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Introduction to Solubility and the Role of Ksp

Understanding the solubility of ionic compounds like zinc hydroxide ($Zn(OH)_2$) in various solutions is fundamental in chemistry, particularly in inorganic chemistry and solution chemistry. Solubility determines how much of a substance can dissolve in a solvent at a given temperature, influencing processes in industrial applications, environmental chemistry, and biological systems.

The solubility product constant, Ksp, is a key parameter that quantifies the solubility of a sparingly soluble compound. It is the equilibrium constant for the dissolution process, representing the maximum product of the ion concentrations in a saturated solution at a specific temperature.

In this article, we will explore how to calculate the solubility of zinc hydroxide in a 1.00 M sodium hydroxide (NaOH) solution, given its Ksp value of 3.0 x 10^-16. This calculation involves understanding complex equilibria, the effect of common ions, and the role of hydroxide ions in solubility enhancement.

Understanding Zinc Hydroxide ($Zn(OH)_2$) and Its Dissolution Equilibrium

Dissolution Reaction of $Zn(OH)_2$

Zinc hydroxide dissolves in water according to the following equilibrium:

$$ \text{Zn(OH)}_2 (s) \rightleftharpoons \text{Zn}^{2+} (aq) + 2 \text{OH}^- (aq) $$

The solubility product expression (Ksp) for this equilibrium is:

$$ K_{sp} = [\text{Zn}^{2+}] [\text{OH}^-]^2 $$

Given:

$$ K_{sp} = 3.0 \times 10^{-16} $$

The challenge in this problem arises because the solution contains 1.00 M NaOH, which supplies a high concentration of hydroxide ions (OH^-). This influences the solubility of $Zn(OH)_2$ due to the common ion effect and potential formation of complexes.

Effect of NaOH on Zinc Hydroxide Solubility

Common Ion Effect

Since NaOH is a strong base that fully dissociates, it contributes 1.00 M of hydroxide ions:

$$ \text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^- $$

This high background concentration of OH^- ions suppresses the dissolution of $Zn(OH)_2$ because the equilibrium shifts toward the solid form, decreasing solubility.

Formation of Zinc Hydroxide Complexes

In strongly basic solutions, zinc can form soluble complexes such as:

- Zn(OH)_3^-
- Zn(OH)_4^{2-}

The formation of these complexes increases the apparent solubility of zinc hydroxide. The extent of complex formation depends on the stability constants of these complexes.

Step-by-Step Calculation Method

To accurately determine the solubility of Zn(OH)_2 in 1.00 M NaOH, we need to:

1. Recognize the initial hydroxide concentration from NaOH.
2. Consider the potential complex formation between zinc ions and hydroxide ions.
3. Write equilibrium expressions for complex formation.
4. Incorporate the complex equilibria into the solubility expression.
5. Solve for the molar solubility of zinc hydroxide (s).

Assumptions and Simplifications

- The solution is at equilibrium.
- The temperature is constant (assumed room temperature unless specified).
- The activity coefficients are approximated to 1, i.e., ideal solution behavior.
- The complex formation constants are known or can be estimated.

Step 1: Establishing the Basic Equilibrium

Let's define:

- s = molar solubility of $Zn(OH)_2$ (mol/L), i.e., the molar amount of $Zn(OH)_2$ that dissolves to reach equilibrium.

From the dissolution:

$$\text{Zn(OH)}_2 (s) \rightarrow \text{Zn}^{2+} + 2 \text{OH}^-$$

- The concentration of Zn^{2+} is s.
- The free OH^- contributed by dissolution is $2s$.

However, because NaOH provides 1.00 M OH^-, the total hydroxide concentration at equilibrium becomes:

$$[\text{OH}^-]_{total} = 1.00\, \text{M} + 2s$$

Step 2: Considering Complex Formation

Zinc hydroxide tends to form complexes in basic solutions. The most relevant are:

- Zn(OH)_3^- with formation constant K_{f1}
- Zn(OH)_4^{2-} with formation constant K_{f2}

Suppose the dominant complex is Zn(OH)_4^{2-}. The complex formation equilibrium can be written as:

$$ \text{Zn}^{2+} + 4 \text{OH}^- \rightleftharpoons \text{Zn(OH)}_4^{2-} $$

with formation constant:

$$ K_{f} = \frac{[\text{Zn(OH)}_4^{2-}]}{[\text{Zn}^{2+}][\text{OH}^-]^4} $$

Values of K_{f} are typically high (e.g., $K_{f} \approx 10^{17}$), indicating strong complex formation.

The total zinc in solution (free + complexed) relates to the solubility:

$$ s_{total} = [\text{Zn}^{2+}] + [\text{Zn(OH)}_3^-] + [\text{Zn(OH)}_4^{2-}] $$

Given the high stability of the Zn(OH)_4^{2-} complex, most zinc might be in this form, especially in strongly basic medium.

Step 3: Formulating the Overall Equilibrium

The total zinc concentration corresponds to the molar solubility s, which includes free zinc ions and complexes. The dominant complex formation shifts zinc into the soluble complex form, effectively increasing zinc's apparent solubility.

The key is to relate the dissolution equilibrium with the complex formation:

$$ K_{sp} = [\text{Zn}^{2+}][\text{OH}^-]^2 $$

But with complex formation, the free zinc ion concentration is reduced, and the total zinc concentration is increased due to complexation.

Step 4: Approximating the Solubility

Given the high K_{f}, most zinc is in the form of Zn(OH)_4^{2-}. The approximate relation is:

$$ s_{total} \approx [\text{Zn(OH)}_4^{2-}] $$

and

$$ [\text{Zn}^{2+}] \approx \frac{s_{total}}{K_{f} [\text{OH}^-]^4} $$

The total hydroxide concentration is approximately 1.00 M (from NaOH) plus a small contribution from the dissolved zinc. Since the dissolved zinc amount is tiny compared to 1 M, we can approximate:

$$ [\text{OH}^-] \approx 1.00\, \text{M} $$

Step 5: Final Calculation of Solubility

Using the equilibrium expression:

$$ K_{sp} = [\text{Zn}^{2+}] [\text{OH}^-]^2 $$

and assuming that:

$$ [\text{Zn}^{2+}] = \frac{s}{\text{complexation factor}} $$

but in the case of strong complexation, the apparent solubility $s_{apparent}$ is:

$$ s_{apparent} \approx \text{(solubility of Zn(OH)}_2\text{ in presence of complexation)} $$

which can be calculated via:

$$ s_{apparent} \approx \frac{K_{f} \times K_{sp}}{[\text{OH}^-]^4} $$

Inserting known values:

$$ s_{apparent} \approx \frac{10^{17} \times 3.0 \times 10^{-16}}{(1.00)^4} $$

$$ s_{apparent} \approx 10^{17} \times 3.0 \times 10^{-16} $$

$$ s_{apparent} \approx 30\, \text{M} $$

This indicates that in the presence of a very high complex formation constant, zinc hydroxide's solubility can be significantly increased, potentially reaching tens of molar concentrations due to complexation.

Summary of Results and Implications

- Without considering complexation, the solubility of $Zn(OH)_2$ in pure water is extremely low, approximately:

$$ s = \sqrt{\frac{K_{sp}}{[\text{OH}^-]^2}} $$

- In 1.00 M

Frequently Asked Questions

What is the solubility of zinc hydroxide ($Zn(OH)_2$) in 1.00 M NaOH solution?

The solubility of $Zn(OH)_2$ in 1.00 M NaOH is approximately 1.26×10^{-4} mol/L.

How does the presence of 1.00 M NaOH affect the solubility of $Zn(OH)_2$?

The high concentration of NaOH suppresses the solubility of $Zn(OH)_2$ due to common ion effect, but using Ksp and OH^- concentration, we can calculate its solubility precisely.

What is the dissociation equation for $Zn(OH)_2$ in aqueous solution?

$Zn(OH)_2(s) \rightleftharpoons Zn^{2+}(aq) + 2OH^-(aq)$.

Given Ksp = 3.0×10^{-16} for $Zn(OH)_2$, how do we set up the solubility expression in 1.00 M NaOH?

Assuming solubility s (mol/L), $[OH^-] = 2s + 1.00$ M (from the added NaOH). The Ksp expression becomes $K_{sp} = [Zn^{2+}][OH^-]^2 = s \times (2s + 1.00)^2$.

How do you solve for the solubility of $Zn(OH)_2$ in 1.00 M NaOH?

By substituting $[OH^-] = 2s + 1.00$ M into the Ksp expression and solving the quadratic equation for s, the approximate solubility is 1.26×10^{-4} mol/L.

Why is the solubility of $Zn(OH)_2$ lower in basic solutions like NaOH?

Because the excess OH^- ions shift the equilibrium according to Le Châtelier's principle, reducing the dissolution of $Zn(OH)_2$ due to the common ion effect.

Can you approximate the [OH^-] concentration in the solution when calculating solubility?

Yes, since NaOH is 1.00 M and the dissolution of Zn(OH)2 adds only a small amount of OH^-, we approximate [OH^-] as mainly from NaOH, i.e., 1.00 M, for simplicity.

What assumptions are made in calculating the solubility of Zn(OH)2 in this scenario?

Assumptions include ignoring the small contribution of OH^- from Zn(OH)2 dissolution compared to the 1.00 M NaOH, and treating the system as at equilibrium with constant [OH^-] from NaOH.

How does the Ksp value influence the solubility calculation of Zn(OH)2?

The Ksp value provides the equilibrium constant that relates the concentrations of ions in solution, allowing calculation of solubility based on known ion concentrations and the dissociation equilibrium.

Additional Resources

Calculate The Solubility Of Zinc Hydroxide, Zn(OH)2 In 1.00 M NaOH

Introduction

The solubility of zinc hydroxide ($Zn(OH)_2$) in aqueous solutions is a classic topic in inorganic chemistry, pivotal for understanding solubility equilibria, complex formation, and the influence of common ions. In particular, the solubility behavior of $Zn(OH)_2$ in highly basic media such as sodium hydroxide (NaOH) solutions provides insight into the stability of zinc hydroxide and related zinc complexes.

This article aims to thoroughly investigate the calculation of $Zn(OH)_2$ solubility in a 1.00 M NaOH solution, given its solubility product constant (Ksp) of $3.0 \times 10^{□1□}$. The analysis involves considering the common ion effect, potential complex formation, and the application of equilibrium principles. The goal is to derive a quantitative value for the solubility of $Zn(OH)_2$ under these conditions, which is essential for applications in analytical chemistry, environmental science, and materials chemistry.

Background and Theoretical Foundations

Solubility Product Constant (Ksp)

The solubility product constant (Ksp) quantifies the solubility equilibrium of a sparingly soluble salt. For $Zn(OH)_2$, the dissociation in water can be represented as:

$Zn(OH)_2(s) \rightleftharpoons Zn^{2+}(aq) + 2OH^-(aq)$

The Ksp expression is:

$Ksp = [Zn^{2+}][OH^-]^2$

Given that $Ksp = 3.0 \times 10^{□1□}$, this value sets an initial measure of the maximum solubility of $Zn(OH)_2$ in pure water or dilute solutions without complexation.

The Effect of Common Ions: NaOH

In a 1.00 M NaOH solution, the hydroxide ion concentration ($[OH^-]$) is significantly higher than in pure water. The common ion effect states that an increase in the concentration of one of the ions involved in a solubility equilibrium shifts the equilibrium toward the solid form, decreasing solubility.

Complex Formation in Basic Media

Zinc ions are known to form soluble complexes with hydroxide ions, notably the tetrahydroxy zincate complex:

$Zn(OH)_4^{2-}$

The formation of such complexes effectively increases zinc's solubility in basic solutions by stabilizing Zn^{2+} ions in solution. The formation equilibrium is:

$Zn^{2+} + 4OH^- \rightleftharpoons [Zn(OH)_4]^{2-}$

The stability of these complexes is characterized by stability constants (K_f). The formation of zincate complexes reduces free Zn^{2+} concentration, thus impacting the solubility of $Zn(OH)_2$.

Step-by-Step Calculation Approach

1. Establish the Baseline Solubility in Pure Water

In pure water, the solubility (s) of $Zn(OH)_2$ is related to Ksp:

$K_{sp} = s \times (2s)^2 = 4s^3$

Therefore,

$s = (K_{sp} / 4)^{(1/3)}$

Plugging in the known Ksp:

s = (3.0 × 10□¹□ / 4)^(1/3)

s = (7.5 × 10□¹□)^(1/3)

s □ (7.5 × 10□¹□)^(1/3)

Calculating cube root:

s □ 4.2 × 10□□ M

This is the solubility in pure water, with hydroxide concentration approximately equal to 2s (from the dissociation):

[OH□] □ 2 × 4.2 × 10□□ □ 8.4 × 10□□ M

which is negligible compared to 1.00 M NaOH.

2. Impact of 1.00 M NaOH – The Common Ion Effect

In a 1.00 M NaOH solution, the hydroxide ion concentration is:

[OH□] = 1.00 M (assuming complete dissociation)

The high hydroxide concentration suppresses Zn(OH)□ dissolution per Le Chatelier's principle, decreasing the free Zn²□ concentration.

Using the Ksp expression:

[Zn²□] = Ksp / [OH□]²

□ [Zn²□] = (3.0 × 10□¹□) / (1.00)² = 3.0 × 10□¹□ M

This indicates that, in the absence of complex formation, the zinc ion concentration is extremely low, and the solubility s is approximately equal to $[Zn^{2+}]$, which is approximately $3.0 \times 10^{-1\square}$ M.

3. Considering Complex Formation: The Tetrahydroxy Zincate

Given the basic environment, zinc can form the complex ion $[Zn(OH)_4]^{2-}$, which significantly enhances zinc's apparent solubility. The formation reaction is:

$$Zn^{2+} + 4OH^- \rightleftharpoons [Zn(OH)_4]^{2-}$$

with a formation constant:

$$Kf = [Zn(OH)_4]^{2-} / ([Zn^{2+}][OH^-]^4)$$

The stability constant (Kf) for zincate complexes varies depending on conditions, but typical values are in the range of $10^{\square}$ to $10^{1\square}$.

Assuming:

$$Kf \approx 10^{1\square}$$

and the total zinc in solution (S) is partitioned between free Zn^{2+} and complexed zinc:

$$S = [Zn^{2+}] + [Zn(OH)_4]^{2-}$$

Given the high stability of the complex, almost all zinc will be complexed, so:

$$[Zn(OH)_4]^{2-} \approx S$$

and

$[Zn^{2+}] \approx S / K_f \times [OH^-]^4$

Rearranged:

$S = [Zn(OH)_4]^{2-} \approx K_f \times [Zn^{2+}] \times [OH^-]^4$

But since the complex stabilizes zinc, the effective solubility of $Zn(OH)_2$ is greatly enhanced compared to the simple Ksp calculation.

4. Quantitative Calculation of Zinc Hydroxide Solubility in 1.00 M NaOH

Assuming that the predominant zinc species is the complex ion $[Zn(OH)_4]^{2-}$, and that the free Zn^{2+} concentration is very low, we can approximate the total zinc concentration (S) as:

$S \approx [Zn(OH)_4]^{2-}$

From the complex formation:

$[Zn^{2+}] = S / (K_f \times [OH^-]^4)$

Substituting:

$S \approx K_f \times [Zn^{2+}] \times [OH^-]^4$

But since $[Zn^{2+}]$ is very small and in equilibrium with the solid $Zn(OH)_2$, the solubility of $Zn(OH)_2$ is essentially governed by the formation of the complex.

Alternatively, considering the equilibrium:

$Zn(OH)_2(s) \rightleftharpoons Zn^{2+} + 2OH^-$

and the complex formation:

$$Zn^{2+} + 4OH^{-} \rightleftharpoons [Zn(OH)_4]^{2-}$$

The overall process can be combined to give an effective solubility expression:

$$S\_total \approx S_0 + S\_complex$$

where S_0 is the solubility in pure water (~$4.2 \times 10^{[illegible]}$ M), and S_complex accounts for zinc bound in complexes.

Given the high stability of the complex, the dominant zinc species in 1.00 M NaOH is $[Zn(OH)_4]^{2-}$, and the free Zn^{2+} concentration is extremely low.

Using the formation constant:

$$K_f \approx 10^{1[illegible]}$$

and $[OH^{-}] = 1.00$ M,

$$S \approx [Zn(OH)_4]^{2-} \approx (K_f) \times [Zn^{2+}] \times (1.00)^4$$

Rearranged:

$$[Zn^{2+}] \approx S / (K_f \times 1.00^4)$$

But also, from the solubility of $Zn(OH)_2$:

$$[Zn^{2+}] \approx s$$

and

$Ksp = s \times ([OH\square])^2$

which gives:

$s = Ksp / [OH\square]^2 = 3.0 \times 10\square^{1}\square / 1.00^2 = 3.0 \times 10\square^{1}\square$ M

This is negligible compared to the complexed zinc concentration.

Therefore, the major zinc concentration is in the complexed form, and the total solubility S_total is approximately:

$S\_total \ \square\ [Zn(OH)\square]^2\square\ \square\ S\_complex$

which can be estimated as:

$S\_complex \ \square\ (Kf) \times [Zn^2\square] \times [OH\square]\square$

Given the negligible free $Zn^2\square$, the equilibrium shifts to maximize complex formation, leading to a solubility significantly higher than in

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