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Understanding the molar solubility of salts in various solutions is a fundamental aspect of chemistry, particularly in the field of solubility equilibria. In this article, we will explore the detailed process of calculating the molar solubility of strontium iodate, Sr(IO3)2, in a solution containing 0.207 M sodium iodate (NaIO3). Given the solubility product constant (Ksp) of Sr(IO3)2 as 3.3 × 10^-7, we will analyze how the common ion effect influences solubility and work through the step-by-step solution to determine the molar solubility in this specific environment.

Understanding the Basics: Solubility, Ksp, and Common Ion Effect

Before diving into calculations, it’s essential to grasp core concepts related to solubility and equilibrium chemistry.

What is Solubility and Molar Solubility?

- Solubility refers to the maximum amount of a substance that dissolves in a solvent at a given temperature to form a saturated solution.
- Molar solubility is expressed in molarity (mol/L) and indicates the number of moles of a compound that dissolve per liter of solution to reach saturation.

What is the Solubility Product Constant (Ksp)?

- The Ksp is an equilibrium constant specific to sparingly soluble salts.
- It represents the product of the concentrations of the ions each raised to the power of their stoichiometric coefficients in the dissolution equilibrium at saturation.

For Sr(IO3)2, the dissociation in water is:

$$ \text{Sr(IO}_3)_2 (s) \rightleftharpoons \text{Sr}^{2+} (aq) + 2 \text{IO}_3^- (aq) $$

And the expression for the Ksp is:

$$ K_{sp} = [\text{Sr}^{2+}][\text{IO}_3^-]^2 $$

Dissolution Equilibrium of $Sr(IO_3)_2$

Knowing the dissociation reaction helps in setting up the problem:

$$ \text{Sr(IO}_3)_2 (s) \leftrightarrow \text{Sr}^{2+} (aq) + 2 \text{IO}_3^- (aq) $$

Let’s define the molar solubility of $Sr(IO_3)_2$ as s mol/L. Then,

- The concentration of Sr^{2+} ions at equilibrium is s mol/L.
- The concentration of IO_3^- ions at equilibrium is 2s mol/L.

In pure water, the solubility would be:

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

But in the presence of a common ion, the solubility decreases due to the common ion effect. Since $NaIO_3$ already provides IO_3^- ions, the initial concentration of IO_3^- is no longer zero.

Impact of Common Ion Effect: $NaIO_3$ as a Source of IO_3^-

In this problem, the solution contains 0.207 M $NaIO_3$, which supplies IO_3^- ions directly, affecting the solubility of $Sr(IO_3)_2$.

Key points:

- $NaIO_3$ is a strong electrolyte, dissociating completely:

$$ \text{NaIO}_3 \rightarrow \text{Na}^+ + \text{IO}_3^- $$

- The initial concentration of IO_3^- in solution is 0.207 M.

- The presence of this common ion shifts the equilibrium, reducing the molar solubility of $Sr(IO_3)_2$ compared to its solubility in pure water.

Setting Up the Calculation for Molar Solubility in 0.207 M NaIO3

Let's define:

- s = molar solubility of $Sr(IO_3)_2$ in mol/L (additional IO_3^- ions from $Sr(IO_3)_2$ dissolution).
- Initial concentration of IO_3^- from $NaIO_3$: 0.207 M.

At equilibrium:

- $[\text{Sr}^{2+}] = s$
- $[\text{IO}_3^-] = 0.207 + 2s$

The Ksp expression becomes:

$$ K_{sp} = [\text{Sr}^{2+}] \times [\text{IO}_3^-]^2 = s \times (0.207 + 2s)^2 $$

Given:

$$ K_{sp} = 3.3 \times 10^{-7} $$

Our goal:

Solve for s in:

$$ 3.3 \times 10^{-7} = s \times (0.207 + 2s)^2 $$

Mathematical Solution of the Equation

This equation involves a quadratic in s when expanded, but since the value of s is expected to be much smaller than 0.207 M, we can make an approximation to simplify the calculation.

Step 1: Justify the approximation

- Because Ksp is small, s will be small compared to 0.207.
- Therefore, $2s$ is negligible compared to 0.207, so:

$$ 0.207 + 2s \approx 0.207 $$

Step 2: Simplify the equation

Using the approximation:

$$ 3.3 \times 10^{-7} \approx s \times (0.207)^2 $$

Calculate $(0.207)^2$:

$$ (0.207)^2 = 0.042849 $$

Now solve for s:

$$ s \approx \frac{3.3 \times 10^{-7}}{0.042849} $$

$$ s \approx 7.7 \times 10^{-6} \text{ mol/L} $$

Step 3: Confirm the validity of the approximation

Calculate $2s$:

$$ 2s = 2 \times 7.7 \times 10^{-6} = 1.54 \times 10^{-5} $$

Compare to 0.207:

$$ \frac{1.54 \times 10^{-5}}{0.207} \approx 7.4 \times 10^{-5} $$

which is less than 1%, confirming the approximation is reasonable.

Final Molar Solubility of Sr(IO3)2 in 0.207 M NaIO3

Based on the calculations:

$$ \boxed{
\text{Molar solubility } s \approx 7.7 \times 10^{-6} \text{ mol/L}
} $$

This value indicates that in a solution containing 0.207 M NaIO3, the molar solubility of Sr(IO3)2 is approximately 7.7 μM.

Conclusions and Implications

- The presence of a common ion (IO3^-) from NaIO3 significantly reduces the solubility of Sr(IO3)2 due to the common ion effect.
- The calculated molar solubility (~7.7 μM) is markedly lower than what would be expected in pure water, emphasizing how common ions influence solubility equilibria.

- Understanding such effects is crucial in fields like analytical chemistry, environmental science, and materials synthesis, where controlling solubility is necessary.

Additional Considerations and Applications

- Effect of Temperature: The Ksp and solubility are temperature-dependent; the calculations here assume a constant temperature.
- Applications in Industry: Managing solubility in processes like crystal growth, wastewater treatment, and pharmaceutical formulation.
- Further Studies: For more precise results, especially when s is not negligible compared to initial ion concentrations, solving the quadratic equation without approximation is necessary.

Summary of Key Steps

1. Write the dissociation equation of $Sr(IO_3)_2$.
2. Express concentrations in terms of molar solubility s.
3. Incorporate initial common ion concentration from $NaIO_3$.
4. Set up the equilibrium expression with the known Ksp.
5. Use approximation for small s relative to initial ion concentration.
6. Calculate s to determine molar solubility.

By following this systematic approach, chemists can accurately determine the molar solubility of salts in complex solutions, facilitating a deeper understanding of solubility equilibria and their practical implications.

Frequently Asked Questions

What is the molar solubility of $Sr(IO_3)_2$ in 0.207 M $NaIO_3$ given its Ksp of 3.3×10^{-7}?

The molar solubility of $Sr(IO_3)_2$ in 0.207 M $NaIO_3$ can be calculated considering the common ion effect. Using the solubility product expression and the initial concentration of IO_3^- from $NaIO_3$, the molar solubility is approximately 7.3×10^{-6} M.

How does the presence of 0.207 M $NaIO_3$ affect the solubility of $Sr(IO_3)_2$?

The $NaIO_3$ provides IO_3^- ions, which suppress the dissolution of $Sr(IO_3)_2$ due to the common ion effect, thereby decreasing its molar solubility compared to pure water.

What is the expression for the solubility product (Ksp) of Sr(IO3)2 in this context?

Ksp = $[Sr^{2+}][IO_3^-]^2$. When $NaIO_3$ is present, $[IO_3^-]$ includes both the initial concentration from $NaIO_3$ and the IO_3^- released from $Sr(IO_3)_2$.

How can the molar solubility be estimated using the Ksp and the initial IO3⁻ concentration?

Assuming the molar solubility is 's', and considering the common ion effect, the equilibrium IO_3^- concentration becomes approximately 0.207 + 2s. Since s is small compared to 0.207, the approximation simplifies the calculation.

What is the significance of the Ksp value 3.3×10^{-7} in this calculation?

The Ksp value indicates the solubility equilibrium constant for $Sr(IO_3)_2$, which helps determine how much of the compound dissolves in a given solution under equilibrium conditions.

What assumptions are made when calculating molar solubility in this scenario?

The main assumptions include that the common ion effect significantly reduces solubility, and that the initial IO_3^- concentration from $NaIO_3$ remains essentially unchanged during dissolution, allowing for the approximation that $[IO_3^-] \approx 0.207$ M.

Why is it important to consider the common ion effect in solubility calculations?

The common ion effect suppresses the solubility of a salt because the presence of an ion already in solution shifts the equilibrium, reducing the amount of salt that can dissolve, which is critical for accurate solubility determination.

Additional Resources

Determine The Molar Solubility Of $Sr(IO_3)_2$ In 0.207 M $NaIO_3$

Understanding the solubility of a salt like strontium iodate ($Sr(IO_3)_2$) in a solution that already contains iodate ions (IO_3^-) is a fundamental problem in solubility equilibria and complex ion chemistry. This problem combines concepts of solubility product constants (Ksp), common ion effect, and equilibrium calculations. In this detailed review, we will delve into the principles, calculations, and implications of determining the molar solubility of $Sr(IO_3)_2$ in a 0.207 M $NaIO_3$ solution.

Introduction to the Problem

The goal is to find the molar solubility of $Sr(IO_3)_2$ in a solution that contains 0.207 M $NaIO_3$. The presence of $NaIO_3$ introduces IO_3^- ions into the solution, which influence the solubility of $Sr(IO_3)_2$ through the common ion effect. The key data points include:

- The solubility product constant (Ksp) of $Sr(IO_3)_2$: 3.3×10^{-7}
- The initial concentration of IO_3^- ions from $NaIO_3$: 0.207 M

The main challenge is to account for the increase in IO_3^- concentration due to both the dissolution of $Sr(IO_3)_2$ and the pre-existing IO_3^- ions from $NaIO_3$.

Understanding the Chemistry of $Sr(IO_3)_2$

1. Dissolution Equation

When $Sr(IO_3)_2$ dissolves in water, it dissociates as:

\[\text{Sr(IO}_3)_2 (s) \leftrightarrow \text{Sr}^{2+} (aq) + 2 \text{IO}_3^- (aq) \]

- The molar solubility of $Sr(IO_3)_2$, denoted as s, represents the molar concentration of Sr^{2+} ions in saturated solution.
- For every mole of $Sr(IO_3)_2$ that dissolves, 1 mole of Sr^{2+} and 2 moles of IO_3^- are produced.

2. Effect of the Common Ion (IO_3^-)

The presence of IO_3^- from $NaIO_3$ affects the solubility through Le Châtelier's principle, which states that the addition of a common ion suppresses the dissociation of the salt, reducing its solubility.

Formulating the Equilibrium Expression

The solubility product expression for $Sr(IO_3)_2$ is:

\[K_{sp} = [\text{Sr}^{2+}] \times [\text{IO}_3^-]^2 \]

In the presence of the common ion (IO_3^-) from $NaIO_3$, the total IO_3^- concentration is:

$$[\text{IO}_3^-]_{total} = 2s + [\text{IO}_3^-]_{initial}$$

where:

- $2s$ is the IO_3^- produced by the dissolution of $Sr(IO_3)_2$,
- $[\text{IO}_3^-]_{initial} = 0.207\, \text{M}$ from $NaIO_3$.

Since the initial IO_3^- concentration is significant, it will influence the solubility calculation.

Calculating the Molar Solubility

1. Defining Variables

Let:

- s = molar solubility of $Sr(IO_3)_2$ (moles per liter),
- Total IO_3^- in solution: $0.207 + 2s$.

The equilibrium concentrations are:

- $[\text{Sr}^{2+}] = s$,
- $[\text{IO}_3^-] = 0.207 + 2s$.

2. Substituting into the Ksp Expression

$$
K_{sp} = s \times (0.207 + 2s)^2
$$

Given that $K_{sp} = 3.3 \times 10^{-7}$, the equation becomes:

$$
3.3 \times 10^{-7} = s \times (0.207 + 2s)^2
$$

This is a quadratic equation in terms of s, which can be solved to find the molar solubility.

Approximations and Solution Strategy

Because the initial IO_3^- concentration from $NaIO_3$ (0.207 M) is quite high compared to the expected molar solubility s, an initial approximation can be made:

- Since s is likely much less than 0.207, the term $2s$ can initially be neglected in comparison to 0.207.

This simplifies the equation to:

$$
3.3 \times 10^{-7} \approx s \times (0.207)^2
$$

Calculating:

$$
s \approx \frac{3.3 \times 10^{-7}}{(0.207)^2}
$$

$$
s \approx \frac{3.3 \times 10^{-7}}{0.0428} \approx 7.71 \times 10^{-6} \, \text{M}
$$

This initial estimate suggests a molar solubility of approximately 7.7×10^{-6} M.

Refining the Calculation

Since the approximation assumes $2s \ll 0.207$, and our initial estimate confirms that s is indeed very small relative to 0.207, the approximation holds. However, to be more precise, we can perform an iterative correction:

- Use $s \approx 7.7 \times 10^{-6}$ M to compute $2s$:

$$
2s = 1.54 \times 10^{-5} \, \text{M}
$$

- Update IO_3^- concentration:

$$
[IO_3^-] = 0.207 + 1.54 \times 10^{-5} \approx 0.2070154\, \text{M}
$$

- Recalculate:

$$
s = \frac{3.3 \times 10^{-7}}{(0.2070154)^2} \approx \frac{3.3 \times 10^{-7}}{0.0428} \approx 7.71 \times 10^{-6} \, \text{M}
$$

The value converges, indicating that the initial approximation is sufficiently accurate.

Implications of the Calculation

The molar solubility of $Sr(IO_3)_2$ in 0.207 M $NaIO_3$ is approximately 7.7×10^{-6} M. This low value reflects the significant suppression of solubility due to the common ion effect—adding IO_3^- ions from $NaIO_3$ greatly reduces the salt's solubility compared to pure water.

Key Insights:

- Common ion effect: The presence of IO_3^- ions from $NaIO_3$ dramatically decreases the solubility of $Sr(IO_3)_2$.
- Quantitative impact: The molar solubility is reduced from what it would be in pure water (which can be estimated separately by setting initial $IO_3^- = 0$).
- Practical significance: This understanding helps in controlling the precipitation of salts in chemical processes, especially in analytical chemistry and industrial applications.

Additional Considerations and Variations

While the calculation above assumes ideal behavior and ignores activity coefficients, real-world systems may require corrections, especially at higher ionic strengths.

Factors influencing solubility:

- Ionic strength: Increased ionic strength from $NaIO_3$ can affect activity coefficients, altering the effective concentrations.
- Temperature dependence: Ksp values are temperature-dependent; the above calculation assumes a standard temperature (usually 25°C).
- Complex formation: Although not indicated here, potential complex ions with Sr^{2+} or IO_3^- could further affect solubility.

Extensions:

- Solubility in pure water: Calculating the molar solubility without the common ion effect provides a baseline for comparison.
- Precipitation conditions: Understanding the molarity at which $Sr(IO_3)_2$ begins to

precipitate in mixed solutions.

Summary

- The molar solubility of $Sr(IO_3)_2$ in 0.207 M $NaIO_3$ is approximately 7.7×10^{-6} M.
- The principal calculation involved understanding the dissolution equilibrium, considering the common ion effect, and solving the resulting quadratic equation.
- The presence of IO_3^- ions from $NaIO_3$ substantially suppresses the solubility of $Sr(IO_3)_2$, illustrating the importance of common ion effects in solubility equilibria.
- This calculation serves

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