

Solubility Equilibria

How much of an 'insoluble' salt actually dissolves, and when does a precipitate form?

●●○ Intermediate

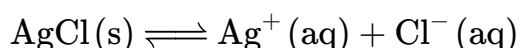
Equilibrium

Last reviewed 6 October 2026

Before you start: it helps to understand [Chemical Equilibrium and K_c](#) and [Precipitation Reactions and Net Ionic Equations](#).

What is it?

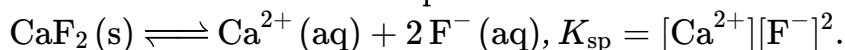
“Insoluble” salts are really **very slightly soluble**. When solid silver chloride is placed in water, a tiny amount dissolves until the solution is **saturated**. Then a dynamic equilibrium is set up between the solid and its ions:



The equilibrium constant is the **solubility product**, K_{sp} :

$$K_{\text{sp}} = [\text{Ag}^+][\text{Cl}^-] = 1.8 \times 10^{-10}$$

The solid does not appear in the expression (its concentration is constant). Each ion concentration is raised to the power of its coefficient: for



🔑 KEY IDEA

K_{sp} is the **ceiling** for the product of the ion concentrations. If the ion product Q is **below** K_{sp} , more solid can dissolve. If Q is **above** K_{sp} , the solution is supersaturated and a **precipitate forms** until $Q = K_{\text{sp}}$.

Why does it matter?

- **Health.** Barium sulfate is swallowed for X-ray imaging of the gut: barium ions are toxic, but BaSO_4 is so insoluble ($K_{\text{sp}} \approx 10^{-10}$) that almost none dissolves. Kidney stones form when calcium oxalate exceeds its K_{sp} .
- **Teeth.** Fluoride converts tooth enamel into a less soluble mineral, so it dissolves less in acid.
- **Analysis and clean-up.** Chemists precipitate one ion while leaving others in solution, and remove heavy metals from waste water.

How does it work?

1. Molar solubility

The **molar solubility**, s , is the number of moles that dissolve per litre of saturated solution. Write the ion concentrations in terms of s :

Salt	Ions	K_{sp} in terms of s
AgCl, BaSO ₄ (1 : 1)	s and s	$K_{sp} = s^2$
CaF ₂ , PbI ₂ , Mg(OH) ₂ (1 : 2)	s and $2s$	$K_{sp} = s(2s)^2 = 4s^3$
Ag ₂ CrO ₄ (2 : 1)	$2s$ and s	$K_{sp} = 4s^3$

Only compare K_{sp} values directly for salts with the **same** ion ratio; otherwise, compare their molar solubilities.

2. The common-ion effect

If the solution already contains one of the ions, the equilibrium shifts to the left (Le Chatelier's principle), and **less** solid dissolves. Silver chloride is far less soluble in salt water than in pure water.

3. Will a precipitate form?

1. Find each ion concentration **after mixing** (allow for the total volume).
2. Calculate the ion product Q in the same form as K_{sp} .
3. Compare: $Q > K_{sp}$: precipitate forms. $Q < K_{sp}$: no precipitate. $Q = K_{sp}$: saturated.

∞ THINK OF IT LIKE THIS

K_{sp} is like the capacity of a crowded room. While there is space (Q below K_{sp}), more people can come in. Once it is full ($Q = K_{sp}$), anyone else who arrives must leave again. If you already filled the room with one group (a common ion), there is less space for anyone else.

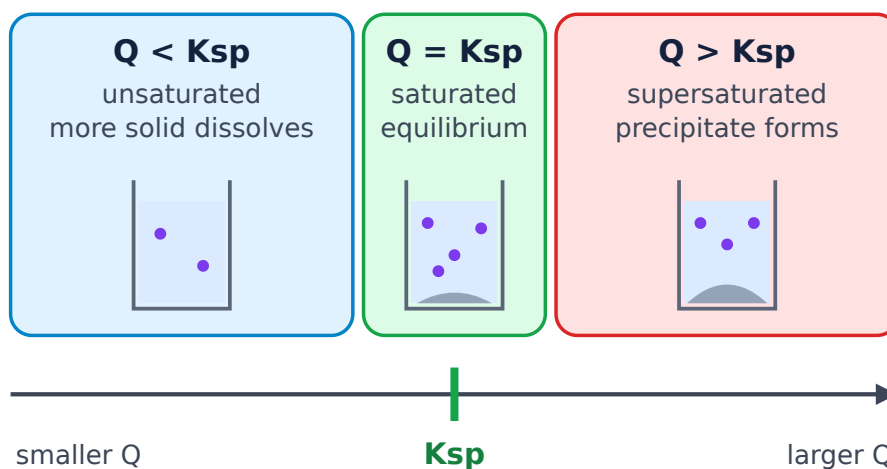
Q MORE PRECISELY

K_{sp} depends on temperature. Simple K_{sp} calculations ignore effects such as ion pairing and reactions of the ions with water: for example, the solubility of salts of weak acids (carbonates, phosphates, fluorides) **increases** in acid, because H^+ removes the anion. Strictly, equilibrium constants use activities rather than concentrations, which matters in concentrated solutions.

Visualise it

Ion product Q versus K_{sp}

Increasing ion concentrations $\rightarrow$



Comparing the ion product Q with K_{sp} predicts what happens.

Worked example

WORKED EXAMPLE: Molar solubility of AgCl

Question: Calculate the molar solubility of AgCl ($K_{sp} = 1.8 \times 10^{-10}$) in mol/L and in g/L.

Step 1 $[\text{Ag}^+] = [\text{Cl}^-] = s$, so $K_{sp} = s^2$.

Step 2 Molar solubility:

$$\begin{aligned}s &= \sqrt{1.8 \times 10^{-10}} \\ &= 1.3 \times 10^{-5} \text{ mol/L}\end{aligned}$$

Step 3 In g/L: $1.34 \times 10^{-5} \text{ mol/L} \times 143.32 \text{ g/mol} = 1.9 \times 10^{-3} \text{ g/L}$, about 2 mg per litre.

 WORKED EXAMPLE: A 1 : 2 salt

Question: Calculate the molar solubility of CaF_2 ($K_{\text{sp}} = 3.9 \times 10^{-11}$).

Step 1 $[\text{Ca}^{2+}] = s$ and $[\text{F}^-] = 2s$, so $K_{\text{sp}} = s(2s)^2 = 4s^3$.

Step 2 Solve:

$$\begin{aligned}s &= \sqrt[3]{\frac{3.9 \times 10^{-11}}{4}} \\ &= 2.1 \times 10^{-4} \text{ mol/L}\end{aligned}$$

 WORKED EXAMPLE: K_{sp} from solubility

Question: The solubility of PbI_2 is 0.62 g/L. Calculate K_{sp} .

Step 1 Molar solubility: $s = \frac{0.62 \text{ g/L}}{461.0 \text{ g/mol}} = 1.34 \times 10^{-3} \text{ mol/L}$

Step 2 $[\text{Pb}^{2+}] = s$, $[\text{I}^-] = 2s$:

$$\begin{aligned}K_{\text{sp}} &= 4s^3 = 4 \times (1.345 \\ &\quad \times 10^{-3})^3 \\ &= 9.7 \times 10^{-9}\end{aligned}$$

 WORKED EXAMPLE: The common-ion effect

Question: Calculate the molar solubility of AgCl in 0.10 mol/L NaCl .

Step 1 $[\text{Cl}^-] \approx 0.10 \text{ mol/L}$ (from the NaCl); $[\text{Ag}^+] = s$.

Step 2 $K_{\text{sp}} = s \times 0.10$, so

$$\begin{aligned}s &= \frac{1.8 \times 10^{-10}}{0.10} \\ &= 1.8 \times 10^{-9} \text{ mol/L}\end{aligned}$$

Step 3 That is about **7500 times less** than in pure water ($1.3 \times 10^{-5} \text{ mol/L}$).

 WORKED EXAMPLE: Will a precipitate form?

Question: Equal volumes of 0.0010 mol/L AgNO_3 and 0.0010 mol/L NaCl are mixed. Does AgCl precipitate?

Step 1 Mixing equal volumes halves each concentration: $[\text{Ag}^+] = [\text{Cl}^-] = 5.0 \times 10^{-4} \text{ mol/L}$.

Step 2 Ion product:

$$\begin{aligned} Q &= (5.0 \times 10^{-4})(5.0 \\ &\quad \times 10^{-4}) \\ &= 2.5 \times 10^{-7} \end{aligned}$$

Step 3 $Q = 2.5 \times 10^{-7}$ is greater than $K_{\text{sp}} = 1.8 \times 10^{-10}$: **a precipitate forms.**

Common mistake

 **COMMON MISTAKE:** Forgetting to double (and square) the second ion

For CaF_2 , $[\text{F}^-] = 2s$, and it is squared in K_{sp} : $K_{\text{sp}} = s(2s)^2 = 4s^3$, not s^2 or $2s^2$.

 **COMMON MISTAKE:** Comparing K_{sp} values of different types of salt

A smaller K_{sp} does not always mean a less soluble salt. Ag_2CrO_4 ($K_{\text{sp}} = 1.1 \times 10^{-12}$) is actually **more** soluble than AgCl (1.8×10^{-10}): 6.5×10^{-5} versus $1.3 \times 10^{-5} \text{ mol/L}$. Compare molar solubilities.

 **COMMON MISTAKE:** Ignoring dilution when solutions are mixed

Mixing two solutions increases the volume, so every concentration falls. Recalculate the concentrations before working out Q .

 NOTATION NOTE

- Solids are left out of K_{sp} expressions; ion concentrations are in mol/L at equilibrium.
- K_{sp} values are usually quoted without units.

Remember this

REMEMBER THIS

- K_{sp} = product of the ion concentrations in a saturated solution, each raised to its coefficient.
- 1 : 1 salts: $K_{\text{sp}} = s^2$. 1 : 2 or 2 : 1 salts: $K_{\text{sp}} = 4s^3$.
- A common ion lowers the solubility.
- $Q > K_{\text{sp}}$: precipitate. $Q < K_{\text{sp}}$: no precipitate. Remember to allow for dilution on mixing.

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

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