Calculate The Concentration Of H+ Ions In A 0.010 M Aqueous Solution Of Sulfuric Acid. H2SO4 (aq)<===>

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Calculate The Concentration Of H+ Ions In A 0.010 M Aqueous Solution Of Sulfuric Acid. H2SO4 (aq) <===> is a fundamental question in acid-base chemistry that involves understanding the dissociation behavior of sulfuric acid (H2SO4) in water, the concept of strong acids, and the ionization process. Since sulfuric acid is a diprotic acid, it can release two protons (H+) per molecule, but its actual dissociation process in aqueous solution is more complex than that of monoprotic acids like hydrochloric acid (HCl). This article explores the dissociation mechanism of sulfuric acid, calculations involved, and the factors influencing the concentration of hydrogen ions (H+), leading to a comprehensive understanding of this important chemical process.

Understanding Sulfuric Acid (H2SO4) and Its Dissociation in Water

Nature of Sulfuric Acid

- Sulfuric acid (H2SO4) is a strong mineral acid widely used in industry, laboratories, and manufacturing processes.
- It is a diprotic acid, meaning each molecule can release two protons (H+ ions) when dissociating in water.
- It has a high degree of ionization, especially for the first proton, making it a strong acid.

Stages of Dissociation in Water

The dissociation of sulfuric acid occurs in two primary steps:

1. **First dissociation:**
   H_2SO_4 (aq) <===> H^+ (aq) + HSO_4^- (aq)

   - This step is complete in dilute solutions, meaning virtually all H_2SO_4 molecules donate their first proton.
   - The first dissociation is strong and essentially irreversible under typical conditions.

2. **Second dissociation:**
   HSO_4^- (aq) <===> H^+ (aq) + SO_4^{2-} (aq)

   - This step is partial; the equilibrium position depends on the concentration and conditions.
   - The second dissociation is weaker compared to the first, but still significant in dilute solutions.

Implications for Concentration of H+ Ions in 0.010 M Solution

First Dissociation: Complete Ionization

In dilute solutions like 0.010 M, the first dissociation of sulfuric acid is considered to be complete. This means:

- All H_2SO_4 molecules release their first proton.
- The concentration of H^+ ions contributed by this step is approximately equal to the initial concentration of the acid, i.e., 0.010 M.

Second Dissociation: Partial Ionization

The second dissociation is not complete and reaches an equilibrium that must be calculated. The degree of ionization depends on the acid dissociation constant (Ka) for the second step, which is smaller than for the first.

Calculating the Concentration of H+ Ions

Step 1: Contribution from the First Dissociation

- Since the first dissociation is complete, it provides 0.010 M of H+ ions directly.

Step 2: Contribution from the Second Dissociation

To find the additional H+ ions contributed by the second dissociation, we need to consider the equilibrium of the second step:

HSO_4^- (aq) <===> H+ (aq) + SO_4^{2-} (aq)

Let's define:

- [HSO_4^-] = initial concentration after first dissociation = 0.010 M
- [H+] = x, the amount of H+ produced from the second dissociation
- [SO_4^{2-}] = x
- [HSO_4^-] at equilibrium = 0.010 - x

Step 3: Acid Dissociation Constant (Ka) for the Second Dissociation

The value of Ka for the second dissociation of H_2SO_4 is approximately:

- $K_{a2} \approx 1.2 \times 10^{-2}$

This value indicates the degree of ionization of HSO_4^- into SO_4^{2-} and H+ ions in dilute solutions.

Step 4: Setting Up the Equilibrium Expression

The equilibrium expression based on the second dissociation is:

$$Ka2 = [H+][SO4^{2-}] / [HSO4-]$$

Substituting the knowns:

$$1.2 \times 10^{-2} = (x)(x) / (0.010 - x)$$

Step 5: Solving for x

Assuming x is small relative to 0.010 M (which is valid here), we approximate:

$$0.010 - x \approx 0.010$$

Thus, the equation simplifies to:

$$1.2 \times 10^{-2} \approx x^2 / 0.010$$

Rearranged:

$$x^2 = (1.2 \times 10^{-2}) \times 0.010 = 1.2 \times 10^{-4}$$

Taking the square root:

$$x = \sqrt{(1.2 \times 10^{-4})} \approx 0.01095 \text{ M}$$

Step 6: Total H+ Concentration

The total concentration of H+ ions is the sum from both dissociation steps:

- From first dissociation: 0.010 M
- From second dissociation: approximately 0.01095 M

Therefore:

$$[H+] \approx 0.010 + 0.01095 \approx 0.02095 \text{ M}$$

Rounding appropriately, the concentration of H+ ions in the solution is approximately 0.021 M.

Summary and Final Remarks

Key Points Recap

- Sulfuric acid is a strong, diprotic acid with the first dissociation being complete in dilute solutions.
- The second dissociation is partial and can be calculated using the equilibrium constant Ka2.
- In a 0.010 M solution, the total H+ concentration is approximately 0.021 M, considering both dissociation steps.

Implications for Acid-Base Chemistry

This calculation illustrates how the strength and dissociation behavior of acids influence the concentration of hydrogen ions, which in turn affects pH and other properties of the solution. Understanding these principles is crucial for applications in chemical analysis, titrations, and industrial processes.

Conclusion

Calculating the concentration of H+ ions in an aqueous sulfuric acid solution requires understanding its dissociation as a diprotic acid, recognizing the difference in the extent of dissociation for each proton, and applying equilibrium constants to quantify the ionization. For a 0.010 M solution of sulfuric acid, the H+ concentration is approximately 0.021 M, reflecting the combined effects of the complete first dissociation and the partial second dissociation. Mastery of these concepts enables chemists to predict acid strengths, calculate pH, and manipulate solutions for various chemical applications.

Frequently Asked Questions

What is the initial concentration of H2SO4 in the solution?

The initial concentration of H2SO4 is 0.010 M.

How does sulfuric acid dissociate in aqueous solution?

Sulfuric acid dissociates in two steps: first, it fully dissociates into H+ and HSO4-, and second, HSO4- partially dissociates into H+ and SO4^{2-}.

What is the contribution of each dissociation step to the H+ ion concentration?

The first dissociation produces 0.010 M of H+ ions, while the second dissociation adds additional H+ ions depending on the extent of ionization of HSO4-.

Why is sulfuric acid considered a strong acid in its first dissociation step?

Because it completely dissociates into H+ and HSO4- in aqueous solution, making the first step a strong acid dissociation.

How do you calculate the concentration of H+ ions from sulfuric acid solution?

Calculate the contribution from both dissociation steps: the first contributes 0.010 M H+, and the second dissociation adds more H+ depending on equilibrium, which can be estimated using the acid dissociation constant (Ka).

What is the approximate pH of a 0.010 M sulfuric acid solution?

Since the first dissociation fully contributes 0.010 M H+ ions, the pH is approximately -log(0.010) ≈ 2.00.

Is the contribution of the second dissociation significant in calculating H+ concentration?

Yes, because HSO4- is a weak acid with a small Ka, it contributes additional H+ ions, slightly lowering the pH further.

How do you account for the second dissociation in calculations?

Use the Ka value for HSO4- to set up an equilibrium expression and solve for the additional H+ ions produced.

What is the final concentration of H+ ions in the solution?

The total H+ concentration is approximately 0.010 M from the first dissociation plus a small amount from the second dissociation, resulting in roughly 0.0102 M.

What is the resulting pH of the solution considering both dissociation

steps?

The pH is approximately -log(0.0102) ≈ 1.99, reflecting the strong initial dissociation and partial second dissociation.

Additional Resources

Calculate The Concentration Of H+ Ions In A 0.010 M Aqueous Solution Of Sulfuric Acid

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Introduction

Sulfuric acid (H_2SO_4) is one of the most widely used industrial chemicals and is renowned for its strong acidic properties. Its extensive applications range from fertilizer manufacturing to chemical synthesis, making understanding its behavior in aqueous solutions essential for both industrial processes and academic research. A fundamental aspect of this understanding involves calculating the concentration of hydrogen ions (H^+) present when sulfuric acid dissolves in water.

This article provides a comprehensive, investigative analysis of how to calculate the concentration of H^+ ions in a 0.010 M aqueous solution of sulfuric acid, examining the underlying chemistry, dissociation mechanisms, and the factors influencing ion concentrations, culminating in a detailed step-by-step calculation.

Background: The Chemistry of Sulfuric Acid in Water

Molecular Structure and Acidic Strength

Sulfuric acid is a diprotic acid, meaning it can donate two protons (H^+) per molecule during dissociation:

$$\mathrm{H_2SO_4} \leftrightarrow 2\, \mathrm{H^+} + \mathrm{SO_4^{2-}}$$

Its strong initial dissociation makes it a typical strong acid in the first dissociation step, whereas the second dissociation is partial, characterized by an equilibrium:

$$\mathrm{HSO_4^-} \leftrightarrow \mathrm{H^+} + \mathrm{SO_4^{2-}}$$

This dual nature influences the overall concentration of H^+ ions in solution.

Dissociation Phases

1. First dissociation (complete):

$$ \mathrm{H_2SO_4} \rightarrow \mathrm{H^+} + \mathrm{HSO_4^-} $$

2. Second dissociation (partial):

$$ \mathrm{HSO_4^-} \leftrightarrow \mathrm{H^+} + \mathrm{SO_4^{2-}} $$

The extent of the second dissociation is governed by its acid dissociation constant, K_{a2}, which is approximately 1.2×10^{-2}.

Investigative Approach to Calculating H^+ Concentration

Step 1: Recognize the Dissociation Behavior

Given the strong nature of the first dissociation, we can assume that all H_2SO_4 molecules dissociate completely in aqueous solution:

$$ \text{[H_2SO_4]}_{\text{initial}} = 0.010\, \mathrm{M} $$
$$ \Rightarrow \text{[H^+]}_{\text{from first dissociation}} \approx 0.010\, \mathrm{M} $$

However, the second dissociation is only partial, contributing additional H^+ ions. To accurately determine the total H^+ concentration, we need to account for this second dissociation.

Step 2: Establish the Equilibrium Expression for the Second Dissociation

The second dissociation can be represented as:

$$ \mathrm{HSO_4^-} \leftrightarrow \mathrm{H^+} + \mathrm{SO_4^{2-}} $$

with the equilibrium constant:

$$ K_{a2} = \frac{[\mathrm{H^+}][\mathrm{SO_4^{2-}}]}{[\mathrm{HSO_4^-}]} $$

Given the initial concentration of H_2SO_4, and assuming complete dissociation of the first hydrogen, the initial concentrations after the first dissociation are:

- $[\mathrm{H^+}] \approx 0.010\, \mathrm{M}$
- $[\mathrm{HSO_4^-}] \approx 0.010\, \mathrm{M}$

- $[\mathrm{SO_4^{2-}}] \approx 0$

The amount of $\mathrm{HSO_4^-}$ that dissociates further is denoted as x. Therefore:

- $[\mathrm{H^+}] = 0.010 + x$
- $[\mathrm{SO_4^{2-}}] = x$
- $[\mathrm{HSO_4^-}] = 0.010 - x$

Using the K_{a2} expression:

$$ K_{a2} = \frac{(0.010 + x) \times x}{0.010 - x} $$

Given that $K_{a2} = 1.2 \times 10^{-2}$, we proceed to solve for x.

Step 3: Solving for x, the Degree of Second Dissociation

The equation simplifies to:

$$ 1.2 \times 10^{-2} = \frac{(0.010 + x) \times x}{0.010 - x} $$

Since K_{a2} is small relative to the initial concentrations, and x is expected to be much less than 0.010, we can approximate:

$$ 0.010 + x \approx 0.010 $$
$$ 0.010 - x \approx 0.010 $$

This simplifies our calculation to:

$$ 1.2 \times 10^{-2} \approx \frac{0.010 \times x}{0.010} = x $$

Therefore:

$$ x \approx 1.2 \times 10^{-2} \, \mathrm{M} $$

But this is inconsistent with the assumption $x \ll 0.010$. To check this, consider that $x \approx 0.012\, \mathrm{M}$ is greater than 0.010 M, which is not possible because the maximum $[\mathrm{HSO_4^-}]$ after dissociation cannot become negative.

Hence, the assumption $x \ll 0.010$ is invalid here, indicating that the second dissociation proceeds significantly, and a more precise calculation using the quadratic formula is needed.

Step 4: Precise Calculation Using Quadratic Equation

Rearranged:

$$ 1.2 \times 10^{-2} = \frac{(0.010 + x) \times x}{0.010 - x} $$

Multiply both sides by $(0.010 - x)$:

$$ 1.2 \times 10^{-2} \times (0.010 - x) = (0.010 + x) \times x $$

Expanding:

$$ 1.2 \times 10^{-2} \times 0.010 - 1.2 \times 10^{-2} \times x = 0.010 x + x^2 $$

Calculating:

$$ 1.2 \times 10^{-2} \times 0.010 = 1.2 \times 10^{-4} $$

So,

$$ 1.2 \times 10^{-4} - 1.2 \times 10^{-2} x = 0.010 x + x^2 $$

Bring all to one side:

$$ 1.2 \times 10^{-4} - 1.2 \times 10^{-2} x - 0.010 x - x^2 = 0 $$

Combine like terms:

$$ 1.2 \times 10^{-4} - (1.2 \times 10^{-2} + 0.010) x - x^2 = 0 $$

Note that $1.2 \times 10^{-2} = 0.012$:

$$ 1.2 \times 10^{-4} - (0.012 + 0.010) x - x^2 = 0 $$
$$ 1.2 \times 10^{-4} - 0.022 x - x^2 = 0 $$

Rewrite as:

$$ x^2 + 0.022 x - 1.2 \times 10^{-4} = 0 $$

Applying the quadratic formula:

$$ x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} $$

where:

- $a = 1$
- $b = 0.022$
- $c = -1.2 \times 10^{-4}$

Calculating discriminant:

$$ D = (0.022)^2 - 4 \times 1 \times (-1.2 \times 10^{-4}) $$
$$ D = 0.000484 + 4.8 \times 10^{-4} $$
$$ D = 0.000484 + 0.00048 = 0.000964 $$

Square root:

$$ \sqrt{D} \approx 0.03105 $$

Calculate x:

$$ x = \frac{-0.022 \pm 0.03105}{2} $$

Two solutions:

1. $ x = \frac{-0.022 + 0.03105}{2} = \frac{0.00905}{2} = 0.004525\, \mathrm{M} $
2. $ x = \frac{-0.022 - 0.03105}{2} = \frac{-0.05305}{2} = -0.026525\, \mathrm{M} $ (discarded as

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