Chemistry 12
Tutorial 10—Ksp Calculations
Solutions

1. Calculate the solubility of SrF_2 in moles/Litre in water.

Equilibrium Equation: $\underset{-s}{SrF_{2(s)}} \rightleftarrows \underset{+s}{Sr^{2+}_{(aq)}} + \underset{+2s}{2F^{-}_{(aq)}}$ (where s = molar solubility)

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

$Ksp = s\ (2s)^2$

$Ksp = s \times 4s^2$

$Ksp = 4s^3$

$s^3 = \frac{Ksp}{4}$

$s = \sqrt[3]{\frac{Ksp}{4}} = \sqrt[3]{\frac{4.3 \times 10^{-9}}{4}} = 1.0244 \times 10^{-3}$ M

Because the Ksp was 2 SD's, the answer will be: Molar Solubility = 1.0×10^{-3} M

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2. Calculate the solubility of Ag_2CO_3 in grams/Litre.

Equilibrium Equation: $\underset{-s}{Ag_2CO_{3(s)}} \rightleftarrows \underset{+2s}{2Ag^{+}_{(aq)}} + \underset{+s}{CO_3^{2-}{}_{(aq)}}$ (where s = molar solubility)

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

$Ksp = (2s)^2 \times s$

$Ksp = 4s^2 \times s$

$Ksp = 4s^3$

$s^3 = \frac{Ksp}{4}$

$s = \sqrt[3]{\frac{Ksp}{4}} = \sqrt[3]{\frac{8.5 \times 10^{-12}}{4}} = 1.286 \times 10^{-4}$ M

Molar Solubility = 1.286×10^{-4} M

To get solubility in g/L:

$$\frac{1.286 \times 10^{-4}\ \text{mol}}{\text{L}} \times \frac{275.8\ \text{g}\ Ag_2CO_3}{1\ \text{mol}\ Ag_2CO_3} = \mathbf{3.5 \times 10^{-2}\ g/L\ \ (or\ \ 0.035\ g/L)}$$

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3. At a certain temperature, the *solubility* of $SrCO_3$ is 7.5×10^{-5} M. Calculate the *Ksp* for $SrCO_3$.

Molar Solubility

$$\begin{array}{ccccc} -7.5 \times 10^{-5}\ M & & +7.5 \times 10^{-5}\ M & & +7.5 \times 10^{-5}\ M \\ SrCO_{3(s)} & \rightleftarrows & Sr^{2+}_{(aq)} & + & CO_3^{2-}{}_{(aq)} \end{array}$$

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

$= (7.5 \times 10^{-5})^2$

$Ksp = 5.6 \times 10^{-9}$

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Answers to Self Test on Tutorial 10

1. The Ksp is just a K_{eq} for an ionic substance dissolving in water.

2. Give the ***Net-Ionic Equation*** and the ***Ksp expression*** for each of the following dissolving in water. (*The first one is done as an example*.)

Substance	Net-Ionic Equation	Ksp Expression
$Ag_2SO_{4(s)}$	$Ag_2SO_{4(s)} \rightleftarrows 2Ag^+_{(aq)} + SO_4^{2-}{}_{(aq)}$	$Ksp = [Ag^+]^2 [SO_4^{2-}]$
$CaCO_{3(s)}$	$CaCO_{3(s)} \rightleftarrows Ca^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)}$	$Ksp = [Ca^{2+}] [CO_3^{2-}]$
$Ca_3(PO_4)_{2(s)}$	$Ca_3(PO_4)_{2(s)} \rightleftarrows 3Ca^{2+}_{(aq)} + 2PO_4^{3-}{}_{(aq)}$	$Ksp = [Ca^{2+}]^3 [PO_4^{3-}]^2$
$AgClO_{3(s)}$	$AgClO_{3(s)} \rightleftarrows Ag^+_{(aq)} + ClO_3^-{}_{(aq)}$	$Ksp = [Ag^+] [ClO_3^-]$

3. a) Calculate the ***molar solubility*** (solubility in moles/Litre) of $Fe(OH)_2$ in water.

$$\underset{-s}{Fe(OH)_{2(s)}} \rightleftarrows \underset{+s}{Fe^{2+}_{(aq)}} + \underset{+2s}{2OH^-_{(aq)}}$$ (where "s" is the molar solubility)

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

$Ksp = s\,(2s)^2$

$Ksp = 4s^3$

$$s = \sqrt[3]{\frac{Ksp}{4}} = \sqrt[3]{\frac{4.9 \times 10^{-17}}{4}} = \mathbf{2.3 \times 10^{-6}\ M}$$

b) What is the $[OH^-]$ in a saturated solution of $Fe(OH)_2$?

$$\underset{-s}{Fe(OH)_{2(s)}} \rightleftarrows \underset{+s}{Fe^{2+}_{(aq)}} + \underset{+2s}{2OH^-_{(aq)}}$$ (where "s" is the molar solubility)

Since $[OH^-] = 2s$; then $[OH^-] = 2(2.3052 \times 10^{-6}) =$ $\mathbf{4.6 \times 10^{-6}\ M}$

4. Calculate the ***solubility*** of $BaCO_3$ in grams per Litre.

$$\underset{-s}{BaCO_{3(s)}} \rightleftarrows \underset{+s}{Ba^{2+}_{(aq)}} + \underset{+s}{CO_3^{2-}{}_{(aq)}}$$ where "s" is the Molar Solubility

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

$Ksp = s^2$

$$s = \sqrt{Ksp} = \sqrt{2.6 \times 10^{-9}} = 5.099 \times 10^{-5}\ M$$

<u>Solubility in g/L:</u>

5.099×10^{-5} mol/L x 197.3 g/mol = $\mathbf{1.0 \times 10^{-2}\ g/L}$

5. The solubility of PbI_2 at a certain temperature is 0.70 grams per Litre.

a) Calculate the solubility in moles/Litre

$$0.70\ \frac{g}{L} \times \frac{1\ mol}{461.0\ g} = 1.5 \times 10^{-3}\ mol/L$$ (unrounded = 1.518×10^{-3} M)

b) Calculate the ***value of Ksp*** for PbI_2 at this temperature

$$\underset{-1.518 \times 10^{-3}\ M}{PbI_{2(s)}} \rightleftarrows \underset{+1.518 \times 10^{-3}\ M}{Pb^{2+}_{(aq)}} + \underset{+3.036 \times 10^{-3}\ M}{2\ I^-_{(aq)}}$$

$Ksp = [Pb^{2+}] \times [I^-]^2$

$Ksp = 1.518 \times 10^{-3}\ (3.036 \times 10^{-3})^2 = \mathbf{1.4 \times 10^{-8}}$

6 It is found that 0.043 grams of $MgCO_3$ is all that can dissolve in 100.0 mL of solution at a certain temperature. From this information, calculate the ***Ksp*** for $MgCO_3$ at this temperature.

Plan: g → mol → M (molar solubility) → Ksp

$$0.043\ g\ MgCO_3 \times \frac{1\ mol\ MgCO_3}{84.3\ g\ MgCO_3} = 5.1008 \times 10^{-4}\ mol\ MgCO_3$$

$$[MgCO_3] = \frac{5.1008 \times 10^{-4}\ mol}{0.1000\ L} = 5.1008 \times 10^{-3}\ M$$ (*This is the molar solubility*)

$$\underset{-5.1008 \times 10^{-3}\ M}{MgCO_{3(s)}} \rightleftarrows \underset{+5.1008 \times 10^{-3}\ M}{Mg^{2+}_{(aq)}} + \underset{+5.1008 \times 10^{-3}\ M}{CO_3^{2-}{}_{(aq)}}$$

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

$= (5.1008 \times 10^{-3})^2$

Ksp = $\mathbf{2.6 \times 10^{-5}}$

7. Two separate experiments were done with combinations of Cu^{2+} and IO_3^- ions. Use the information given to fill in the missing value.

The Net-Ionic Equation for equilibrium is: $Cu(IO_3)_{2(s)} \rightleftarrows Cu^{2+}_{(aq)} + 2IO_{3\ (aq)}^{-}$

Experiment #	$[Cu^{2+}]$	$[IO_3^-]$
1	0.00327 M	0.00654 M
2	0.00240 M	?

Use the results from Experiment #1 to calculate the Ksp for $Cu(IO_3)_2$:

$$Ksp = [Cu^{2+}]\ [IO_3^-]^2$$

$$= 0.00327 \times (0.00654)^2$$

$$= 1.399 \times 10^{-7}$$

Use the Ksp and $[Cu^{2+}]$ to calculate $[IO_3^-]$ in Experiment #2:

$$Ksp = [Cu^{2+}]\ [IO_3^-]^2$$

$$[IO_3^-]^2 = \frac{Ksp}{[Cu^{2+}]}$$

$$[IO_3^-] = \sqrt{\frac{Ksp}{[Cu^{2+}]}} = \sqrt{\frac{1.399 \times 10^{-7}}{0.00240}}$$

$[IO_3^-] = 7.63 \times 10^{-3}$ M