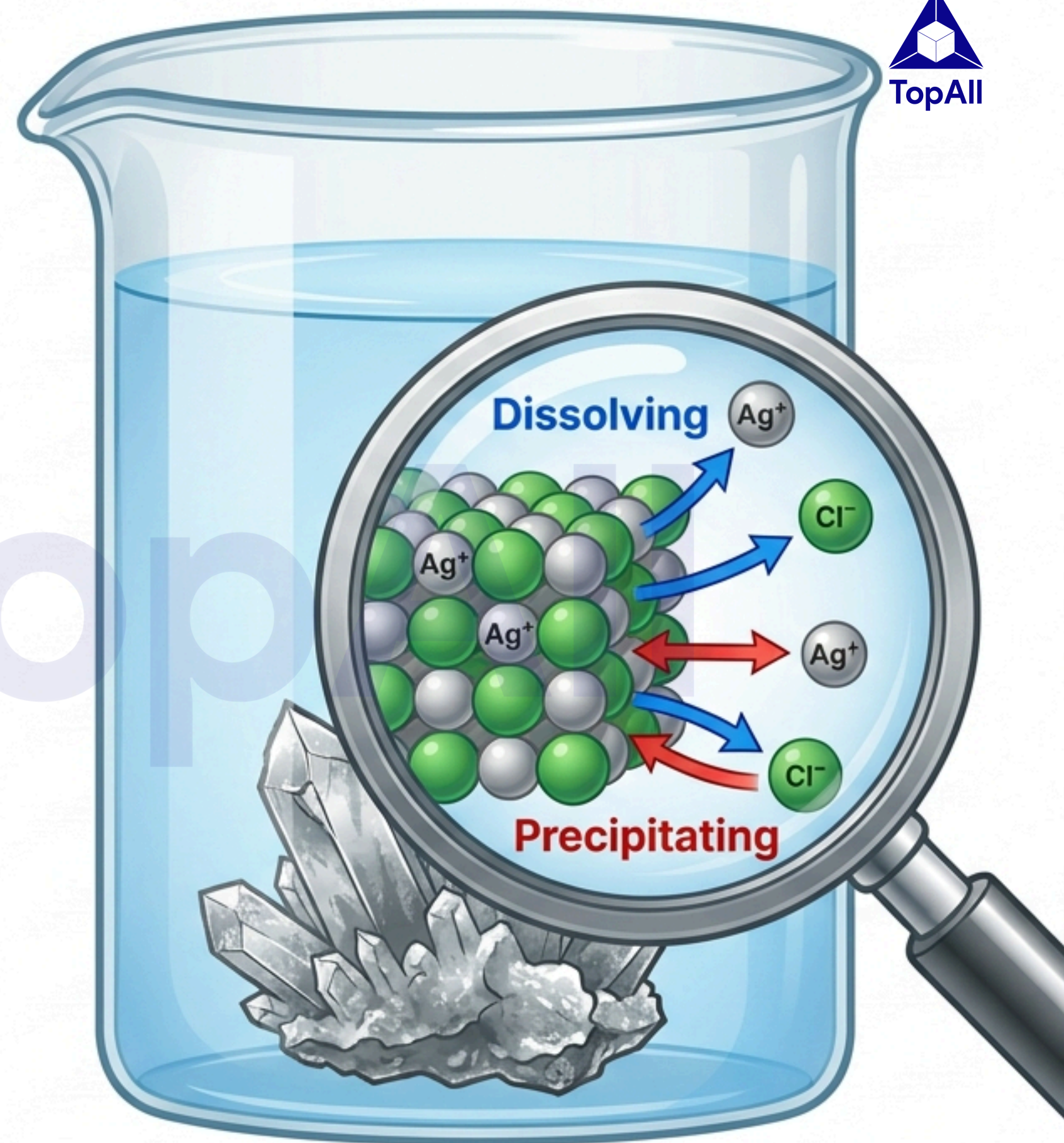


Why Some Salts Dissolve and Some Precipitate

Mastering Solubility Equilibrium & Solubility Product (K_{sp})

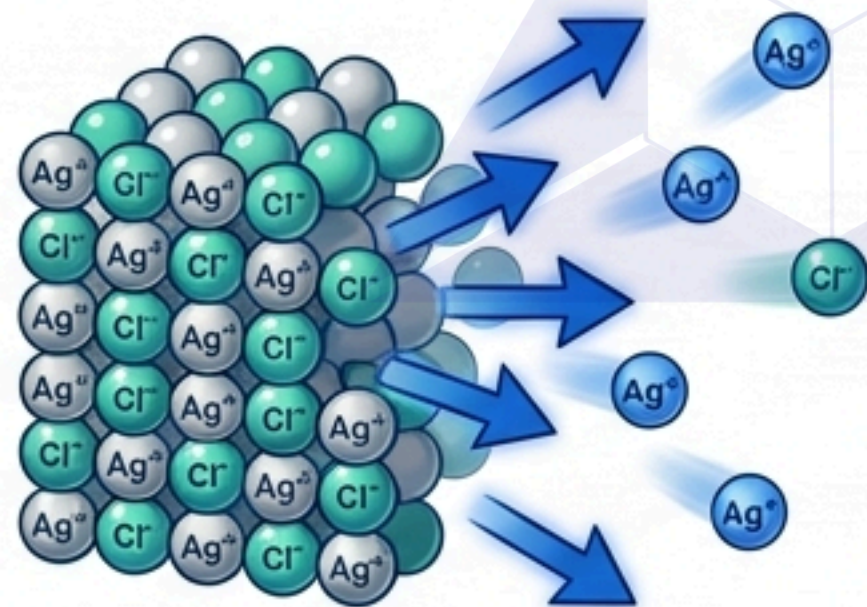


The Hidden Tug-of-War

Solubility is not a one-way street. It is a continuous, dynamic competition between two opposing forces.

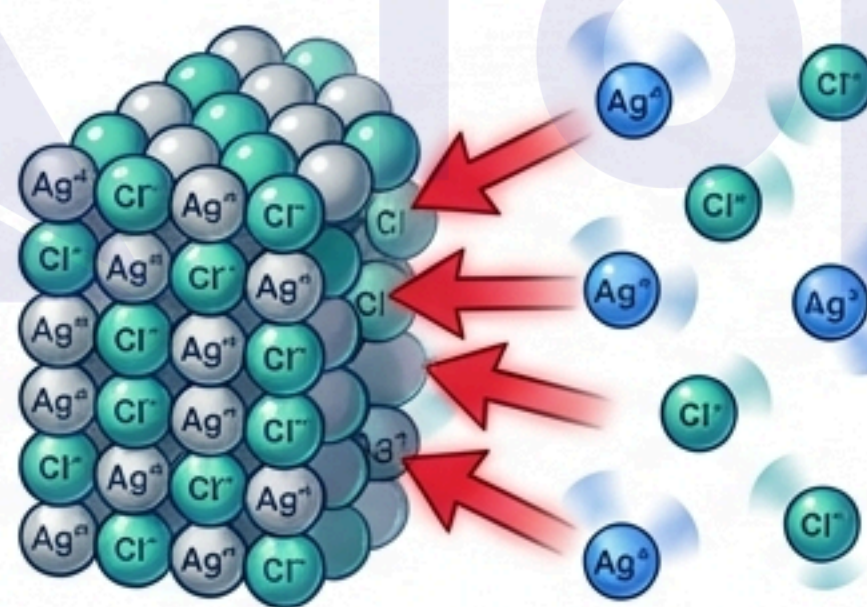


Dissolution (Forward Rate)



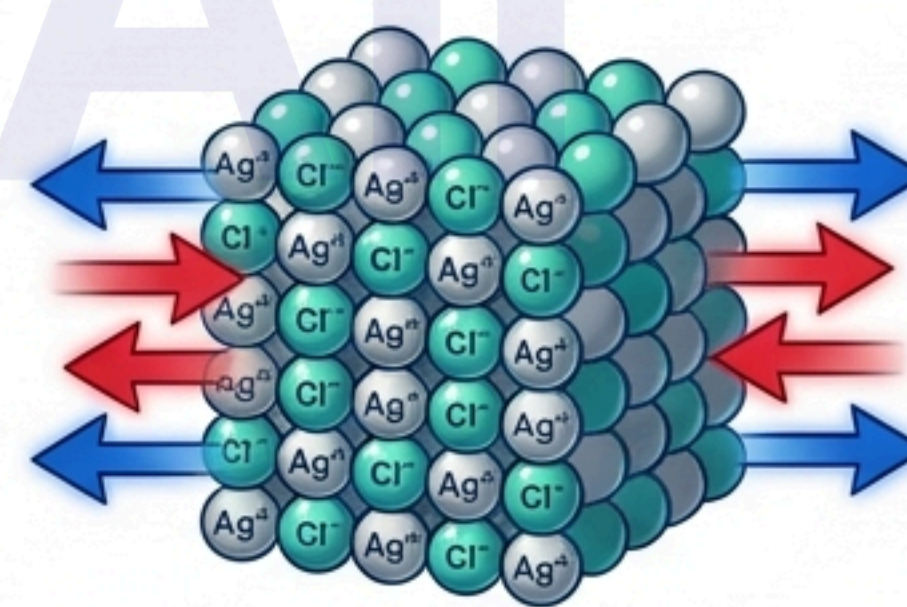
Ions break free from the solid lattice into the surrounding solvent.

Recrystallisation (Backward Rate)



Dissolved ions collide with the solid surface and reattach.

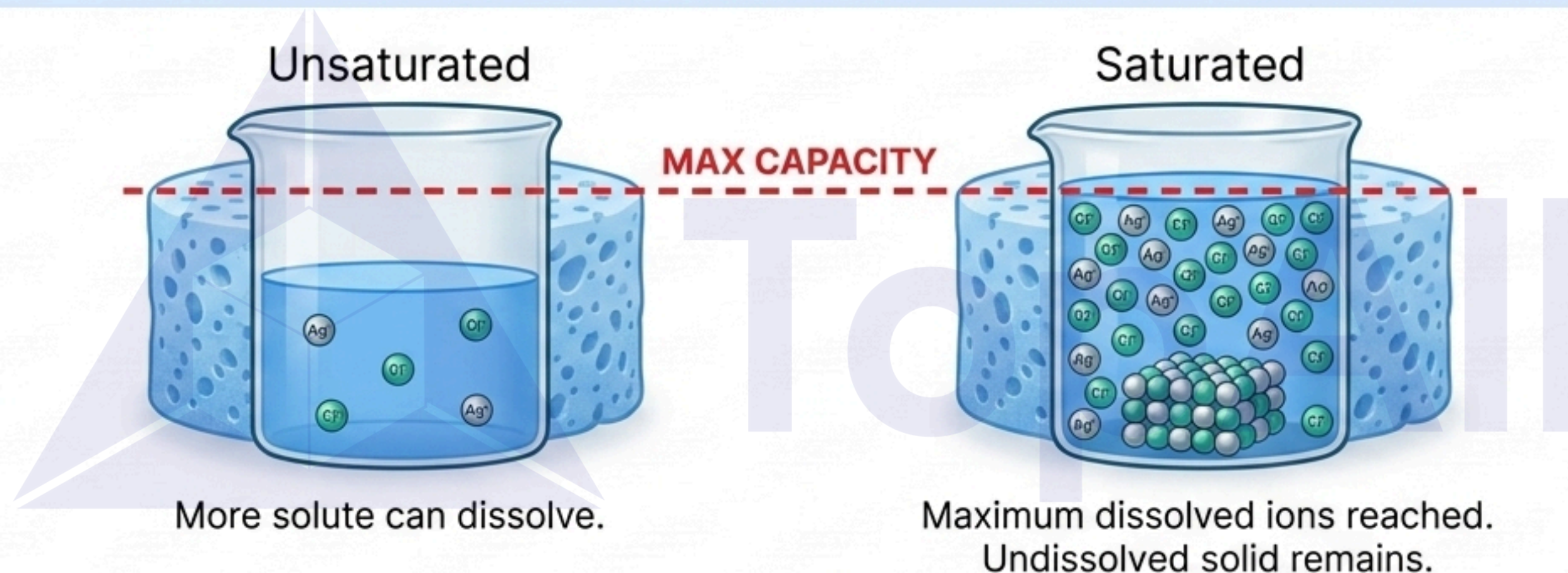
Dynamic Balance



Rate of Dissolution = Rate of Recrystallisation.
The system looks static macroscopically but is highly active microscopically.

The Saturated Solution

Every solvent has a strict maximum holding capacity for dissolved ions at a specific temperature.



The Golden Rule of Equilibrium

Solubility equilibrium ONLY exists in a **saturated solution** where pure solid is in direct contact with its dissolved ions.



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$$E=mc^2$$



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Translating Physical Reality to Math

Dependent ONLY on temperature.

Solubility Product Constant



Concentrations of dissolved ions strictly at equilibrium (in mol/L).



Why is $AgCl_{(s)}$ missing?

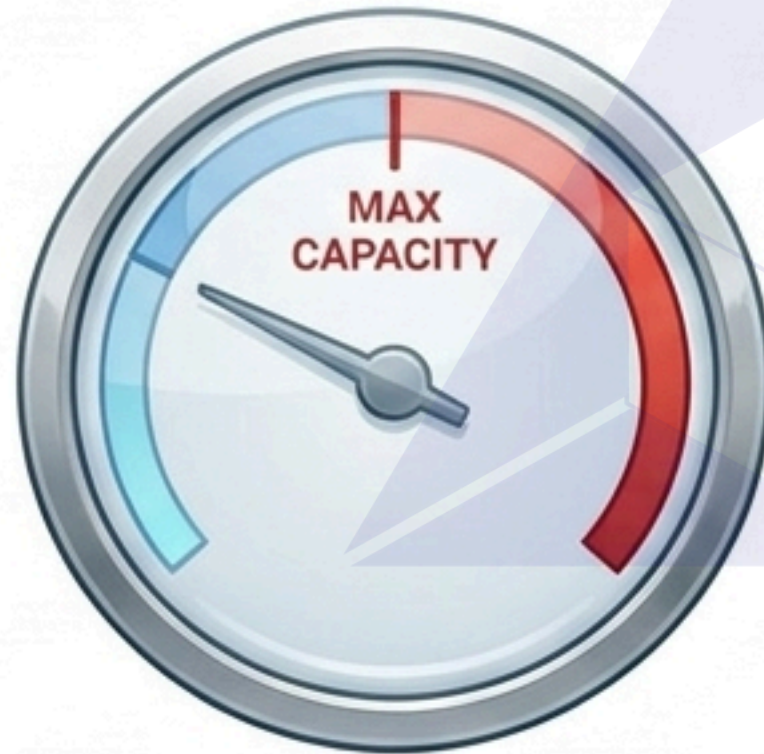
Pure solids have a constant active mass (concentration). They are mathematically absorbed into the equilibrium constant and **EXCLUDED** from the K_{sp} expression!

The Predictive Engine: Q_{sp} vs K_{sp}

K_{sp} = Maximum capacity limit (at equilibrium).

Q_{sp} = Actual current conditions (Ionic Product at any given moment).

$$Q_{sp} < K_{sp}$$



Unsaturated Solution

No Precipitation
(More solid can actively dissolve).

$$Q_{sp} = K_{sp}$$



Saturated Solution

Dynamic Equilibrium
(No net change in the system).

$$Q_{sp} > K_{sp}$$



Supersaturated Solution

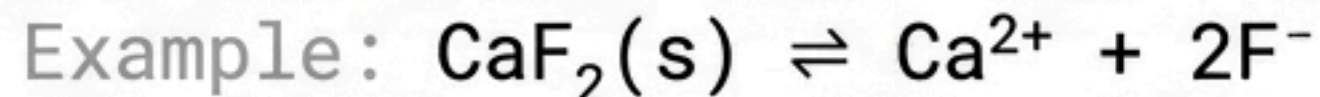
Precipitation
(Excess dissolved ions are forced to form a solid).

JEE Exam Strategy: Predicting Precipitation



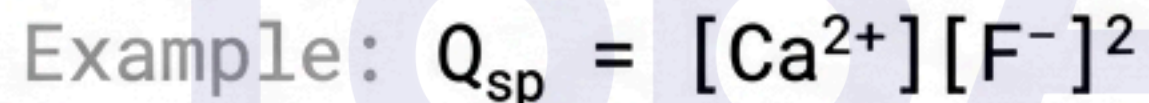
Step 1: Write Ion Equation

Write the balanced dissociation equation for the sparingly soluble salt.



Step 2: Calculate Q_{sp}

Find the initial concentrations of mixed ions. Raise them to their respective stoichiometric coefficients.



Step 3: Compare with K_{sp}

Check the newly calculated Q_{sp} against the provided K_{sp} value.



Step 4: Predict Outcome

If $Q_{\text{sp}} > K_{\text{sp}}$, precipitation occurs immediately.

High-Yield Misconception



The Trap: "Salt A has a lower K_{sp} than Salt B, therefore Salt A will ALWAYS precipitate first."



WRONG MINDSET

"Lower K_{sp} = Automatic precipitation."

This totally ignores stoichiometry! You cannot directly compare K_{sp} values if salts produce different numbers of ions (e.g., comparing AgCl with Ag_2CrO_4).



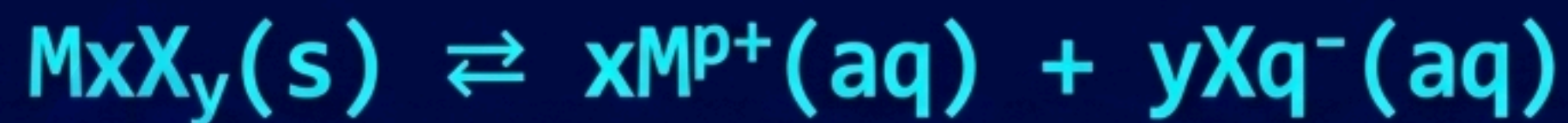
CORRECT MINDSET

"Calculate and compare Q_{sp} with K_{sp} for each salt."

Always calculate the exact concentration of the common ion needed to trigger precipitation to find the true answer.

Exam Snapshot: Formula Bank

General Salt Dissociation:



Solubility Product Equation:

$$K_{sp} = [M^{p+}]^x [X^{q-}]^y$$

Relationship with Molar Solubility (S):

$$K_{sp} = (xS)^x (yS)^y = x^x \cdot y^y \cdot S^{(x+y)}$$

[] = Concentration in mol/L

S = Molar Solubility

Pure Solids = Constant (Excluded)



You've Learned the Theory.

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