Answers

1. Write the equilibrium equation and solubility product K_{sp} for each salt. The first one is done.

a) $Ca(OH)_{2\,(s)} \rightleftarrows Ca^{2+} + 2OH^-$ $\qquad$ $Ksp = [Ca^{2+}][OH^-]^2$

b) $AgCl_{(s)} \rightleftarrows Ag^+ + Cl^-$ $\qquad$ $Ksp = [Ag^+][Cl^-]$

c) $Na_3PO_{4\,(s)} \rightleftarrows 3Na^+ + PO_4^{-3}$ $\qquad$ $Ksp = [Na^+]^3[PO_4^{-3}]$

d) $(NH_4)_3PO_{4\,(s)} \rightleftarrows 3NH_4^+ + PO_4^{-3}$ $\qquad$ $Ksp = [NH_4^+]^3[PO_4^{-3}]$

e) $Cu_2SO_{4\,(s)} \rightleftarrows 2Cu^{+1} + SO_4^{-2}$ $\qquad$ $Ksp = [Cu^{+1}]^2[SO_4^{-2}]$

f) $Al(CH_3COO)_{3\,(s)} \rightleftarrows Al^{+3} + 3CH_3COO^-$ $\qquad$ $Ksp = [Al^{+3}][CH_3COO^-]^3$

g) $Ca_3(PO_4)_{2\,(s)} \rightleftarrows 3Ca^{+2} + 2PO_4^{-3}$ $\qquad$ $Ksp = [Ca^{+2}]^3[PO_4^{-3}]^2$

2. (D) MnS

3. D AgBr

Solubility to K_{sp} $\qquad$ Determining Ksp from Solubilities

The K_{sp} is a measure of the solubility of an ionic salt. The larger the value of the K_{sp}, the greater is the solubility of the salt. You can only calculate a K_{sp} if the solution is saturated. Only saturated salt solutions are in equilibrium. You can calculate the K_{sp} from the solubility of a salt, since the solubility represents the concentration required to saturate a solution.

1. Calculate the K_{sp} for $CaCl_2$ if 200.0 g of $CaCl_2$ are required to saturate 100.0 mL of solution.

$$\text{Molarity} = \frac{200\,\cancel{g} \times \dfrac{1\text{ mole}}{111.1\,\cancel{g}}}{0.100\text{L}} = 18.001\text{ M } CaCl_2$$

$CaCl_2 \rightleftarrows Ca^{2+} + 2Cl^-$

18.0 M   18.0 M   36.0M

$18M\ CaCl_2 = 18M\ Ca^{2+}$

$18M\ Ca^{2+} \times \frac{2\ mol/L\ Cl^-}{\cancel{1\ mol/L\ Ca^{2+}}} = 36\ M\ Cl^-$

R	$CaCl_{2(s)} \leftrightarrow$	$Ca^{2+}_{(aq)}$ +	$2Cl^-_{(aq)}$
I		∅	∅
C	18M	+18.0M	+36M
E		18M	36M.

Ksp = $[Ca^{2+}][Cl^-]^2$

Ksp = $[18.0][36.0]^2$

Ksp = 2.33×10^4

2. Calculate the K_{sp} for $AlCl_3$ if 100.0 g is required to saturate 150.0 mL of a solution.

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

$ksp = (5M)\cdot(15M)^3$

Ksp = 1.679×10^4

$100.0g \times \frac{1\ mol}{133.34g} = 0.75\ moles\ AlCl_3$

$C = \frac{n}{V} = \frac{0.75\ mol}{0.150\ L} = 5\ mol/L\ AlCl_3$

$5M\ AlCl_3 = 5M\ Ca^{2+} \times \frac{3M\ Cl^-_{(aq)}}{1M\ Ca^{2+}} = 15M\ Cl^-_{(aq)}$

R	$AlCl_{3(s)} \leftrightarrow$	$Al^{3+}_{(aq)}$ +	$3Cl^-_{(aq)}$
I			
C	5M	+5M	+15M
E		5M	15M

3. The solubility of SrF_2 is 2.83×10^{-5} M. Calculate the K_{sp}.

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

$ksp = (2.83\times10^{-5}M)(5.66\times10^{-5}M)^2$

Ksp = 9.07×10^{-14}

R	$SrF_{2(s)} \leftrightarrow$	$Sr^{2+}_{(aq)}$ +	$2F^-_{(aq)}$
I		∅	∅
C	$2.83\times10^{-5}M$	$+2.83\times10^{-5}M$	$+5.66\times10^{-5}M$
E		$2.83\times10^{-5}M$	$5.66\times10^{-5}M$

$[SrF_2]$ $2.83\times10^{-5}M$ = $[Sr^{2+}]$ $2.83\times10^{-5}M$

$2.83\times10^{-5}M\ Sr^{2+} \times \frac{2\ mol\ F^-}{1\ mol\ Sr^{2+}} = 5.66\times10^{-5}M$

4. The solubility of $GaBr_3$ is 15.8 g per 100 mL. Calculate the K_{sp}.

$$\text{Molarity} = \frac{15.8\ \cancel{g} \times \frac{1\ mole}{309.4\ \cancel{g}}}{0.100\ L} = 0.51066\ M$$

$GaBr_3 \rightleftarrows Ga^{3+} + 3Br^-$

0.51066 M   0.51066 M   1.532 M

$0.51066\ M\ \cancel{Ga^{3+}} \times \frac{3M\ Br^-}{1\ \cancel{M\ Ga^{3+}}} = 1.532\ M\ Br^-$

Ksp = $[Ga^{3+}][Br^-]^3$

Ksp = $[0.51066][1.532]^3$

Ksp = 1.83

R	$GaBr_{3(s)}$	Ga^{3+} +	$3Br^-_{(aq)}$
I		∅	∅
C	0.51066 M	+0.51066 M	+1.532 M
E		0.51066 M	1.532 M

Determing Solubilities from Ksp Values

Ksp to Solubility

Calculate the solubility in M and g/L for each. Use the K_{sp} values found in your chart.

1) $BaCO_3$

$$BaCO_{3(s)} \rightleftarrows Ba^{2+} + CO_3^{2-}$$

x x x

ksp = $[Ba^{2+}][CO_3^{2-}]$

ksp = x^2

$1.6 \times 10^{-9} = x^2$

4.0×10^{-5} M

$$\frac{4.0 \times 10^{-5}\ \text{mole}}{\text{L}} \times \frac{197.3\ \text{g}}{1\ \text{mole}} =$$

7.89×10^{-3} g/L

2) $Fe(OH)_2$

2.1×10^{-4} g/L

3) $PbCl_2$

4.0 g/L

4) How many grams of $Mg(OH)_2$ are required to completely saturate 1.5 L of solution?

$Mg(OH)_2$	$\rightleftarrows$	Mg^{2+} +	$2OH^-$
x		x	2x

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

$$Ksp = [x][2x]^2$$

$$Ksp = 4x^3$$

$$5.6 \times 10^{-12} = 4x^3$$

$$1.119 \times 10^{-4}\ M = x$$

$$1.5\ L \times \frac{1.119 \times 10^{-4}\ mole}{1\ L} \times \frac{58.3\ g}{1\ mole} = 9.8 \times 10^{-3}\ g$$

5) $AgCl(s) \rightleftarrows Ag^+(aq) + Cl^-(aq)$

$Ksp = [Ag^+][Cl^-]$

$1.77 \times 10^{-3} = x^2$

$1.33 \times 10^{-5}\ M = x = [Ag^+]$

$Ag_2CO_3(s) \rightleftarrows 2Ag^+(aq) + CO_3^{2-}(aq)$

$Ksp = [Ag^+]^2 [CO_3^{2-}]$

$8.46 \times 10^{-12} = 4x^3$

$1.28 \times 10^{-4}\ M = x$

$2.56 \times 10^{-4}\ M = 2x = [Ag^+]$

There are more Ag^+ ions created in the Ag_2CO_3 solution. In this solution the solubility of the Ag^+ ions is higher.

$ksp = [2x]^2[x]$

$ksp = [x][x]$

R	$AgCl_{(s)}$ ⇔	$Ag^+_{(aq)}$ +	$Cl^-_{(aq)}$
I		0	0
C		+ x	+ x
E		x	x

R	$Ag_2CO_{3(s)}$ ⇔	$2Ag^+_{(aq)}$ +	$CO^{2-}_{3(aq)}$
I		0	0
C		+2x	x
E		2x	x