Determine The Molar Solubility Of CuCl In A Solution Containing 0.060 M KCl. K Sp (CuCl) = 1.0 10 -6.1.0

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Understanding the molar solubility of copper(I) chloride (CuCl) in various solutions is a fundamental aspect of solubility equilibria in chemistry. When CuCl is introduced into a solution already containing potassium chloride (KCl), the common ion effect significantly influences its solubility. In this detailed article, we will explore how to determine the molar solubility of CuCl in a 0.060 M KCl solution, considering the solubility product constant (Ksp) of CuCl is 1.0 x 10^-6.1.0. This comprehensive guide will cover the theoretical background, step-by-step calculations, and practical implications, optimized for SEO to assist students, educators, and professionals alike.

Understanding the Basics: Solubility Equilibrium and Ksp

What Is Solubility Equilibrium?

Solubility equilibrium occurs when a salt dissolves in water until the rate of dissolution equals the rate of precipitation. At this point, the solution is saturated, and the concentrations of ions in solution remain constant.

For CuCl, the dissolution can be represented as:

$$ \mathrm{CuCl(s)} \leftrightarrow \mathrm{Cu^+ (aq)} + \mathrm{Cl^- (aq)} $$

What Is the Solubility Product Constant (Ksp)?

The Ksp is an equilibrium constant specific to the solubility of a sparingly soluble salt. It expresses the maximum product of the molar concentrations of the constituent ions in a saturated solution at equilibrium.

For CuCl:

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

Given:

$$ K_{sp} = 1.0 \times 10^{-6.1.0} $$

(Note: The value appears to be a formatting error; assuming the correct Ksp is 1.0 x 10^-6.1 for illustration purposes.)

Impact of the Common Ion Effect on Solubility

What Is the Common Ion Effect?

When a solution contains an ion that is also a product of a salt's dissociation, the solubility of that salt decreases. This phenomenon is known as the common ion effect.

In our case:
- The solution contains KCl, providing a common ion, Cl^-.
- The Cl^- concentration from KCl will influence the solubility of CuCl.

Why Does the Common Ion Effect Matter?

The presence of additional Cl^- ions shifts the equilibrium position, reducing the molar solubility of CuCl because the system favors the formation of the solid phase to reduce ion concentrations in solution.

Step-by-Step Calculation of Molar Solubility of CuCl

Step 1: Write the Dissociation Equation and Expression for Ksp

The dissociation:
$$\mathrm{CuCl(s)} \leftrightarrow \mathrm{Cu^+ (aq)} + \mathrm{Cl^- (aq)}$$

Since the solution already contains Cl^- from KCl, the total Cl^- concentration will be:
$$[\mathrm{Cl^-}]_{total} = [\mathrm{Cl^-}]_{from\, CuCl} + [\mathrm{Cl^-}]_{from\, KCl}$$

Let:
- s = molar solubility of CuCl (moles per liter), i.e., the molar concentration of Cu^+ ions.
- $[\mathrm{Cl^-}]_{from\, CuCl} = s$

Given:
$$[\mathrm{Cl^-}]_{from\, KCl} = 0.060\, M$$

Total chloride concentration:

$$[\mathrm{Cl^-}]_{total} = s + 0.060$$

The Ksp expression:

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

$$1.0 \times 10^{-6.1.0} = s \times (s + 0.060)$$

Step 2: Set Up the Equation and Solve for s

Assuming the Ksp value is 1.0×10^{-6}, the equation simplifies to:

$$1.0 \times 10^{-6} = s (s + 0.060)$$

This is a quadratic in s:

$$s^2 + 0.060s - 1.0 \times 10^{-6} = 0$$

Applying the quadratic formula:

$$s = \frac{ -b \pm \sqrt{b^2 - 4ac} }{2a}$$

Where:
- $a = 1$
- $b = 0.060$
- $c = -1.0 \times 10^{-6}$

Calculating discriminant:

$$\Delta = (0.060)^2 - 4 \times 1 \times (-1.0 \times 10^{-6})$$

$$\Delta = 0.0036 + 4 \times 10^{-6}$$

$$\Delta = 0.0036 + 0.000004 = 0.003604$$

Square root:

$$\sqrt{\Delta} \approx 0.06003$$

Calculating s:

$$s = \frac{ -0.060 \pm 0.06003 }{2}$$

The negative root is physically meaningless for molar solubility, so:

$$s = \frac{ -0.060 + 0.06003 }{2}$$

$$s \approx \frac{0.00003}{2} = 1.5 \times 10^{-5} \, M$$

Therefore, the molar solubility of CuCl in 0.060 M KCl is approximately $1.5 \times 10^{-5} \, M$.

Practical Implications and Summary

Key Points to Remember

- The common ion effect decreases the solubility of salts like CuCl when additional Cl^- ions are present.
- The calculation involves setting up the solubility equilibrium expression and solving the resulting quadratic.
- The molar solubility is significantly lower in solutions with high concentrations of common ions.

Applications of This Calculation

- Designing precipitation reactions in analytical chemistry.
- Understanding mineral solubility in geochemistry.
- Developing pharmaceutical formulations where solubility impacts drug delivery.

Summary of the Process

1. Write the dissociation reaction and the expression for Ksp.
2. Account for the initial concentration of common ions.
3. Set up the quadratic equation to solve for molar solubility.
4. Use the quadratic formula to find the positive root.
5. Interpret the result in the context of solution chemistry.

Conclusion

Determining the molar solubility of CuCl in a solution containing 0.060 M KCl involves understanding the principles of solubility equilibria, the impact of the common ion effect, and careful algebraic calculation. The key takeaway is that the presence of common ions like Cl^- significantly reduces the solubility of sparingly soluble salts, a phenomenon crucial in various scientific and industrial applications. By following the step-by-step approach outlined in this article, students and professionals can accurately estimate solubility values, enhancing their understanding of solution chemistry and its practical applications.

If you need further assistance with solubility calculations, chemical equilibria, or related topics, don't hesitate to explore more chemistry resources or consult with educational experts.

Frequently Asked Questions

What is the molar solubility of CuCl in a 0.060 M KCl solution given Ksp = 1.0 x 10^-6?

The molar solubility of CuCl in the 0.060 M KCl solution is approximately 9.7 x 10^-4 M.

How does the presence of KCl affect the solubility of CuCl?

The KCl provides additional Cl- ions, which suppress the dissolution of CuCl due to the common ion effect, reducing its molar solubility.

What is the role of the common ion effect in calculating the molar solubility of CuCl?

The common ion effect decreases solubility because the additional Cl- ions from KCl shift the dissolution equilibrium of CuCl towards the solid form, lowering its molar solubility.

How do you set up the expression to determine the molar solubility of CuCl in this scenario?

You write the solubility product expression, incorporating the initial KCl concentration to account for the common ion effect, and solve for the molar solubility s.

What is the simplified formula to calculate the molar solubility of CuCl in the presence of KCl?

s = sqrt(Ksp / (1 + [Cl-] from KCl)), where [Cl-] from KCl is 0.060 M, used to adjust the solubility calculation.

What assumptions are made in calculating the molar solubility of CuCl in this context?

Assumptions include that the molar solubility is small compared to the initial KCl concentration, and that activity coefficients are approximated as 1.

Could the molar solubility of CuCl be significantly affected if the KCl concentration was higher?

Yes, increasing KCl concentration would further suppress CuCl solubility due to the common ion effect, resulting in a lower molar solubility.

Why is the value of Ksp important in calculating the solubility of CuCl?

Ksp defines the maximum product of ion concentrations at equilibrium; it is essential for calculating the molar solubility based on the dissolution reaction of CuCl.

Additional Resources

Determining the molar solubility of copper(I) chloride (CuCl) in a potassium chloride (KCl) solution involves understanding complex interactions between solute and solvent, the influence of common ions, and the application of solubility product constants (Ksp). This process is fundamental in inorganic chemistry, especially when analyzing ionic equilibria in aqueous solutions. In this comprehensive review, we will explore the principles underlying solubility calculations, analyze the specific case of CuCl in a 0.060 M KCl solution, and demonstrate the step-by-step method to determine its molar solubility, incorporating the relevant thermodynamic and chemical considerations.

Understanding the Fundamentals of Solubility and Ksp

Solubility and Molar Solubility: Definitions and Significance

Solubility refers to the maximum amount of a substance that can dissolve in a given solvent at a specific temperature to form a saturated solution. It is commonly expressed in grams per liter (g/L) or molarity (mol/L). Molar solubility, in particular, indicates the number of moles of a solute that dissolve per liter of solution to reach saturation.

For ionic compounds like CuCl, the molar solubility is directly related to the dissociation of the compound into its constituent ions:

$$
\text{CuCl (s)} \leftrightarrow \text{Cu}^+ (aq) + \text{Cl}^- (aq)
$$

The molar solubility (denoted as s) corresponds to the molar concentration of Cu^+ and Cl^- ions in the saturated solution.

The Solubility Product Constant (Ksp)

The solubility product constant (K_{sp}) characterizes the equilibrium between a sparingly soluble salt and its ions in solution:

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

For CuCl, K_{sp} gives the maximum product of ion concentrations at equilibrium before the salt begins to precipitate out of solution.

In this scenario, the given K_{sp} is 1.0×10^{-6}, indicating very low solubility. This small value signifies the compound is only slightly soluble in water under standard conditions, which makes the effect of common ions and solution conditions particularly significant.

Impact of Common Ions on Solubility

The Common Ion Effect

The presence of ions already in the solution can significantly influence the solubility of a salt. This phenomenon, known as the common ion effect, suppresses the dissolution of the salt due to Le Châtelier's principle.

In this case, the solution contains 0.060 M KCl, which introduces additional Cl^- ions into the medium. Since Cl^- is a common ion with CuCl, the increased chloride ion concentration shifts the equilibrium position, reducing the molar solubility of CuCl.

This effect must be quantitatively incorporated into the calculations to accurately determine the molar solubility of CuCl in this environment.

The Role of KCl in the Solution

Potassium chloride (KCl) is a strong electrolyte that dissociates completely:

$$
\text{KCl} \rightarrow \text{K}^+ + \text{Cl}^-
$$

The resulting chloride ion concentration in solution is elevated, contributing to the common ion effect. The potassium ion (K^+) is generally considered inert in this context because it does not participate in the equilibrium of CuCl dissolution.

The key focus is on the chloride ion concentration, which impacts the solubility equilibrium of CuCl.

Step-by-Step Calculation of Molar Solubility

Establishing the Equilibrium Expression

The dissociation of CuCl can be represented as:

$$
\text{CuCl (s)} \leftrightarrow \text{Cu}^+ (aq) + \text{Cl}^- (aq)
$$

At equilibrium, the concentrations of ions are:
- $[\text{Cu}^+] = s$
- $[\text{Cl}^-] = s + [\text{Cl}^-]_{\text{initial}}$

Given that the initial chloride ion concentration is due to KCl:
$$
[\text{Cl}^-]_{\text{initial}} = 0.060\, \text{M}
$$

The solubility product expression becomes:
$$
K_{sp} = [\text{Cu}^+][\text{Cl}^-] = s \times (s + 0.060)
$$

Since s is expected to be very small (due to the low K_{sp}), the term $s + 0.060$ can be approximated as 0.060 for practical calculation purposes, simplifying the expression to:
$$
K_{sp} \approx s \times 0.060
$$

This approximation is valid because $s \ll 0.060$.

Calculating the Molar Solubility

Using the approximation:
$$
s = \frac{K_{sp}}{0.060}
$$
substituting the known K_{sp}:
$$
s = \frac{1.0 \times 10^{-6}}{0.060} = \frac{1.0 \times 10^{-6}}{6.0 \times 10^{-2}} \approx 1.67 \times 10^{-5} \, \text{M}
$$

This result indicates that the molar solubility of CuCl in a 0.060 M KCl solution is approximately 1.67×10^{-5} mol/L.

Refining the Calculation: Considering the Exact Expression

While the approximation provides a quick estimate, for greater accuracy, the exact quadratic solution should be considered:
$$
K_{sp} = s (s + 0.060)
$$
which rearranges to:
$$
s^2 + 0.060 s - K_{sp} = 0
$$

Applying the quadratic formula:

$$
s = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}
$$

where:

- $a = 1$
- $b = 0.060$
- $c = -K_{sp} = -1.0 \times 10^{-6}$

Calculating the discriminant:

$$
\Delta = (0.060)^2 - 4 \times 1 \times (-1.0 \times 10^{-6}) = 3.6 \times 10^{-3} + 4.0 \times 10^{-6} = 3.604 \times 10^{-3}
$$

Taking the square root:

$$
\sqrt{\Delta} \approx 0.06002
$$

Now, the solution for s:

$$
s = \frac{-0.060 + 0.06002}{2} \approx \frac{0.00002}{2} = 1.0 \times 10^{-5} \, \text{M}
$$

This refined calculation yields a molar solubility of approximately 1.0×10^{-5} mol/L, slightly lower than the approximation but within similar magnitude, confirming the impact of the common ion effect.

Implications of the Findings and Broader Context

Effect of the Common Ion on Solubility

The calculations vividly demonstrate how the presence of common ions in solution suppresses the solubility of sparingly soluble salts. The 0.060 M of chloride ions from KCl significantly diminishes the molar solubility of CuCl compared to its solubility in pure water, which can be estimated by setting the chloride ion concentration solely to s (i.e., ignoring the initial chloride).

In pure water:

$$
K_{sp} = s^2 \Rightarrow s = \sqrt{K_{sp}} = \sqrt{1.0 \times 10^{-6}} = 1.0 \times 10^{-3} \, \text{M}
$$

which is about 100 times larger than in the KCl solution, emphasizing the profound influence of common ions.

Applications and Practical Significance

Understanding the molar solubility of compounds like CuCl in the presence of other ions is crucial in various fields:
- Analytical Chemistry: Accurate determination of metal ions in complex matrices.
- Environmental Science: Predicting metal ion solubility and mobility in natural waters containing various salts.
- Industrial Processes: Controlling precipitation and dissolution in manufacturing, electroplating, and wastewater treatment.
- Pharmaceuticals and Material Science: Designing solutions with controlled solubility profiles to optimize product stability and efficacy.

Conclusion

The determination of the molar solubility of CuCl in a 0.060 M KCl solution exemplifies fundamental principles of chemical equilibria, solubility, and ionic interactions. By understanding how common ions influence solub

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