solution pH, especially when strong acids or bases are involved.

Conclusion

Calculating the pH of a solution formed by mixing hydrofluoric acid and sodium hydroxide involves understanding the stoichiometry of the neutralization reaction, the resultant ion concentrations, and the hydrolysis behavior of the conjugate base. In this example, since equivalent molar amounts of HF and NaOH are mixed, and all acid is neutralized, the pH is determined primarily by the hydro

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