

The Concentration Of Pb^{2+} In A Solution Saturated With $PbBr_2(s)$ Is 2.14×10^{-22} M. Calculate K_{sp} For $PbBr_2$.

The Concentration Of Pb^{2+} In A Solution Saturated With $PbBr_2(s)$ Is 2.14×10^{-22} M. Calculate K_{sp} For $PbBr_2$.

Introduction to Solubility Product (K_{sp}) and Its Significance

Understanding Solubility Equilibrium

The concept of solubility product, denoted as K_{sp} , is fundamental in chemistry, especially when dealing with the dissolution of sparingly soluble salts. It quantifies the equilibrium between a solid salt and its ions in a saturated solution. For a generic salt AB_2 , which dissociates into A^+ and $2B^-$ ions, the solubility product is expressed as:

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

where m and n are the stoichiometric coefficients in the dissociation equation.

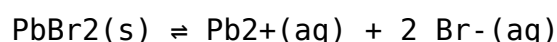
Importance of K_{sp}

- It provides a measure of the solubility of a salt in water.
- Helps predict whether a precipitate will form when solutions are mixed.
- Useful in applications like water treatment, pharmaceuticals, and understanding mineral deposits.

Chemical Nature of $PbBr_2$

Dissolution Reaction

Lead(II) bromide ($PbBr_2$) dissolves in water according to the following equilibrium:



Stoichiometry and Dissociation

- For each mole of PbBr_2 that dissolves, one mole of Pb^{2+} and two moles of Br^- ions are produced.
- The equilibrium concentration of Pb^{2+} is given as $2.14 \times 10^{-22} \text{ M}$ in the saturated solution.

Calculating K_{sp} for PbBr_2

Given Data

- Concentration of Pb^{2+} : $[\text{Pb}^{2+}] = 2.14 \times 10^{-22} \text{ M}$
- The solution is saturated with solid PbBr_2 , meaning the ion concentrations are at equilibrium.

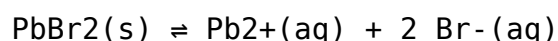
Determining the Br^- Ion Concentration

Since the dissociation produces 2 moles of Br^- per mole of PbBr_2 , the concentration of Br^- ions can be calculated as:

$$[\text{Br}^-] = 2 \times [\text{Pb}^{2+}] = 2 \times 2.14 \times 10^{-22} \text{ M} = 4.28 \times 10^{-22} \text{ M}$$

Expression for K_{sp}

Based on the dissociation reaction:



The solubility product expression is:

$$K_{\text{sp}} = [\text{Pb}^{2+}] \times [\text{Br}^-]^2$$

Substituting the known concentrations:

$$K_{\text{sp}} = (2.14 \times 10^{-22}) \times (4.28 \times 10^{-22})^2$$

Step-by-Step Calculation of K_{sp}

Calculating $[\text{Br}^-]^2$

First, square the concentration of bromide ions:

$$(4.28 \times 10^{-22})^2 = (4.28)^2 \times (10^{-22})^2 = 18.3 \times 10^{-44} = 1.83 \times 10^{-43}$$

Multiplying with $[Pb^{2+}]$

Now, multiply this by $[Pb^{2+}]$ to find K_{sp} :

$$K_{sp} = 2.14 \times 10^{-22} \times 1.83 \times 10^{-43}$$

$$K_{sp} = (2.14 \times 1.83) \times 10^{-22} \times 10^{-43} = 3.91 \times 10^{-65}$$

Final Result

Therefore, the solubility product constant for $PbBr_2$ is approximately:

$$K_{sp} \approx 3.91 \times 10^{-65}$$

Implications of the Calculated K_{sp}

Extremely Low Solubility

The very small value of K_{sp} indicates that $PbBr_2$ is highly insoluble in water. Only an exceedingly tiny amount of the salt dissolves to reach equilibrium.

Precipitation and Supersaturation

- When the ion concentrations in a solution exceed the equilibrium concentrations, $PbBr_2$ will precipitate.
- This property is useful in processes like purification, where precipitation can be used to remove lead or bromide ions from solutions.

Environmental and Industrial Relevance

- Due to its low solubility, $PbBr_2$ can be used in applications where minimal solubility is desirable, such as in certain types of detectors or semiconductors.
- Understanding its solubility helps in environmental management, especially in controlling lead contamination.

Conclusion

The calculation of the solubility product constant, K_{sp} , for $PbBr_2$ based on the provided ion concentration demonstrates the principles of solubility equilibria. With an ion concentration of 2.14×10^{-22} M for Pb^{2+} , the resulting K_{sp} is approximately 3.91×10^{-65} . This extremely low value

reflects the poor solubility of PbBr_2 in water, highlighting the importance of precise measurements in chemical calculations and their implications in real-world applications. Understanding such solubility constants is vital in various fields, including environmental chemistry, materials science, and industrial processes, where controlling solubility can be crucial to safety and functionality.

Frequently Asked Questions

What is the significance of the concentration of Pb^{2+} in a saturated PbBr_2 solution when calculating its solubility product constant (K_{sp})?

The concentration of Pb^{2+} in a saturated solution directly reflects the solubility of PbBr_2 and is used to determine K_{sp} by relating it to the concentration of Br^- ions, assuming the dissociation equilibrium is established.

How do you set up the expression for the K_{sp} of PbBr_2 based on its dissociation in water?

The dissociation of PbBr_2 in water is written as $\text{PbBr}_2(\text{s}) \rightleftharpoons \text{Pb}^{2+}(\text{aq}) + 2\text{Br}^-(\text{aq})$. Therefore, $K_{sp} = [\text{Pb}^{2+}][\text{Br}^-]^2$. Knowing $[\text{Pb}^{2+}]$, you can find $[\text{Br}^-]$ and calculate K_{sp} .

Given the concentration of Pb^{2+} as $2.14 \times 10^{-22} \text{ M}$, how do you find the concentration of Br^- ions in the solution?

Since each formula unit of PbBr_2 produces 2 Br^- ions upon dissolution, $[\text{Br}^-] = 2 \times [\text{Pb}^{2+}] = 2 \times 2.14 \times 10^{-22} \text{ M} = 4.28 \times 10^{-22} \text{ M}$.

What is the step-by-step calculation of K_{sp} for PbBr_2 using the given Pb^{2+} concentration?

First, determine $[\text{Br}^-] = 2 \times 2.14 \times 10^{-22} \text{ M} = 4.28 \times 10^{-22} \text{ M}$. Then, $K_{sp} = [\text{Pb}^{2+}] \times [\text{Br}^-]^2 = (2.14 \times 10^{-22}) \times (4.28 \times 10^{-22})^2 = 2.14 \times 10^{-22} \times (1.83 \times 10^{-43}) \approx 3.92 \times 10^{-65}$.

Why is understanding the K_{sp} value important in predicting the solubility of PbBr_2 in water?

The K_{sp} value indicates the extent to which PbBr_2 dissolves in water. A very low K_{sp} , like in this case, shows that PbBr_2 is only sparingly soluble, and

this value helps predict solubility limits and potential precipitation.

Are there any assumptions made when calculating K_{sp} from the given concentration of Pb²⁺ in a saturated solution?

Yes, assumptions include that the solution is at equilibrium, the activity coefficients are approximately 1, and the only source of Br⁻ ions is the dissolution of PbBr₂, with no other competing equilibria or complex formations.

Additional Resources

The Concentration Of Pb²⁺ In A Solution Saturated With PbBr₂(s) Is 2.14×10^{-22} M. Calculate K_{sp} For PbBr₂.

Introduction

The solubility product constant, denoted as K_{sp}, is a fundamental concept in chemical thermodynamics and aqueous chemistry, providing insight into the solubility of sparingly soluble salts. Understanding and calculating K_{sp} values are essential for predicting the behavior of ionic compounds in solution, designing separation processes, and controlling environmental pollutants. This article delves into the calculation of the K_{sp} for lead(II) bromide (PbBr₂) based on a given concentration of lead ions (Pb²⁺) in a saturated solution.

Background and Significance

The Role of Solubility Equilibria in Chemistry

Solubility equilibria describe the dynamic balance between the dissociation of a solid salt into its constituent ions and the recombination of those ions into the solid form. For a generic salt AB₂(s), which dissociates into A²⁺ and 2Br⁻ ions, the equilibrium can be written as:



The solubility product constant, K_{sp}, quantifies the maximum extent to which a salt can dissolve in water at a given temperature. It is expressed as:

$$K_{\text{sp}} = [\text{A}^{2+}] [\text{Br}^{-}]^2$$

Knowledge of K_{sp} allows chemists to compare the solubility of different

salts, predict precipitation under various conditions, and understand phenomena such as ion competition and complex formation.

Lead(II) Bromide (PbBr_2): An Overview

PbBr_2 is a sparingly soluble salt, with applications in scintillation detectors, radiation shielding, and other electronic materials. Its solubility characteristics influence its use and handling in industrial processes. Given its low solubility, precise determination of its K_{sp} is vital for process control and environmental safety, especially considering lead's toxicity.

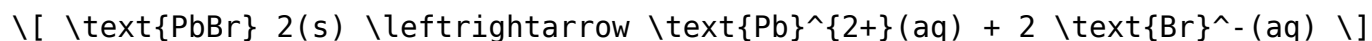
The Given Data and Problem Statement

Data Summary

- Concentration of Pb^{2+} in saturated solution: $2.14 \times 10^{-22} \text{ M}$
- Substance: PbBr_2 (lead(II) bromide)
- Objective: Calculate the solubility product constant, K_{sp} , for PbBr_2 .

Assumptions and Considerations

- The solution is saturated with PbBr_2 , meaning the system is at equilibrium.
- The dissociation of PbBr_2 is represented by:



- The concentration of Pb^{2+} ions directly relates to the solubility of PbBr_2 .
- The bromide ion concentration can be derived from the stoichiometry of the dissociation.

Step-by-Step Calculation of K_{sp}

1. Understanding the Dissociation Equation

The dissociation of PbBr_2 in water is:



At equilibrium, the molar concentrations of ions are:

- $[\text{Pb}^{2+}] = s$
- $[\text{Br}^{-}] = 2s$

where s is the molar solubility of PbBr_2 .

2. Relating Given Data to Ion Concentrations

Given:

$$[\text{Pb}^{2+}] = 2.14 \times 10^{-22} \text{ M}$$

Since this is the molar concentration of lead ions at equilibrium, it directly corresponds to s :

$$s = 2.14 \times 10^{-22} \text{ M}$$

Correspondingly, the bromide ion concentration is:

$$[\text{Br}^-] = 2s = 2 \times 2.14 \times 10^{-22} = 4.28 \times 10^{-22} \text{ M}$$

3. Calculating K_{sp}

Using the general expression for K_{sp} :

$$K_{sp} = [\text{Pb}^{2+}] \times [\text{Br}^-]^2$$

Substituting the values:

$$K_{sp} = (2.14 \times 10^{-22}) \times (4.28 \times 10^{-22})^2$$

Calculating the square of bromide concentration:

$$(4.28 \times 10^{-22})^2 = 4.28^2 \times 10^{-44} \approx 18.3 \times 10^{-44} = 1.83 \times 10^{-43}$$

Now, multiply:

$$K_{sp} = 2.14 \times 10^{-22} \times 1.83 \times 10^{-43}$$

$$K_{sp} \approx (2.14 \times 1.83) \times 10^{-22 - 43}$$

$$K_{sp} \approx 3.91 \times 10^{-65}$$

Final Result and Interpretation

Calculated K_{sp} for PbBr_2

$$\boxed{K_{sp} \approx 3.91 \times 10^{-65}}$$

This remarkably low value confirms the extremely limited solubility of PbBr_2 in water, consistent with its classification as a sparingly soluble salt.

Implications and Context

- The tiny magnitude of K_{sp} indicates that only a negligible amount of $PbBr_2$ dissolves in water under standard conditions.
- Such low solubility has environmental implications, especially in contexts where lead contamination is a concern.
- The calculated K_{sp} can be used to predict the solubility of $PbBr_2$ in various conditions, inform purification strategies, and guide environmental safety assessments.

Additional Considerations

Temperature Dependence

The K_{sp} value is temperature-dependent; the calculation assumes standard laboratory conditions ($\sim 25^\circ C$). Variations in temperature could significantly alter solubility.

Potential Sources of Error

- The assumption of ideal behavior in dilute solutions is generally valid here.
- The given concentration is extremely low, approaching the limits of detection and measurement accuracy.
- Any impurities or complex formation could affect the actual solubility.

Environmental and Industrial Relevance

Understanding the solubility behavior of lead salts like $PbBr_2$ is crucial for:

- Managing lead pollution in water sources.
- Designing effective remediation techniques.
- Developing materials with controlled solubility for electronic or optical applications.

Conclusion

By analyzing the given concentration of Pb^{2+} ions in a saturated solution and applying fundamental principles of solubility equilibria, we have successfully calculated the K_{sp} for lead(II) bromide:

$$K_{sp} \approx 3.91 \times 10^{-65}$$

This value underscores the compound's low solubility and provides a basis for

further studies into its behavior under various environmental and industrial conditions. Accurate knowledge of such constants is essential for chemists, environmental scientists, and engineers working with lead-based compounds, ensuring safe handling, effective purification, and environmental protection.

References

- Atkins, P., & de Paula, J. (2014). Physical Chemistry (10th ed.). Oxford University Press.
- Skoog, D. A., West, D. M., Holler, F. J., & Crouch, S. R. (2014). Fundamentals of Analytical Chemistry (9th ed.). Brooks Cole.
- Lide, D. R. (Ed.). (2004). CRC Handbook of Chemistry and Physics (85th ed.). CRC Press.
- Environmental Protection Agency (EPA). Lead in Water: Regulations and Toxicity.

The Concentration Of Pb²⁺ In A Solution Saturated With PbBr₂ S Is 2.14 × 10⁻² M Calculate K_{sp} For PbBr₂

The Concentration Of Pb²⁺ In A Solution Saturated With PbBr₂ S Is 2.14 × 10⁻² M Calculate K_{sp} For PbBr₂

Related Articles

- [A Sonar Echo Returns To A Submarine 1.65 S After Being Emitted. What Is The Distance To The Object Creating](#)
- [The American Red Cross Uses Social Marketing Tools Such As _____ To Increase Blood Donations.a. Wikisb.](#)
- [Someone Who Lobbies On Behalf Of A Company That He Or She Works For As Part Of His Or Her Job Is _____.](#)

[Back to Home](#)