If Silver Chloride (ksp=1.81010) Is Placed In A 0.35 M Solution Of Sodium Chloride, What Will Be The

If Silver Chloride (ksp=1.81010) Is Placed In A 0.35 M Solution Of Sodium Chloride, What Will Be The fate of its solubility and how do various factors influence its behavior? This question touches on fundamental concepts in solubility equilibria, common ion effect, and ionic interactions in aqueous solutions. Understanding these principles provides insight into the dynamics of silver chloride in chloride-rich environments and is vital in fields ranging from analytical chemistry to materials science.

Introduction to Silver Chloride and Its Solubility

What Is Silver Chloride?

Silver chloride (AgCl) is an insoluble salt known for its bright white appearance and photographic application. It belongs to a class of compounds called sparingly soluble salts, characterized by low solubility products (Ksp). The Ksp value provides a quantitative measure of the salt's solubility, indicating the extent to which it dissolves in water.

Understanding the Solubility Product (Ksp)

The solubility product constant, symbolized as Ksp, is defined for a salt's dissociation equilibrium. For AgCl, the dissolution can be represented as:

$AgCl(s) \rightleftharpoons Ag^{+}(aq) + Cl^{-}(aq)$

The corresponding Ksp expression:

$K_{sp} = [Ag^{+}][Cl^{-}]$

Given that Ksp for AgCl is 1.8×10^{-10}, this indicates that at equilibrium, the concentrations of Ag^{+} and Cl^{-} ions in solution are very low, reflecting its limited solubility.

Impact of Common Ions on Solubility

The Common Ion Effect Explained

The common ion effect describes the reduction in solubility of a salt when a solution already contains one of its constituent ions. When ions common to the salt's dissociation are present in solution at significant concentrations, they shift the equilibrium position, favoring the formation of the solid phase.

Silver Chloride in a Sodium Chloride Solution

Sodium chloride (NaCl) is a highly soluble salt that dissociates completely in water:

$NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$

In a 0.35 M NaCl solution, the chloride ion concentration is already high. When AgCl is introduced into this solution, the pre-existing Cl^{-} ions influence the equilibrium, typically decreasing the solubility of AgCl.

Quantitative Analysis: Calculating the Solubility of AgCl in 0.35 M NaCl

Setting Up the Equilibrium Expression

Let’s denote the molar solubility of AgCl in the NaCl solution as 's' mol/L. This is the concentration of Ag^+ ions that result from the dissolution of AgCl. Since the solution already contains 0.35 M Cl^- ions from NaCl, the total chloride ion concentration becomes:

$[Cl^-] = 0.35 + s$

The silver ion concentration is equal to 's' because each mole of AgCl dissolves to produce one mole of Ag^+:

$[Ag^+] = s$

The solubility equilibrium expression:

$K_{sp} = [Ag^+][Cl^-] = s\ (0.35 + s)$

Given that $K_{sp} = 1.8 \times 10^{-10}$, the equation becomes:

$1.8 \times 10^{-10} = s\ (0.35 + s)$

Solving for 's'

Since Ksp is very small and s is expected to be much less than 0.35 M (due to the common ion effect), the term s^2 can be neglected compared to 0.35 s. This simplifies the equation to:

$$1.8 \times 10^{-10} \approx 0.35\,s$$

Solving for 's':

$$s \approx (1.8 \times 10^{-10}) / 0.35 \approx 5.14 \times 10^{-10}\ M$$

This molar concentration 's' represents the solubility of AgCl in 0.35 M NaCl solution.

Conclusion from the Calculation

The solubility of AgCl in a 0.35 M NaCl solution is approximately 5.14×10^{-10} M, which is significantly lower than its solubility in pure water (approximately 1.3×10^{-5} M). This illustrates how the presence of common ions (Cl^- in this case) drastically decreases the solubility of sparingly soluble salts.

Factors Influencing Silver Chloride Solubility

Role of Ionic Strength

The ionic strength of the solution, which depends on the concentration of ions present, affects activity coefficients and thus solubility. Higher ionic strength typically decreases solubility due to increased electrostatic interactions.

Temperature Effects

As with most salts, temperature influences solubility. For AgCl, increasing temperature generally increases solubility slightly, although the effect is minimal compared to the impact of the common ion effect.

Complex Formation

In some cases, Ag^+ ions can form complexes with other ligands (like ammonia or cyanide), which can increase the solubility of AgCl. However, in chloride-rich solutions like NaCl, complex formation with chloride ions does not significantly affect solubility.

Practical Applications and Significance

Analytical Chemistry

Understanding the solubility behavior of AgCl in chloride-rich solutions is crucial in analytical procedures, especially in gravimetric and titrimetric analyses where silver chloride is used as a precipitate for determining chloride content.

Corrosion and Environmental Chemistry

The behavior of silver chloride can influence corrosion processes, especially in marine environments where chloride ions are abundant. Silver-based materials may undergo surface transformations affected by chloride ions, impacting their durability.

Photographic and Material Science Applications

Silver chloride's sensitivity to light makes it useful in photographic film. Knowledge of its solubility and stability in chloride solutions can influence manufacturing processes and storage conditions.

Summary and Key Takeaways

- Silver chloride has a low solubility in water, with a Ksp of approximately 1.8 × 10^(-10).
- In a 0.35 M sodium chloride solution, the solubility of AgCl decreases dramatically due to the common ion effect, dropping to about 5.14 × 10^(-10) M.
- Factors such as temperature, ionic strength, and complex formation can influence the solubility of AgCl.
- Understanding these principles is essential in analytical chemistry, environmental science, and material applications involving silver compounds.

Conclusion

The interaction between silver chloride and chloride-rich environments exemplifies fundamental concepts of solubility and equilibrium. When AgCl is placed in a 0.35 M NaCl solution, its solubility diminishes significantly, illustrating the powerful effect of common ions. This knowledge is crucial for chemists and material scientists working with silver salts, ensuring proper handling and application in various technological and scientific fields.

References:

- Atkins, P., & de Paula, J. (2014). Physical Chemistry (10th Edition). Oxford University Press.
- Chang, R. (2010). Chemistry. McGraw-Hill Education.
- Lide, D. R. (2004). CRC Handbook of Chemistry and Physics. CRC Press.
- Silver chloride solubility data from standard reference tables.

Note: Always remember that real-world conditions may slightly vary, and actual measurements are essential for precise applications.

Frequently Asked Questions

What happens when silver chloride is placed in a 0.35 M sodium chloride solution?

Silver chloride will undergo further dissolution or precipitation depending on the common ion effect, but given the solubility product (Ksp = 1.8×10^-10) and high chloride ion concentration, it is more likely to precipitate or remain in equilibrium with the solution.

How does the presence of 0.35 M NaCl affect the solubility of silver chloride?

The high chloride ion concentration from NaCl suppresses the solubility of silver chloride due to the common ion effect, causing less AgCl to dissolve in the solution.

Will silver chloride precipitate or dissolve in the 0.35 M NaCl solution?

Silver chloride will tend to precipitate or remain largely insoluble because the chloride ion concentration exceeds the solubility limits dictated by its Ksp, promoting precipitation.

What is the approximate concentration of dissolved silver ions in this solution?

The concentration of Ag+ ions will be very low, approaching the value dictated by the solubility equilibrium with the chloride ions present, likely close to 1.8×10^-10 M, due to the common ion effect.

How can we determine if silver chloride will precipitate in this solution?

Compare the ionic product [Ag+][Cl-] to the Ksp; if it exceeds 1.8×10^-10, precipitation occurs. Given the high Cl- concentration, silver chloride is expected to precipitate or be in equilibrium with a very low Ag+ concentration.

What is the significance of the common ion effect in this scenario?

The common ion effect reduces the solubility of silver chloride because the excess chloride ions from NaCl shift the equilibrium toward the formation of solid AgCl, decreasing dissolved silver ion concentration.

Additional Resources

Understanding the Solubility of Silver Chloride in a Sodium Chloride Solution: A Comprehensive Analysis

When examining the interplay between silver chloride (AgCl) and sodium chloride (NaCl) solutions, it is crucial to understand the underlying principles of solubility equilibria, common ion effects, and the implications of solubility product constants (Ksp). This detailed review aims to dissect the question: If silver chloride (Ksp = 1.8×10^{-10}) is placed in a 0.35 M sodium chloride solution, what will be the solubility of AgCl under these conditions?

By exploring the chemistry involved, we will clarify how the presence of a common ion (Cl^-) affects the solubility of AgCl, interpret the relevant calculations, and discuss broader implications for solubility equilibria in aqueous solutions.

Fundamentals of Solubility and Solubility Product (Ksp)

What is Ksp?

- The solubility product constant, Ksp, quantifies the maximum product of ion concentrations in a saturated solution of an ionic compound at equilibrium.
- For AgCl, the dissociation can be written as:

$$
\mathrm{AgCl (s)} \rightleftharpoons \mathrm{Ag^+ (aq)} + \mathrm{Cl^- (aq)}
$$

- The expression for Ksp:

$$
K_{sp} = [\mathrm{Ag^+}] \times [\mathrm{Cl^-}]
$$

- For AgCl, the given Ksp is:

$$
K_{sp} = 1.8 \times 10^{-10}
$$

- This small Ksp indicates that AgCl is sparingly soluble in water.

Molar Solubility (s)

- Molar solubility (s) is the concentration of the dissolved solid in mol/L when the solution is saturated.
- In pure water, the molar solubility of AgCl is:

$$
s = [\mathrm{Ag^+}] = [\mathrm{Cl^-}]
$$

- Therefore, for pure water:

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

$$
s = \sqrt{K_{sp}} = \sqrt{1.8 \times 10^{-10}} \approx 1.34 \times 10^{-5} \ \text{M}
$$

- This is the molar solubility in pure water, but the presence of additional chloride ions will influence it.

Impact of Common Ions on Solubility

The Common Ion Effect

- When a solution already contains ions that are part of the equilibrium, the solubility of the ionic compound decreases.
- In this scenario, sodium chloride (NaCl) provides a high concentration of chloride ions:

$$
[\mathrm{Cl^-}]_{initial} = 0.35\, \text{M}
$$

- The common ion effect occurs because the chloride ion concentration is already high, shifting the equilibrium position to reduce the dissolution of AgCl.

Le Chatelier's Principle

- Adding more chloride ions pushes the equilibrium:

$$
\mathrm{AgCl (s)} \rightleftharpoons \mathrm{Ag^+ (aq)} + \mathrm{Cl^- (aq)}
$$

to the left, decreasing the solubility of AgCl.

Quantitative Analysis: Calculating the Solubility in 0.35 M NaCl

Setting Up the Equilibrium Expression

- Let's denote the molar solubility of AgCl in the presence of the chloride ion concentration $[\mathrm{Cl^-}]_{initial} = 0.35\, \text{M}$.

- The total chloride ion concentration at equilibrium will be:

$$
[\mathrm{Cl^-}]_{total} = 0.35\, \text{M} + s
$$

- Since s (the molar solubility of AgCl) is expected to be very small relative to 0.35 M, we can approximate:

$$
[\mathrm{Cl^-}]_{total} \approx 0.35\, \text{M}
$$

- The equilibrium concentrations are:

$$
[\mathrm{Ag^+}] = s
$$

$$
[\mathrm{Cl^-}] \approx 0.35\, \text{M}
$$

- Therefore, the Ksp expression becomes:

$$
K_{sp} = s \times [\mathrm{Cl^-}] \approx s \times 0.35
$$

- Solving for s:

$$
s = \frac{K_{sp}}{[\mathrm{Cl^-}]} = \frac{1.8 \times 10^{-10}}{0.35}
$$

$$
s \approx 5.14 \times 10^{-10}\, \text{M}
$$

Interpretation of Results

- The molar solubility of AgCl in a 0.35 M NaCl solution is approximately 5.14×10^{-10} M, which is significantly lower than the solubility in pure water (~1.34×10^{-5} M).

- This dramatic decrease exemplifies the profound impact that common ions have on solubility.

Broader Implications and Additional Considerations

Effect of Increasing Chloride Concentration

- As chloride ion concentration increases, the solubility of AgCl decreases further.
- For extremely high chloride concentrations, the solubility could become negligible, leading to almost complete precipitation.

Practical Applications

- This phenomenon is exploited in qualitative analysis: to precipitate silver chloride selectively in the presence of other halides.
- It’s also relevant in industrial processes where controlling the solubility of silver salts is necessary.

Limitations of the Approximation

- The approximation that $[\mathrm{Cl^-}]_{total} \approx 0.35\, \text{M}$ assumes $s \ll 0.35\, \text{M}$.
- Because the calculated s (~5×10^{-10} M) is many orders of magnitude smaller, the approximation is valid.
- For more precise calculations, iterative methods or more complex models could be employed, but they are unnecessary here.

Additional Factors Affecting Solubility

- Temperature: The solubility of AgCl varies with temperature, generally increasing with higher temperatures.
- Ionic Strength: The overall ionic strength of the solution impacts activity coefficients, slightly modifying the effective concentrations.
- Presence of Complexing Agents: Ligands such as ammonia can form complexes with silver, increasing solubility despite high chloride levels.

Summary and Final Remarks

- The key takeaway is that in a 0.35 M sodium chloride solution, the solubility of silver chloride is drastically reduced due to the common ion effect.
- Quantitatively, the molar solubility drops from approximately 1.34×10^{-5} M in pure water to about 5.14×10^{-10} M in the presence of 0.35 M chloride ions.
- This example underscores the importance of solution composition in predicting solubility behavior and designing processes involving ionic compounds.

In conclusion, placing AgCl in a 0.35 M NaCl solution results in an extremely low solubility of silver chloride, effectively preventing significant dissolution. This understanding is foundational in analytical chemistry, environmental science, and industrial applications where controlling ionic equilibria is essential.

Note: Always consider the specific conditions and potential complexities in real-world scenarios, including temperature variations and the presence of other ions or complexing agents, which can influence the solubility behavior beyond simplified calculations.

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