For The Following Equilibrium, If The Concentration Of Barium Ion Is X, What Will Be The Molar Solubility

For The Following Equilibrium, If The Concentration Of Barium Ion Is X, What Will Be The Molar Solubility?

Understanding solubility equilibria is fundamental in chemistry, especially when dealing with insoluble salts and their dissolution behaviors. When a salt dissolves in water, it establishes an equilibrium between its solid form and its ions in solution. Determining the molar solubility—the number of moles of a salt that dissolve in a liter of solution to produce a saturated solution—is crucial for applications in analytical chemistry, environmental science, pharmaceuticals, and materials science.

This article aims to explore the concept of molar solubility in the context of a specific equilibrium involving barium ions, providing detailed explanations, step-by-step calculations, and insightful tips to understand how the concentration of one ion influences the overall solubility of the salt.

Understanding the Basic Concepts

Before delving into the specific problem, it's essential to review some foundational ideas:

Solubility Product Constant (Ksp)

- The solubility product constant, denoted as Ksp, is an equilibrium constant that describes the solubility of a sparingly soluble salt.

- For a generic salt AB_2 (where A and B are ions), dissolving as:

$$
AB_2 (s) \leftrightarrow A^{2+} (aq) + 2 B^- (aq)
$$

the Ksp is expressed as:

$$
K_{sp} = [A^{2+}][B^-]^2
$$

- The lower the Ksp, the less soluble the salt.

Molar Solubility

- Molar solubility refers to the number of moles of the salt that dissolve per liter of solution to reach saturation.
- It is directly related to the concentrations of ions in solution at equilibrium.

Le Châtelier's Principle and Common Ion Effect

- The presence of ions already in solution (common ions) can suppress the dissolution of the salt, affecting its molar solubility.
- If an ion common to the salt's dissociation is added or already present, the equilibrium shifts to the left, reducing solubility.

Examining the Specific Equilibrium Involving Barium Ions

Suppose we are given the equilibrium:

\[
BaSO_4 (s) \leftrightarrow Ba^{2+} (aq) + SO_4^{2-} (aq)
\]

with a known Ksp value.

The problem states:

"For the following equilibrium, if the concentration of barium ion ($[Ba^{2+}]$) is X, what will be the molar solubility of $BaSO_4$?"

This scenario involves understanding how the existing concentration of Ba^{2+} ions affects the solubility of $BaSO_4$.

Standard Dissolution of Barium Sulfate

- The dissolution can be represented as:

\[
\text{BaSO}_4 (s) \leftrightarrow \text{Ba}^{2+} (aq) + \text{SO}_4^{2-} (aq)
\]

- The solubility (S) is the molar amount of $BaSO_4$ that dissolves per liter, resulting in:

\[

$$
[Ba^{2+}] = S
$$

$$
[SO_4^{2-}] = S
$$

- Therefore, the solubility product expression becomes:

$$
K_{sp} = S^2
$$

- Solving for S gives:

$$
S = \sqrt{K_{sp}}
$$

However, if there's already a known concentration of Ba^{2+} (say, from an external source or previous dissolution), the equilibrium is disturbed, and the molar solubility will differ.

Impact of Existing Barium Ion Concentration (X) on Molar Solubility

To analyze the problem, consider:

- The initial concentration of Ba^{2+} is X.
- When $BaSO_4$ dissolves, it adds additional Ba^{2+} ions, but the total concentration will be affected by the common ion effect.

- The new equilibrium concentration of Ba^{2+} becomes:

$$
[Ba^{2+}]_{total} = X + s
$$

where:

- s is the molar solubility of $BaSO_4$ in this scenario (the amount that dissolves to restore equilibrium).

- The sulfate ion concentration will be:

$$
[SO_4^{2-}] = s
$$

Given that the initial Ba^{2+} concentration is X, and the dissolution of $BaSO_4$ adds s more, the overall ion concentrations at equilibrium are:

$$
[Ba^{2+}] = X + s
$$

$$
[SO_4^{2-}] = s
$$

The solubility product expression becomes:

$$
K_{sp} = (X + s) \times s
$$

\]

Deriving the Expression for Molar Solubility

To find s, the molar solubility, rearranged as:

\[

K_{sp} = s (X + s)

\]

which simplifies to a quadratic equation:

\[

s^2 + X s - K_{sp} = 0

\]

This quadratic equation can be solved as:

\[

s = \frac{-X \pm \sqrt{X^2 + 4 K_{sp}}}{2}

\]

Since molar solubility (s) cannot be negative, we take the positive root:

\[

s = \frac{-X + \sqrt{X^2 + 4 K_{sp}}}{2}

\]

This formula provides the molar solubility of $BaSO_4$ in the presence of an initial Ba^{2+} concentration of X.

Step-by-Step Calculation Example

Suppose:

- The Ksp of $BaSO_4$ is 1.1×10^{-10}
- The initial concentration of Ba^{2+} (X) is 1.0×10^{-4} mol/L

Calculating molar solubility (s):

$$
s = \frac{-X + \sqrt{X^2 + 4 K_{sp}}}{2}
$$

Plugging in the values:

$$
s = \frac{-1.0 \times 10^{-4} + \sqrt{(1.0 \times 10^{-4})^2 + 4 \times 1.1 \times 10^{-10}}}{2}
$$

Calculating inside the square root:

$$
(1.0 \times 10^{-4})^2 = 1.0 \times 10^{-8}
$$

\[

$$
4 \times 1.1 \times 10^{-10} = 4.4 \times 10^{-10}
$$

$$
\sqrt{1.0 \times 10^{-8} + 4.4 \times 10^{-10}} = \sqrt{1.044 \times 10^{-8}} \approx 1.021 \times 10^{-4}
$$

Now, compute s:

$$
s = \frac{-1.0 \times 10^{-4} + 1.021 \times 10^{-4}}{2} = \frac{2.1 \times 10^{-6}}{2} = 1.05 \times 10^{-6} \text{ mol/L}
$$

Result: The molar solubility of $BaSO_4$ in the presence of an initial Ba^{2+} concentration of 1.0×10^{-4} mol/L is approximately 1.05×10^{-6} mol/L.

Implications and Practical Applications

Understanding how the initial concentration of ions affects solubility has several practical implications:

- Precipitation Processes: Controlling ion concentrations can facilitate or inhibit the formation of precipitates in industrial processes.
- Water Treatment: Removal of specific ions is influenced by their solubility and existing ion concentrations, critical in softening and purification.
- Pharmaceuticals: Drug solubility can be affected by the presence of common ions, impacting bioavailability.
- Environmental Chemistry: The mobility and bioavailability of metal ions like Ba^{2+} depend on their solubility under various conditions.

Summary and Key Takeaways

- The molar solubility of a salt in the presence of an existing ion concentration can be derived from the solubility product constant and the initial ion concentration.
- The quadratic formula provides a straightforward method to calculate molar solubility under these conditions.
- When the initial concentration of the common ion (X) is high, the solubility decreases due to the common ion effect.
- The approach outlined here applies broadly to various sparingly soluble salts and their equilibria.

Final Notes

- Always verify the units and initial conditions before performing calculations.
- Remember that the quadratic solution must be physically meaningful (positive molar solubility).
- Adjustments may be necessary for complex equilibria involving multiple

Frequently Asked Questions

What is the general approach to determine the molar solubility of a salt involving barium ion when its concentration is given?

The approach involves writing the solubility equilibrium expression for the salt, expressing the concentrations of ions in terms of the molar solubility (X), and then solving for X using the given concentration of the barium ion.

How does the concentration of barium ion (X) influence the calculation

of molar solubility in equilibrium?

The concentration of barium ion (X) is used directly in the solubility product expression; knowing X allows us to relate it to the molar solubility by setting up an equilibrium expression and solving for the total molar solubility.

What is the typical equilibrium expression for a barium salt, such as $BaSO_4$, in terms of molar solubility?

For $BaSO_4$, the equilibrium expression is Ksp = $[Ba^{2+}][SO_4^{2-}]$, where $[Ba^{2+}]$ is given as X, and $[SO_4^{2-}]$ can be expressed in terms of the molar solubility, allowing calculation of the molar solubility.

If the concentration of Ba^{2+} is X, and the solubility product constant (Ksp) is known, how can we find the molar solubility?

We can set up the equilibrium expression Ksp = $[Ba^{2+}][SO_4^{2-}]$, substitute $[Ba^{2+}]$ = X, and solve for $[SO_4^{2-}]$, which relates directly to the molar solubility.

What role does the common ion effect play in calculating molar solubility when the Ba^{2+} concentration is already X?

The common ion effect can decrease the molar solubility of the salt because the presence of additional Ba^{2+} ions shifts the equilibrium, reducing solubility; this must be considered when calculating molar solubility with known X.

Can you provide a step-by-step method to determine the molar solubility given the concentration of Ba^{2+} as X?

Yes. First, write the solubility equilibrium expression for the salt. Then, express the concentrations of all ions in terms of X and molar solubility. Use the known X value and Ksp to set up an equation and solve for the molar solubility, considering any additional ion concentrations affecting the equilibrium.

Additional Resources

Understanding the Relationship Between Barium Ion Concentration and Molar Solubility in Equilibrium

When delving into the intricacies of solubility equilibria, especially involving sparingly soluble salts like barium sulfate ($BaSO_4$), it's essential to understand how the concentration of ions in solution relates to the compound's molar solubility. The question—"For the following equilibrium, if the concentration of Barium ion is X, what will be the molar solubility?"—sits at the core of analytical chemistry and equilibrium principles. This discussion aims to explore this concept thoroughly, dissecting the underlying chemistry, mathematical relationships, and practical applications.

Fundamentals of Solubility Equilibrium

Defining Solubility and Molar Solubility

- Solubility refers to the maximum amount of a substance that can dissolve in a solvent at a given temperature to form a saturated solution.
- Molar solubility is a specific expression, denoting the number of moles of a solute that dissolve per liter of solution to reach equilibrium saturation.

Key Point: Molar solubility is directly related to the concentrations of ions in a saturated solution, governed by the solubility product constant (Ksp).

The Concept of Solubility Product Constant (Ksp)

- For a generic salt AB_{s}, dissociating as:

$$
AB_{s} \leftrightarrow A^{+} + B^{-}
$$

the solubility product expression is:

$$
K_{sp} = [A^{+}][B^{-}]
$$

- For sparingly soluble salts like barium sulfate, which dissociate as:

$$
BaSO_4 (s) \leftrightarrow Ba^{2+} (aq) + SO_4^{2-} (aq)
$$

the Ksp expression is:

$$
K_{sp} = [Ba^{2+}][SO_4^{2-}]
$$

- Note: The solid phase activity is considered constant and thus omitted from the expression.

Implication: Knowing either the molar solubility or the concentration of one ion allows calculation of the other, provided the Ksp is known.

Analyzing Barium Sulfate Equilibrium

Equilibrium Expression for $BaSO_4$

- The dissociation of barium sulfate is represented as:

$$
BaSO_4 (s) \leftrightarrow Ba^{2+} (aq) + SO_4^{2-} (aq)
$$

- The solubility product constant (Ksp) is:

$$
K_{sp} = [Ba^{2+}][SO_4^{2-}]
$$

Assumption: When a saturated solution is formed, the molar solubility of $BaSO_4$, denoted as s, is the molar concentration of both Ba^{2+} and SO_4^{2-} ions, assuming no other sources or sinks.

Thus:

$$
[Ba^{2+}] = s
$$

$$
[SO_4^{2-}] = s
$$

and

$$
K_{sp} = s^2
$$

This straightforward relationship holds when the solution contains only the ions from $BaSO_4$.

Impact of Common Ions and Additional Sources

Presence of a Common Ion: Effect on Solubility

- When an external source of Ba^{2+} or SO_4^{2-} is introduced, it shifts the equilibrium.
- Le Châtelier's principle states that an increase in one ion's concentration decreases the solubility of $BaSO_4$ due to suppression of further dissolution.

Example:
- Adding $BaCl_2$ increases Ba^{2+} concentration.
- This increases the product $[Ba^{2+}][SO_4^{2-}]$ beyond Ksp, prompting precipitation of $BaSO_4$ and reducing the molar solubility.

Hence, the molar solubility in such cases depends on the initial ion concentrations in the solution.

Mathematical Model with Additional Ion Sources

Suppose:
- The concentration of Ba^{2+} ions in solution is X (given).
- The initial concentration of SO_4^{2-} ions is Y (which may be zero or non-zero).

The equilibrium expression becomes:

$$K_{sp} = [Ba^{2+}][SO_4^{2-}]$$

\[

$$
K_{sp} = X \times [SO_4^{2-}]
$$

$$
\Rightarrow [SO_4^{2-}] = \frac{K_{sp}}{X}
$$

If the only source of SO_4^{2-} is $BaSO_4$ dissolution:

$$
[SO_4^{2-}] = s
$$

$$
[Ba^{2+}] = X
$$

and the solubility product relationship is:

$$
K_{sp} = X \times s
$$

which rearranges to:

$$
s = \frac{K_{sp}}{X}
$$

This formula allows calculating molar solubility given the concentration of Ba^{2+} ions.

Calculating Molar Solubility When $[Ba^{2+}] = X$

Step-by-Step Calculation

1. Identify the knowns:

- The concentration of Ba^{2+} ions in solution, X.
- The solubility product constant, Ksp, for $BaSO_4$.

2. Express the relationship:

$$
s = \frac{K_{sp}}{X}
$$

where:

- s = molar solubility of $BaSO_4$.
- X = given concentration of Ba^{2+} ions.

3. Insert known values:

- For example, K_{sp} of $BaSO_4$ $\approx$ 1.1×10^{-10} at 25°C.
- If $X = 1 \times 10^{-3} \text{ M}$,

$$
s = \frac{1.1 \times 10^{-10}}{1 \times 10^{-3}} = 1.1 \times 10^{-7} \text{ M}
$$

4. Interpretation:

- The molar solubility is inversely proportional to the concentration of Ba^{2+}.
- As X increases, s decreases, reflecting the suppressive effect of common ions.

Practical Applications and Considerations

Real-World Contexts of Molar Solubility Calculations

- Water Treatment: Understanding how adding ions affects the precipitation or dissolution of salts.
- Pharmaceuticals: Solubility of salts influences drug formulation.
- Environmental Chemistry: Predicting mineral behavior in natural waters with varying ion compositions.
- Analytical Chemistry: Determining unknown concentrations based on solubility equilibria.

Limitations and Assumptions in Calculations

- Assumption of ideal behavior: activity coefficients are often ignored or assumed to be 1.
- Temperature dependence: Ksp varies with temperature.
- Presence of complex ions: Can alter ion activities and thus impact solubility.
- Ionic strength effects: High ionic strength can influence activity coefficients.

Always verify the conditions under which the data are valid before applying these calculations.

Summary and Key Takeaways

- The molar solubility of a salt like $BaSO_4$ depends critically on the concentrations of its constituent ions in solution.
- When the concentration of Ba^{2+} ions is known (X), the molar solubility s can be calculated using:

$$
s = \frac{K_{sp}}{X}
$$

- The relationship highlights the inverse proportionality between the ion concentration and the solubility.
- The presence of common ions (either Ba^{2+} or SO_4^{2-}) significantly reduces molar solubility due to

Le Châtelier’s principle.

- Understanding these principles allows chemists to manipulate conditions to control solubility for various scientific and industrial applications.

Final Thoughts

The question of "If the concentration of Barium ion is X, what will be the molar solubility?" underscores a fundamental aspect of equilibrium chemistry. It demonstrates the delicate balance between ion concentrations and the solubility of sparingly soluble salts. This knowledge is not only academically interesting but also practically vital across multiple scientific disciplines. Mastery of these relationships enables precise control and prediction of mineral behavior in complex systems, ultimately advancing technology and environmental management.

Note: For specific numerical problems, always ensure to use the most accurate and relevant Ksp values, consider activity coefficients in high ionic strength solutions, and account for temperature effects to refine your calculations.

For The Following Equilibrium If The Concentration Of Barium Ion Is X What Will Be The Molar Solubility

For The Following Equilibrium If The Concentration Of Barium Ion Is X What Will Be The Molar Solubility

Related Articles

- For Each Of The Values Below, Assume That The Represent An Error Correction Code Using An Encoding Scheme
- Find The Coordinates Of The Midpoint M Of ST???????. Then Find The Distance Between

[Points S And T. Round]
- [An Led Emits Green Light. Increasing The Size Of The Band Gap Could Change The Color Of The Emitted Light]

[Back to Home]