The Molar Solubility Of CuI Is 2.26 10 -6 M In Pure Water. Calculate The K Sp For CuI.5.11 10-12 M4.62

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Understanding the solubility of ionic compounds like copper(I) iodide (CuI) is fundamental in chemistry, especially when analyzing equilibrium processes in aqueous solutions. In this article, we will explore how to determine the solubility product constant (Ksp) of CuI based on given molar solubility data. We will also delve into the concepts of solubility, equilibrium, and how to interpret experimental measurements to calculate Ksp accurately.

Overview of CuI and Its Solubility

Copper(I) iodide (CuI) is an insoluble salt that dissociates in water according to the following equilibrium:

$$\text{CuI (s)} \leftrightarrow \text{Cu}^+ (aq) + \text{I}^- (aq)$$

The extent to which CuI dissolves in water is quantified by its molar solubility, which is the molar concentration of CuI that dissolves to reach equilibrium. The solubility product constant (Ksp) expresses the equilibrium concentrations of the ions in solution.

Understanding Molar Solubility and Ksp

What Is Molar Solubility?

- Molar solubility refers to the number of moles of a solute that dissolve per liter of solution to form a saturated solution.
- For CuI, molar solubility (denoted as s) is the molarity of CuI in a saturated solution.

What Is Ksp?

- The solubility product constant, Ksp, is an equilibrium expression that relates the

concentrations of ions in a saturated solution.
- For CuI, the Ksp expression is:

$$ K_{sp} = [\text{Cu}^+][\text{I}^-] $$

- Since CuI dissociates into one Cu$^+$ and one I$^-$ ion, their concentrations are equal at equilibrium, both equal to the molar solubility $ s $.

Calculating Ksp from Molar Solubility

Given the molar solubility $ s = 2.26 \times 10^{-6} \, M $, the calculation of Ksp is straightforward:

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

$$ K_{sp} = (2.26 \times 10^{-6})^2 $$

$$ K_{sp} = 5.11 \times 10^{-12} $$

This matches the given value of approximately $ 5.11 \times 10^{-12} $, confirming the consistency of the data.

Interpreting the Data and Additional Values

The mention of "4.62" in the original prompt appears to be a fragment or an incomplete reference. It might relate to another measurement or a different calculation, but based on the context, the primary focus is on calculating Ksp from molar solubility.

Additional Calculations and Context

Comparing Experimental Values

- Experimental determination of Ksp provides insight into how soluble CuI is in pure water.
- The low value indicates that CuI is very insoluble, consistent with its role as a precipitate in analytical chemistry.

Estimating Solubility in Other Conditions

- Factors such as temperature, presence of complexing agents, or ionic strength can influence solubility.
- For example, in the presence of complexing agents, the effective solubility may increase, affecting the Ksp indirectly.

Practical Applications of Ksp and Molar Solubility

Understanding and calculating Ksp has various practical implications:

- Precipitation Reactions: Determining conditions to precipitate or dissolve salts.
- Environmental Chemistry: Predicting mineral solubility in natural waters.
- Pharmaceuticals: Controlling solubility for drug formulation.
- Analytical Chemistry: Quantitative analysis based on solubility equilibria.

Step-by-Step Summary for Calculating Ksp

1. Identify the molar solubility s: Given as $2.26 \times 10^{-6} \, M$.
2. Write the dissociation equation: $\text{CuI (s)} \leftrightarrow \text{Cu}^+ (aq) + \text{I}^- (aq)$.
3. Express ion concentrations at equilibrium: Both are equal to s.
4. Calculate Ksp: $K_{sp} = s^2$.
5. Compute the value: $(2.26 \times 10^{-6})^2 = 5.11 \times 10^{-12}$.

Conclusion

In summary, the molar solubility of CuI in pure water is $2.26 \times 10^{-6} \, M$. Using this value, the solubility product constant (Ksp) can be calculated as:

$$\boxed{K_{sp} = 5.11 \times 10^{-12}}$$

This extremely low Ksp value underscores the insolubility of CuI, aligning with its common use as a precipitate in various chemical processes. Accurate knowledge of solubility and Ksp is vital for predicting the behavior of ionic compounds in aqueous solutions, whether in laboratory settings, environmental studies, or industrial applications.

Additional Reading and Resources

- "Chemistry: The Central Science" by Brown, LeMay, Bursten, and Murphy
- Online calculators for solubility product computations
- Research articles on copper halide solubility and complex formation

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Keywords: CuI solubility, molar solubility, Ksp calculation, ionic equilibrium, insoluble salts, solubility product constant, aqueous chemistry

Frequently Asked Questions

What is the molar solubility of CuI in pure water?

The molar solubility of CuI in pure water is 2.26×10^{-6} M.

How is the solubility product constant (Ksp) related to molar solubility?

Ksp is calculated using the concentrations of ions in saturated solution; for CuI, Ksp = $[Cu^+][I^-]$, which depends on the molar solubility.

What is the formula to calculate Ksp for CuI given its molar solubility?

Ksp = (molar solubility)^2, because CuI dissociates into Cu^+ and I^- ions in a 1:1 ratio.

What is the value of Ksp for CuI based on its molar solubility?

Ksp = $(2.26 \times 10^{-6})^2 = 5.11 \times 10^{-12}$.

How does the molar solubility of CuI relate to its solubility product constant?

The molar solubility directly influences Ksp; higher solubility results in a larger Ksp, indicating greater solubility.

Why is the square of molar solubility used in calculating Ksp for CuI?

Because CuI dissociates into two ions in a 1:1 ratio, so Ksp = $[Cu^+][I^-]$ = (molar solubility)^2.

What are the units of Ksp for CuI?

The units of Ksp are typically expressed as M^2 (molar squared), since it involves the product of ion concentrations.

Could the given value 4.62 relate to the calculation of Ksp?

No, based on the data, the correct Ksp is 5.11×10^{-12}; 4.62 may be a distractor or unrelated value in this context.

What is the significance of knowing the Ksp of CuI?

Knowing the Ksp helps determine the solubility of CuI in water and predict whether it will precipitate under certain conditions.

Additional Resources

The solubility product constant (Ksp) serves as a fundamental parameter in understanding the solubility behavior of ionic compounds in aqueous solutions. When considering compounds like copper(I) iodide (CuI), which exhibit limited solubility, calculating Ksp provides insights into their dissolution equilibrium and stability in water. This article delves into the process of determining the Ksp of CuI from its molar solubility, elucidates the underlying chemical principles, and explores the significance of these calculations in broader chemical contexts.

Understanding the Concept of Molar Solubility and Ksp

What is Molar Solubility?

Molar solubility refers to the number of moles of a solute that can dissolve in one liter of solvent to form a saturated solution. For ionic compounds like CuI, molar solubility indicates how much of the compound dissolves before the solution becomes saturated, reaching an equilibrium where the rate of dissolution equals the rate of precipitation.

For CuI, the molar solubility is given as 2.26 × 10-6 M in pure water. This means that in one liter of water, approximately 2.26 micro-moles of CuI dissolve to produce a saturated solution.

What is Ksp?

The solubility product constant (Ksp) is an equilibrium constant that quantifies the extent to which a salt dissolves in water. It is specific to a particular temperature and provides a measure of the compound's solubility.

For a generic salt AB that dissociates as:
$AB(s) \rightleftharpoons A^+(aq) + B^-(aq)$
the Ksp expression is:
$K_{sp} = [A^+][B^-]$

In the case of CuI, which dissociates as:
$CuI(s) \rightleftharpoons Cu^+(aq) + I^-(aq)$

the Ksp expression is:
$K_{sp} = [Cu^+][I^-]$

Since the molar solubility (s) of CuI is the concentration of Cu^+ and I^- ions in solution at equilibrium, the solubility product can be directly calculated once s is known.

Calculating Ksp for CuI from Molar Solubility

Step 1: Recognize the Dissociation Equation

The dissociation of CuI in water is straightforward:
$CuI(s) \rightleftharpoons Cu^+(aq) + I^-(aq)$

Because the solid dissolves into equal molar amounts of Cu^+ and I^-, their concentrations at equilibrium are both equal to the molar solubility, s.

Step 2: Express the Concentrations in Terms of Molar Solubility

Given:
$s = 2.26 \times 10^{-6}$ M

Thus:
$[Cu^+] = s = 2.26 \times 10^{-6}$ M
$[I^-] = s = 2.26 \times 10^{-6}$ M

Step 3: Write the Ksp Expression

Using the dissociation equation:
$K_{sp} = [Cu^+][I^-]$

Substituting the concentrations:

$Ksp = (2.26 \times 10-6) \times (2.26 \times 10-6) = (2.26 \times 10-6)^2$

Step 4: Perform the Calculation

Calculating:
$Ksp = (2.26)^2 \times (10-6)^2$
$= 5.1076 \times 10-12$

Rounding appropriately:
$Ksp \approx 5.11 \times 10-12$

Therefore, the solubility product constant for CuI at the given temperature is approximately $5.11 \times 10-12$.

Significance and Implications of the Ksp Value

Understanding the Low Solubility of CuI

The calculated Ksp value of approximately $5.11 \times 10-12$ indicates that CuI is very sparingly soluble in water. This low Ksp reflects a strong lattice energy relative to the hydration energy, making the compound resistant to dissolution. Such properties are crucial in applications where insolubility is desired, such as in semiconductors, or in contexts where controlled release or precipitation is necessary.

Comparison with Other Salts

When compared to more soluble salts, CuI's Ksp is orders of magnitude lower. For instance, common salts like sodium chloride (NaCl) have a Ksp around $1.8 \times 10-5$. The stark difference underscores the stability of the ionic lattice in CuI and its minimal tendency to dissolve.

Predicting Precipitation and Saturation

Knowing the Ksp allows chemists to predict whether a given mixture will result in precipitation. For example, if solutions containing $Cu+$ and $I-$ ions are mixed and their concentrations exceed the molar solubility corresponding to the Ksp, CuI will precipitate out of solution.

Additional Context: Computing the Ksp for Other Concentrations

The problem statement mentions a concentration of 4.62 (probably $4.62 \times 10-12$ M), but without full context, it's likely referring to an experimental value or a different scenario. For completeness, understanding how such values relate to the molar solubility and Ksp is important.

If a solution has ion concentrations exceeding the molar solubility, precipitation occurs. Conversely, if ion concentrations are below the solubility limit, the compound remains dissolved.

Effect of Temperature on Ksp and Solubility

It is essential to note that Ksp is temperature-dependent. As temperature increases, the solubility of most salts changes, affecting the Ksp. For CuI, precise Ksp calculations at different temperatures require experimental data, but the principles remain consistent.

Practical Applications and Broader Relevance

Industrial and Environmental Significance

Understanding the solubility of CuI is vital in industries such as electronics, where copper iodide finds use in semiconductors and photographic materials. Its insolubility ensures stability in these applications.

In environmental chemistry, the low solubility of copper compounds influences their mobility and bioavailability, impacting remediation strategies for copper contamination.

Analytical Techniques Leveraging Ksp

The concept of Ksp underpins many analytical techniques, including gravimetric analysis and titrations. By knowing the solubility limits, analysts can design experiments to quantify ionic species in complex mixtures.

Conclusion

The calculation of the Ksp for CuI, based on its molar solubility of 2.26 × 10-6 M, yields a value of approximately 5.11 × 10-12. This remarkably low Ksp underscores the compound's limited solubility in water and highlights the importance of solubility product constants in predicting and understanding the behavior of ionic compounds in aqueous environments. These insights not only deepen our fundamental comprehension of chemical equilibria but also inform practical applications across industries, environmental science, and analytical chemistry. Recognizing the delicate balance between lattice energy, hydration energy, and temperature effects remains essential in harnessing the properties of such compounds for technological and environmental advancements.

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