Solubility Equilibria Problems Key

General approach to solving solubility problems:

I. Write salt dissolution reaction (use Equilibrium arrow!)
   $\Rightarrow$ break solid salt (reactant) into ions (product); include charges and coefficients for # of ions!
   $\Rightarrow$ don't include water as a reactant!

II. Write K_{sp} expression
   $\Rightarrow$ don't include solid salt in equilibrium constant!

III. Set up ICE.
   $\Rightarrow$ x = molar solubility
   $\Rightarrow$ Can use Molar mass to convert from molar solubility to gram solubility.

1) If a saturated solution prepared by dissolving CaF_2 in water has $[Ca^{2+}] = 3.3\times10^{-4}$ M, what is the value of K_{sp}?

Dissolution eq: $CaF_{2(s)} \leftrightarrows Ca^{2+}(aq) + 2F^{-}(aq)$ **breaks into 2 F^- ions not F_2!**

	$CaF_{2(s)} \leftrightarrows$	$Ca^{2+}(aq)$ +	$2F^{-}(aq)$
I		0	0
C		+x	+2x
E		x	2x

Initial [ ]'s = 0

$[F^-]$ is double the $[Ca^{2+}]$

$x = [Ca^{2+}] = 3.3\times10^{-4}$ M

$K_{sp} = [Ca^{2+}][F^-]^2$ $\qquad [F^-] = 2x = 6.6\times10^{-4}$ M

$K_{sp} = (3.3\times10^{-4}\text{ M})(6.6\times10^{-4})^2$ $\qquad$ **$K_{sp} = 1.4\times10^{-10}$**

Or $K_{sp} = x(2x)^2 = 4x^3$; **plug in $x = 3.3\times10^{-4}$ M** $\quad K_{sp} = 4(3.3\times10^{-4})^3 =$ **1.4×10^{-10}**

2) K_{sp} for $Al(OH)_3$ is 1.9×10^{-33}. Calculate the molar solubility.

$Al(OH)_{3\,(s)} \leftrightarrows Al^{3+}(aq) + 3OH^{-}(aq)$

	$Al(OH)_{3\,(s)} \leftrightarrows$	$Al^{3+}(aq)$ +	$3OH^{-}(aq)$
I		0	0
C		+x	+3x
E		x	3x

$K_{sp} = [Al^{3+}][OH^-]^3$

$1.9\times10^{-33} = x(3x)^3$

$1.9\times10^{-33} = 27x^4$

$$x = \left(\frac{1.9\times10^{-33}}{27}\right)^{1/4} = (7.04\text{x}10^{-35})^{1/4} = 2.9\times10^{-9}\text{ M}$$

solubility = 2.9×10^{-9} M

3) The solubility of $Ca(OH)_2$ is found to be 0.233 g/L. Calculate K_{sp}.

Dissolution eq: $Ca(OH)_{2(s)} \leftrightarrows Ca^{2+}(aq) + 2OH^{-}(aq)$

	$Ca(OH)_{2(s)} \leftrightarrows$	$Ca^{2+}(aq)$ +	$2OH^{-}(aq)$
I		0	0
C		+x	+2x
E		x	2x

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

$K_{sp} = x(2x)^2 = 4x^3$

Need mol/L of $Ca(OH)_2$

$$x = \frac{0.233g}{L}\times\frac{1\,mol}{74.10g} = 3.14\times10^{-3}\,M$$

$K_{sp} = 4(3.14\times10^{-3})^3 =$ **1.24×10^{-7}**

4) a) Calculate the molar solubility of SrF_2 in pure water ($K_{sp} = 4.3\times10^{-9}$).

Dissolution eq: $SrF_{2}(s) \leftrightarrows Sr^{2+}(aq) + 2F^{-}(aq)$ **breaks into 2 F^- ions not F_2!**

	$SrF_2(s)$ ⇆	$Sr^{2+}(aq)$ +	$2F^-(aq)$
I		0	0
C		+x	+2x
E		x	2x

$K_{sp} = [Sr^{2+}][F^-]^2$

$4.3\times10^{-9} = x(2x)^2 = 4x^3$

$x = (4.3\times10^{-9}/4)^{1/3} = (1.075\times10^{-9})^{1/3}$

$x = 1.0\times10^{-3}$ M = molar solubility of SrF_2

b) Calculate the molar solubility of SrF_2 in 0.010 M NaF.

Assume SrF_2 is added to a NaF solution so $[Sr^{2+}]$ is initially 0, but NaF provides $[F^-]$.

NaF is soluble salt: $NaF \rightarrow Na^+ + F^-$ F^- is common ion

1:1 ratio so $[NaF] = [F^-] = 0.010$ M

	$SrF_2(s)$ ⇆	$Sr^{2+}(aq)$ +	$2F^-(aq)$
I		0	0.010
C		+x	+2x
E		x	0.010 + 2x

F^- comes from NaF, so don't double $[F^-]$

$K_{sp} = [Sr^{2+}][F^-]^2$

$4.3\times10^{-9} = x(0.010 + 2x)^2$

Assume 2x is a lot smaller than 0.010 $\Rightarrow$ $4.3\times10^{-9} = x(0.010)^2$

$x = \frac{4.3\times10^{-9}}{(0.010)^2}$ **$x = 4.3\times10^{-5}$ M = molar solubility of SrF_2 with common ion**

$\Rightarrow$ Adding common ion $\downarrow$ solubility of SrF_2

5) **Will a precipitate form** when 0.150 L of 0.10 M lead (II) nitrate and 0.100 L of 0.20 M sodium chloride are mixed? For $PbCl_2$, $K_{sp} = 1.2\times10^{-5}$

DR precipitation reaction: $Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow$ **$PbCl_2(s)$** $+ 2NaNO_3(aq)$

Solubility reaction: $PbCl_2(s) \leftrightarrows Pb^{2+} + 2Cl^-$

Will a precipitate form – indicates we need to find Q for $PbCl_2$!

$Q = [Pb^{2+}][Cl^-]^2$

1:1 mole ratios so initial $[Pb^{2+}] = [Pb(NO_3)_2] = 0.10$ M; initial $[Cl^-] = [NaCl] = 0.20$ M

After combining solutions: $[Pb^{2+}] = \left(\frac{0.10\text{M } Pb^{2+} \times 0.150\text{L}}{0.250\text{L}}\right) = 0.060\text{ M } Pb^{2+}$

$[Cl^-] = \left(\frac{0.20\text{M} Cl^- \times 0.100\text{L}}{0.250\text{L}}\right) = 0.080\text{ M } Cl^-$

$Q = (0.060)(0.080)^2 =$ **3.8×10^{-4}**

$Q > K_{sp}$, so $PbCl_2$ does precipitate. Eq shifts $\leftarrow$