In A Saturated Solution That Is In Contact With Solid Mg(oh)2, The Concentration Of Mg2 Is 3.7 105 M.

In A Saturated Solution That Is In Contact With Solid Mg(OH)2, The Concentration Of Mg^{2+} Is 3.7×10^{-5} M.

Introduction to Magnesium Hydroxide and Its Saturation Properties

Magnesium hydroxide, $Mg(OH)_2$, is an important inorganic compound with diverse applications ranging from antacids to wastewater treatment. Understanding its solubility behavior, especially in saturated solutions, is critical for various scientific and industrial processes. When a solution is saturated with $Mg(OH)_2$, it contains the maximum amount of dissolved Mg^{2+} and OH^- ions that can exist in equilibrium with the solid phase. The concentration of Mg^{2+} ions, as indicated, is 3.7×10^{-5} M, providing essential insight into the solubility product (Ksp) and the equilibrium dynamics of the system.

Fundamentals of Solubility Equilibrium

Solubility Product Constant (Ksp)

The solubility product constant, Ksp, quantifies the solubility of ionic compounds like $Mg(OH)_2$ in water. It is defined as the equilibrium expression for the dissolution process:

$Mg(OH)_2 (s) \rightleftharpoons Mg^{2+} (aq) + 2 OH^- (aq)$

The Ksp expression:

$Ksp = [Mg^{2+}][OH^-]^2$

where the brackets denote molar concentrations at equilibrium.

Significance of Ksp

- Determines the maximum solubility of the compound in water.
- Helps predict whether a precipitate will form when two solutions are mixed.
- Assists in calculating ion concentrations at equilibrium.

Calculating the Solubility of $Mg(OH)_2$

Given the concentration of Mg^{2+} ions in the saturated solution, we can derive the solubility of $Mg(OH)_2$.

Known Data:

- $[Mg^{2+}] = 3.7 \times 10^{-5}$ M
- The dissolution releases 1 mol of Mg^{2+} per mol of $Mg(OH)_2$ dissolved.

Let's denote the solubility of $Mg(OH)_2$ as S (in mol/L). Since each mole of $Mg(OH)_2$ yields 1 mole of Mg^{2+} and 2 moles of OH^-:

- $[Mg^{2+}] = S$
- $[OH^-] = 2S$

From the given data:

$S = 3.7 \times 10^{-5}$ M

Therefore, the hydroxide ion concentration:

$[OH^-] = 2 \times 3.7 \times 10^{-5} = 7.4 \times 10^{-5}$ M

Calculating Ksp

Using the equilibrium expression:

$K_{sp} = [Mg^{2+}][OH^-]^2$

Plugging in the values:

$K_{sp} = (3.7 \times 10^{-5}) \times (7.4 \times 10^{-5})^2$

Calculating:

$Ksp = 3.7 \times 10^{-5} \times (5.476 \times 10^{-9})$

$Ksp \approx 3.7 \times 5.476 \times 10^{-14}$

$Ksp \approx 20.25 \times 10^{-14}$

$Ksp \approx 2.025 \times 10^{-13}$

This value aligns with typical literature values for $Mg(OH)_2$, confirming the accuracy of the given ion concentration.

Factors Affecting the Solubility of $Mg(OH)_2$

Understanding what influences the solubility of magnesium hydroxide is essential for controlling its behavior in various settings.

pH of the Solution

- Acidic conditions increase solubility by shifting equilibrium toward dissolution.
- Basic conditions decrease solubility by suppressing ionization.

Common Ion Effect

- Presence of additional Mg^{2+} or OH^- ions from other sources reduces $Mg(OH)_2$ solubility due to Le Châtelier’s principle.

Temperature

- $Mg(OH)_2$ is relatively insoluble; temperature changes have minimal impact, but in some cases, solubility may slightly increase with temperature.

Presence of Complexing Agents

- Certain ligands can form complexes with Mg^{2+}, enhancing solubility.

Industrial and Environmental Significance

Application in Water Treatment

- $Mg(OH)_2$ is used to precipitate and remove heavy metals and other impurities.
- Its low solubility makes it effective for neutralization and pH adjustment.

Medical Uses

- As an antacid, $Mg(OH)_2$ neutralizes stomach acid effectively due to its insolubility and buffering capacity.

Environmental Impact

- Understanding $Mg(OH)_2$ solubility is vital for managing soil and water pH in environmental remediation efforts.

Practical Considerations in Handling $Mg(OH)_2$ Solutions

Preparation of Saturated Solutions

- Dissolve solid $Mg(OH)_2$ in water until no more dissolves.
- Maintain stirring and control temperature to ensure equilibrium.

Measuring Ion Concentrations

- Use techniques like atomic absorption spectroscopy (AAS) or inductively coupled plasma (ICP) analysis for precise measurements.
- Ensure solutions are properly filtered to avoid undissolved particles.

Safety Precautions

- $Mg(OH)_2$ is relatively safe but can cause skin irritation.
- Handle with appropriate protective gear and dispose of waste according to safety regulations.

Conclusion and Summary

The concentration of Mg^{2+} ions in a saturated $Mg(OH)_2$ solution is a critical parameter, with a typical value of 3.7×10^{-5} M indicating the solubility limit under standard conditions. By understanding this concentration, scientists and engineers can calculate the solubility product (Ksp), predict the behavior of magnesium hydroxide in various environments, and optimize its applications across industries. The equilibrium dynamics of $Mg(OH)_2$, influenced by factors such as pH, temperature, and presence of other ions, are fundamental to leveraging its properties effectively. Whether in water treatment, medicine, or environmental management, mastering the principles of solubility and saturation of $Mg(OH)_2$ is essential for achieving desired outcomes while maintaining safety and efficiency.

Keywords: $Mg(OH)_2$, magnesium hydroxide, solubility, saturation, ion concentration, Ksp, equilibrium, water treatment, industrial applications, environmental safety

Frequently Asked Questions

What is the significance of the Mg^{2+} concentration being 3.7×10^{5} M in a saturated $Mg(OH)_2$ solution?

This high concentration indicates the solution has reached its saturation point, where no more $Mg(OH)_2$ can dissolve, and the Mg^{2+} ions are in equilibrium with the solid phase.

How can the solubility product constant (Ksp) for $Mg(OH)_2$ be calculated from the given Mg^{2+} concentration?

Since $Mg(OH)_2$ dissociates as $Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$, $Ksp = [Mg^{2+}][OH^-]^2$. Given $[Mg^{2+}] = 3.7 \times 10^{5}$ M, if $[OH^-]$ is known, Ksp can be calculated accordingly.

Is the given Mg^{2+} concentration consistent with typical solubility values of $Mg(OH)_2$?

No, a Mg^{2+} concentration of 3.7×10^{5} M is unrealistically high; typical solubility values are around 1.2×10^{-5} M, suggesting a possible error or different conditions.

What does the contact between the saturated solution and solid $Mg(OH)_2$ imply about the system's equilibrium?

It indicates that the system has reached dynamic equilibrium, where the rate of $Mg(OH)_2$ dissolving equals the rate of precipitation, maintaining constant concentrations.

How would adding more solid $Mg(OH)_2$ affect the concentration of Mg^{2+} in the solution?

Since the solution is already saturated, adding more solid $Mg(OH)_2$ would not increase Mg^{2+} concentration; instead, it would remain at equilibrium until conditions change.

What factors could influence the solubility of $Mg(OH)_2$ in water?

Temperature, pH, ionic strength of the solution, and presence of complexing agents can all affect the solubility of $Mg(OH)_2$.

How does pH affect the solubility of $Mg(OH)_2$ in solution?

Higher pH (more basic conditions) generally decrease $Mg(OH)_2$ solubility because excess OH^- shifts the equilibrium towards the solid phase, reducing Mg^{2+} concentration.

Can the given Mg^{2+} concentration be used to determine the pH of the solution?

Not directly, but if the solubility and Ksp are known, and assuming equilibrium, the pOH (and thus pH) can be calculated from the hydroxide ion concentration, which relates to Mg^{2+} concentration.

What practical applications involve saturated $Mg(OH)_2$ solutions with known Mg^{2+} concentrations?

Such solutions are used in water treatment, antacids, and as a source of magnesium in chemical manufacturing, where precise solubility and concentration control are important.

Additional Resources

In a saturated solution that is in contact with solid $Mg(OH)_2$, the concentration of Mg^{2+} is 3.7×10^{-5} M. This statement encapsulates a fundamental aspect of solubility equilibria involving magnesium hydroxide, a sparingly soluble salt. Understanding the implications of this concentration, the underlying chemistry, and how it influences various processes is essential for chemists, environmental scientists, and engineers alike. This review aims to explore the detailed chemistry behind this scenario, analyze factors affecting the solubility, and discuss practical applications and limitations.

Understanding the Solubility of $Mg(OH)_2$

Basic Chemistry of Magnesium Hydroxide

Magnesium hydroxide, $Mg(OH)_2$, is an inorganic compound characterized by its low solubility in water.

It is often used as an antacid and in water treatment due to its ability to neutralize acids and precipitate impurities. Its solubility product constant (Ksp) is a key parameter that quantifies its solubility equilibrium:

$$ \text{Mg(OH)}_2 (s) \leftrightarrow \text{Mg}^{2+} (aq) + 2 \text{OH}^- (aq) $$

The solubility product constant (Ksp) for $Mg(OH)_2$ at 25°C is approximately 1.8×10^{-11}. This low value indicates that only a tiny amount of $Mg(OH)_2$ dissolves in water under ideal conditions, reaching a dynamic equilibrium where the rate of dissolution equals the rate of precipitation.

Implications of the Given Mg^{2+} Concentration

Given that the Mg^{2+} concentration in the saturated solution is 3.7×10^{-5} M, it suggests that the solution is at or near equilibrium with solid $Mg(OH)_2$. This concentration is consistent with the solubility limit dictated by the Ksp:

$$
K_{sp} = [\text{Mg}^{2+}][\text{OH}^-]^2
$$

Substituting the known Mg^{2+} concentration:

$$
1.8 \times 10^{-11} = (3.7 \times 10^{-5}) \times [\text{OH}^-]^2
$$

Solving for $[OH^-]$:

$$
[\text{OH}^-]^2 = \frac{1.8 \times 10^{-11}}{3.7 \times 10^{-5}} \approx 4.86 \times 10^{-7}
$$

$$
[\text{OH}^-] \approx \sqrt{4.86 \times 10^{-7}} \approx 6.97 \times 10^{-4}\, \text{M}
$$

Thus, the hydroxide ion concentration is approximately 6.97×10^{-4} M, which aligns with expectations for a saturated $Mg(OH)_2$ solution.

Factors Influencing the Solubility of $Mg(OH)_2$

pH and Ionic Environment

The solubility of $Mg(OH)_2$ is heavily influenced by the pH of the solution. Since it is an amphoteric hydroxide, its solubility increases in acidic conditions due to protonation of hydroxide ions, shifting the equilibrium towards dissolution. Conversely, in basic solutions with high [OH^-], the solubility decreases due to the common ion effect, which suppresses further dissolution.

Key points:

- Acidic pH increases $Mg(OH)_2$ solubility by consuming hydroxide ions, shifting equilibrium right.
- Basic pH suppresses solubility via the common ion effect, reducing Mg^{2+} release.

Features & Pros:

- Controlled solubility allows $Mg(OH)_2$ to be used in applications requiring limited dissolution.
- Buffering capacity in environmental systems: $Mg(OH)_2$ can neutralize acids efficiently.

Cons:

- In highly alkaline environments, $Mg(OH)_2$'s solubility is minimal, which can limit its efficacy for certain applications.
- pH-dependent solubility means precise control is necessary in processes like water treatment.

Temperature Effects

Temperature influences solubility, though $Mg(OH)_2$ exhibits only slight solubility changes with temperature variations. Typically, increasing temperature can marginally increase solubility, but the effect is minor due to the low Ksp.

- Higher temperatures may slightly increase $Mg(OH)_2$ solubility, enhancing dissolution rates.
- Temperature fluctuations can affect equilibrium concentrations, important in industrial processes.

Practical Applications and Implications

Water Treatment

Magnesium hydroxide is widely used as a neutralizing agent in water treatment facilities. Its low solubility ensures gradual pH adjustments, precipitating impurities and controlling acidity.

Features:

- Acts as a neutralizer, raising pH in acidified water systems.
- Precipitates heavy metals and other contaminants through co-precipitation.

Pros:

- Eco-friendly and relatively inexpensive.
- Produces minimal sludge compared to other hydroxides like $Al(OH)_3$.

Cons:

- Limited solubility requires sufficient contact time.
- Precise dosing is necessary to avoid over-alkalization.

Biomedical and Pharmaceutical Uses

$Mg(OH)_2$'s properties make it suitable as an antacid to neutralize stomach acid, providing relief from indigestion. Its low solubility ensures a slow release of Mg^{2+} ions, prolonging its buffering effect.

Features:

- Well-tolerated and safe for human consumption.
- Provides magnesium supplementation.

Pros:

- Slow dissolution minimizes side effects.
- Neutralizes stomach acid effectively.

Cons:

- Limited solubility means rapid relief might require adjuncts.
- Excessive intake can cause magnesium overdose symptoms.

Limitations and Challenges

While $Mg(OH)_2$ offers several advantages, certain limitations are inherent due to its low solubility:

- Limited Dissolution Rate: Slow dissolution can impede processes requiring rapid neutralization or magnesium release.
- Sensitivity to pH: Variability in environmental pH can alter solubility, complicating process control.
- Precipitation Risks: In systems where Mg^{2+} concentration rises beyond equilibrium, unwanted precipitation can occur, leading to scaling or clogging.

Conclusion

The statement that in a saturated solution in contact with solid $Mg(OH)_2$, the Mg^{2+} concentration is 3.7×10^{-5} M encapsulates fundamental principles of solubility equilibria. Magnesium hydroxide's low solubility is a defining feature that influences its applications across environmental, industrial, and biomedical fields. Understanding the factors affecting this equilibrium—such as pH, temperature, and ionic strength—is crucial for optimizing its use. While its limited solubility presents some challenges, it also offers benefits like controlled release and minimal sludge formation. Future developments in process engineering may focus on enhancing dissolution rates or modifying $Mg(OH)_2$ properties to expand its utility further. Overall, $Mg(OH)_2$ remains a versatile compound whose solubility profile is pivotal to its diverse applications.

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