

## Answers

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



→ 2. (D) MnS

→ 3. D AgBr

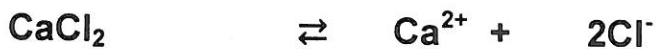
Solubility to  $K_{sp}$

## Determining $K_{sp}$ 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  $\text{CaCl}_2$  if 200.0 g of  $\text{CaCl}_2$  are required to saturate 100.0 mL of solution.

$$\text{Molarity} = \frac{200 \text{ g} \times \frac{1 \text{ mole}}{111.1 \text{ g}}}{0.100 \text{ L}} = 18.001 \text{ M } \text{CaCl}_2$$



18.0 M                      18.0 M      36.0 M

$$K_{sp} = [\text{Ca}^{2+}][\text{Cl}^-]^2$$

$$K_{sp} = [18.0][36.0]^2$$

$$K_{sp} = 2.33 \times 10^4$$

$$18\text{M CaCl}_2 = 18\text{M Ca}^{2+}$$

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

R	$\text{CaCl}_2(s)$	$\text{Ca}^{2+}_{(aq)}$	$2\text{Cl}^-_{(aq)}$
I		∅	∅
C	18M	+18.0M	+36M
E		18M	36M

2. Calculate the  $K_{sp}$  for  $\text{AlCl}_3$  if 100.0 g is required to saturate 150.0 mL of a solution.

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

$$K_{sp} = (5\text{M}) \cdot (15\text{M})^3$$

$$K_{sp} = 1.679 \times 10^4$$

$$100.0\text{g} \times \frac{1\text{ mol}}{133.34\text{g}} = 0.75\text{ moles AlCl}_3$$

$$C = \frac{n}{V} = \frac{0.75\text{ mol}}{0.150\text{ L}} = 5\text{ mol/L AlCl}_3$$

$$5\text{M AlCl}_3 = 5\text{M Ca}^{2+} \times \frac{3\text{M Cl}^-_{(aq)}}{1\text{M Ca}^{2+}} = 15\text{M Cl}^-_{(aq)}$$

R	$\text{AlCl}_3(s)$	$\text{Al}^{3+}$	$3\text{Cl}^-$
I		∅	∅
C	5M	+5M	+15M
E		5M	15M

3. The solubility of  $\text{SrF}_2$  is  $2.83 \times 10^{-5}\text{ M}$ . Calculate the  $K_{sp}$ .

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

$$K_{sp} = (2.83 \times 10^{-5}\text{M})(5.66 \times 10^{-5}\text{M})^2$$

$$K_{sp} = 9.07 \times 10^{-14}$$

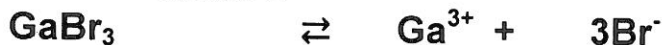
R	$\text{SrF}_2(s)$	$\text{Sr}^{2+}$	$2\text{F}^-$
I		∅	∅
C	$2.83 \times 10^{-5}\text{M}$	$+2.83 \times 10^{-5}\text{M}$	$+5.66 \times 10^{-5}\text{M}$
E		$2.83 \times 10^{-5}\text{M}$	$5.66 \times 10^{-5}\text{M}$

$$[\text{SrF}_2] = 2.83 \times 10^{-5}\text{M} = [\text{Sr}^{2+}]$$

$$2.83 \times 10^{-5}\text{M} \times \frac{2\text{ mol F}^-}{1\text{ mol Sr}^{2+}} = 5.66 \times 10^{-5}\text{M}$$

4. The solubility of  $\text{GaBr}_3$  is 15.8 g per 100 mL. Calculate the  $K_{sp}$ .

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



0.51066 M                      0.51066 M      1.532 M

$$K_{sp} = [\text{Ga}^{3+}][\text{Br}^-]^3$$

$$K_{sp} = [0.51066][1.532]^3$$

$$K_{sp} = 1.83$$

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

R	$\text{GaBr}_3(s)$	$\text{Ga}^{3+}$	$3\text{Br}^-$
I		∅	∅
C	0.51066 M	+0.51066 M	+1.532 M
E		0.51066 M	1.532 M

# Determining Solubilities from K<sub>sp</sub> Values

## K<sub>sp</sub> to Solubility

Calculate the solubility in M and g/L for each. Use the K<sub>sp</sub> values found in your chart.

1) BaCO<sub>3</sub>



$$K_{sp} = [\text{Ba}^{2+}][\text{CO}_3^{2-}]$$

$$K_{sp} = x^2$$

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

$$4.0 \times 10^{-5} \text{ M}$$

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

$$7.89 \times 10^{-3} \text{ g/L}$$

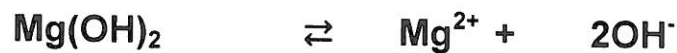
2) Fe(OH)<sub>2</sub>

$$2.1 \times 10^{-4} \text{ g/L}$$

3) PbCl<sub>2</sub>

$$4.0 \text{ g/L}$$

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



$$x \qquad \qquad \qquad x \qquad \qquad 2x$$

$$K_{sp} = [\text{Mg}^{2+}][\text{OH}^-]^2$$

$$K_{sp} = [x][2x]^2$$

$$K_{sp} = 4x^3$$

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

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

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



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

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

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



$$K_{sp} = [\text{Ag}^+]^2 [\text{CO}_3^{2-}]$$

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

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

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

There are more  $\text{Ag}^+$  ions created in the  $\text{Ag}_2\text{CO}_3$  solution. In this solution the solubility of the  $\text{Ag}^+$  ions is higher.

$$k_{sp} = [x][x]$$

$$k_{sp} = [2x]^2 [x]$$

R	$\text{AgCl}(\text{s})$	$\rightleftharpoons$	$\text{Ag}^+(\text{aq})$	$+$	$\text{Cl}^-(\text{aq})$
I	<del>/</del>		0		0
C	<del>/</del>		+x		+x
E	<del>/</del>		x		x

R	$\text{Ag}_2\text{CO}_3(\text{s})$	$\rightleftharpoons$	$2\text{Ag}^+(\text{aq})$	$+$	$\text{CO}_3^{2-}(\text{aq})$
I	<del>/</del>		0		0
C	<del>/</del>		+2x		x
E	<del>/</del>		2x		x