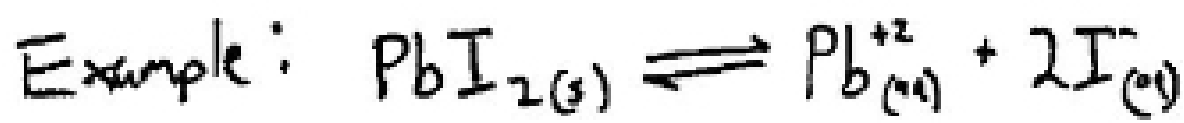


Everything I Need To Know About:

Solubility & Complex-Ion Equilibria

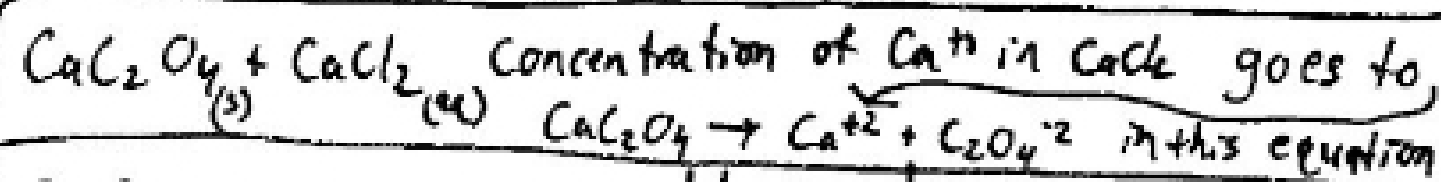
The Solubility Product Constant - K_{sp} - the equilibrium constant for the solubility equilibrium of a slightly-soluble (nearly insoluble) ionic compound



$K_{sp} = [Pb^{2+}][I^{-}]^2$ * only the products are included in K_{sp} because equilibrium expressions exclude pure liquids/solids and the reactants is a solid.

The same rules for equilibrium ICE tables apply to solubility problems

When a solution contains two of the same cations ~~and~~ from two different salts, the solid salt becomes less soluble



When solving ICE table problems for common ions, the cation concentration (instead of being 0 initially) is the concentration of the common ion.

Will Precipitation occur?

$Q_c = \frac{\text{products}}{\text{reactants}}$

technically, since reactants are solids, they = 1 and $Q_c = [\text{product}][\text{product}]$ same for K_c

If $Q_c < K_c \rightarrow$ forward direction, no precipitate

If $Q_c > K_c \rightarrow$ reverse direction, precipitate

If $Q_c = K_c \rightarrow$ equilibrium, no precipitate

if Q_c is smaller than K_c , more solid than solution, forward direction

if Q_c is larger than K_c , more solution than solid, precipitate

Fractional Precipitation - separating 2 or more ions from a solution by adding a reactant that precipitates first one ion then the next by manipulating K_{sp} .

Example: K_2CrO_4 is added to $.1M Ba^{2+}$ & $.1M Sr^{2+}$

$[Ba^{2+}][CrO_4^{2-}] = K_{sp}$ for $BaCrO_4$
then

$[Sr^{2+}][CrO_4^{2-}] = K_{sp}$ for $SrCrO_4$

the smaller concentration value precipitates first

% of ^{one} ~~each~~ ion remaining when the ^{the 2nd one} others begins to precipitate:

$\frac{[Original\ ion]}{X} \times [Anion\ when\ 2nd\ begins\ to\ precipitate] = K_{sp}$ for $BaCrO_4$

$\frac{X}{original\ concentration} = \% \text{ remaining when 2nd cation begins to precipitate}$