Equilibrium
7.10 Reaction Quotient and Le Chatelier's Principle
7.11 Introduction to Solubility Equilibria
7.12 Common-Ion Effect
Worksheet Key

1) Which of the following is more soluble in water at 25°C?

$Ba(OH)_2$	$K_{sp} = 5 \times 10^{-3}$
$Ca(OH)_2$	$K_{sp} = 5.5 \times 10^{-6}$

$Ba(OH)_2$ is more soluble, as it has a larger K_{sp} value.

2) The solubility product constant, K_{sp}, for nickel (II) carbonate is 1.3×10^{-7} at 25°C.
   a. Write the balance chemical equation for this dissolving process.

$NiCO_3(s) \rightarrow Ni^{2+}(aq) + CO_3^{2-}(aq)$

   b. Write the equilibrium expression.

$K_{sp} = [Ni^{2+}][CO_3^{2-}]$

   c. Find the maximum molar concentration of all ions in solution and the molar solubility of the solution.

$$K_{sp} = [Ni^{2+}][CO_3^{2-}] = (s)(s) \qquad s = \text{molar solubility}$$
$$K_{sp} = s^2$$
$$1.3\times10^{-7} = s^2$$
$$s = \sqrt{1.3\times10^{-7}}$$
$$s = [Ni^{2+}] = [CO_3^{2-}] = 3.6\times10^{-4}\,M$$

3) The solubility product constant, K_{sp}, for calcium hydroxide is 6.5×10^{-6} at 25°C.
   a. Write the balance chemical equation for this dissociation in water.

$Ca(OH)_2(s) \rightarrow Ca^{2+}(aq) + 2\ OH^-(aq)$

   b. Write the equilibrium expression.

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

c. Find the molar concentration of all ions in this solution and the molar solubility of the solution.

$$K_{sp} = \left[Ca^{2+}\right]\left[OH^{-}\right]^2$$
$$K_{sp} = (s)(2s)^2 \qquad s = \text{molar solubility}$$
$$K_{sp} = 4s^3$$
$$6.5\times10^{-6} = 4s^3$$
$$s = \sqrt[3]{\frac{6.5\times10^{-6}}{4}} = [Ca^{2+}] = 1.2\times10^{-2}\,M;\ \left[OH^{-}\right] = 2s = 2(1.2\times10^{-2}\,M) = 2.4\times10^{-2}\,M$$

d. Calculate the maximum mass of calcium hydroxide that will dissolve in 100 mL of distilled water at 25°C.

$$100\text{ mL}\times\frac{1\text{ L}}{1000\text{ mL}}\times\frac{0.012\text{ mol Ca(OH)}_2}{1\text{ L}}\times\frac{74.10\text{ g Ca(OH)}_2}{1\text{ mol Ca(OH)}_2} = 8.9\times10^{-2}\text{g Ca(OH)}_2$$

4) A 250 mL saturated solution of silver chloride is prepared at 25°C. The solubility product constant, K_{sp}, for silver chloride is 1.8 x 10^{-10} at 25°C.

a. Write the balance chemical equation for this dissolving process.

$AgCl(s) \rightarrow Ag^{+}(aq) + Cl^{-}(aq)$

b. Write the equilibrium expression.

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

c. Find the molar concentrations of all ions in this solution and the molar solubility of the solution.

$$K_{sp} = \left[Ag^{+}\right]\left[Cl^{-}\right]$$
$$K_{sp} = (s)(s) \qquad s = \text{molar solubility}$$
$$K_{sp} = s^2$$
$$1.8\times10^{-10} = s^2$$
$$s = \sqrt{1.8\times10^{-10}}$$
$$s = [Ag^{+}] = [Cl^{-}] = 1.3\times10^{-5}\,M$$

d. What are the molar concentrations of all ions in a 500 mL saturated solution of silver chloride at 25°C?

$$[Ag^+] = [Cl^-] = 1.3 \times 10^{-5} M$$

Concentration in independent of volume. It is always measured in moles / 1 L.

5) What is minimum volume of distilled water that is needed to completely dissolve a 3.0 g sample of $CuCO_3$ at 25°C? (K_{sp} for $CuCO_3$ is 2.5 x 10^{-10} at 25°C.)

$$K_{sp} = 2.5 \times 10^{-10} = s^2,\ s = 1.6 \times 10^{-5} M$$

$$3.0 \text{ g } CuCO_3 \times \frac{1 \text{ mol } CuCO_3}{123.56 \text{ g } CuCO_3} = 0.024 \text{ mol } CuCO_3$$

$$\frac{0.024 \text{ mol } CuCO_3}{V} = 1.6 \times 10^{-5} M$$

$$V = 1500 \text{ L}$$

6) At a certain temperature, 9.5 x 10^{-5} g of zinc carbonate will dissolve in 150 mL of water.

a. Write the balanced chemical equation for the dissociation of zinc carbonate in water.

$$ZnCO_3(s) \rightarrow Zn^{2+}(aq) + CO_3^{2-}(aq)$$

b. Calculate the molar solubility of zinc carbonate in water at this temperature.

$$9.5 \times 10^{-5} \text{ g } ZnCO_3 \times \frac{1 \text{ mol } ZnCO_3}{125.40 \text{ g } ZnCO_3} = 7.6 \times 10^{-7} \text{ mol } ZnCO_3$$

$$s = \frac{7.6 \times 10^{-7} \text{ mol } ZnCO_3}{0.15 \text{ L}} = 5.1 \times 10^{-6} M$$

c. Calculate the value of the solubility product constant, K_{sp}, for zinc carbonate at this temperature.

$$K_{sp} = [Zn^{2+}][CO_3^{2-}]$$

$$K_{sp} = (s)(s) = s^2 \qquad s = \text{molar solubility}$$

$$K_{sp} = (5.1 \times 10^{-6})^2 = 2.6 \times 10^{-11}$$

7) A 1.0 L saturated solution of zinc hydroxide is prepared, and the concentration of Zn^{2+} is measured to be 4.22 x 10^{-6} *M* at 25°C.

a. Write the balanced chemical equation for the dissociation of zinc hydroxide in water.

$$Zn(OH)_2(s) \rightarrow Zn^{2+}(aq) + 2\ OH^-(aq)$$

b. Calculate the molar concentration of OH^- in solution.

$$[OH^-] = 2[Zn^{2+}] = 2\ (4.22 \text{ x } 10^{-6}\ M) = 8.44 \text{ x } 10^{-6}\ M$$

c. Calculate the value of the solubility product constant, K_{sp}.

$$K_{sp} = [Zn^{2+}][OH^-]^2 = (4.22 \text{ x } 10^{-6})(8.44 \text{ x } 10^{-6}\ )^2 = 3.01\times10^{-16} M$$

d. Calculate the maximum mass of zinc hydroxide that will dissolve in 150 mL of distilled water at 25°C

$$150 \text{ mL}\times\frac{1 \text{ L}}{1000 \text{ mL}}\times\frac{4.22\times10^{-6} \text{mol Zn(OH)}_2}{1 \text{ L}}\times\frac{99.41 \text{ g Zn(OH)}_2}{1 \text{ mol Zn(OH)}_2} = 6.3\times10^{-5} \text{g Zn(OH)}_2$$

e. What is the molar concentration of Zn^{2+} if 200 mL of water evaporates from the solution?

$[Zn^{2+}]$ = 4.22 x 10^{-6} *M*; concentration is not affected by changes in volume.

8) The solubility product constant, K_{sp}, for barium sulfate is 1.1 x 10^{-10} at 25°C. Will a precipitate of $BaSO_4(s)$ form when 210 mL of 4.75 x 10^{-2} *M* $Ba(NO_3)_2$ is mixed with 315 mL of 0.450 *M* Li_2SO_4.

Find $[Ba^{2+}]$ $\quad 0.21 \text{ L solution}\times\frac{4.75\times10^{-2} \text{mol Ba}^{2+}}{1 \text{ L solution}} = 1.0\times10^{-2} \text{mol Ba}^{2+}$

$$[Ba^{2+}] = \frac{1.0\times10^{-2} \text{mol Ba}^{2+}}{0.21 \text{ L} + 0.315 \text{ L}} = 2.0\times10^{-2} M$$

Find $[SO_4^{2-}]$ $\quad 0.315 \text{ L solution}\times\frac{0.450 \text{ mol SO}_4^{2-}}{1 \text{ L solution}} = 0.142 \text{ mol SO}_4^{2-}$

$$[SO_4^{2-}] = \frac{0.142 \text{ mol SO}_4^{2-}}{0.21 \text{ L} + 0.315 \text{ L}} = 0.27M$$

$$Q = [Ba^{2+}][SO_4^{2-}] = (2.0\times10^{-2})(0.27) = 5.4\times10^{-3}$$

$Q > K_{sp}$, so a precipitate will form

9) Find the solubility product constant for lead (II) iodide if the concentration of I^- is found to be 1.25 x 10^{-3} *M* when the solution is saturated.

$$PbI_{2(s)} \rightarrow Pb^{2+}{}_{(aq)} + 2\ I^-{}_{(aq)}$$

$$K_{sp} = [Pb^{2+}][I^-]^2 = (0.5 \times [I^-])[I^-]^2 = (0.5 \times 1.25 \times 10^{-3})(1.25 \times 10^{-3})^2 = 9.77 \times 10^{-10} M$$

10) Use the K_{sp} value obtained in the previous question to determine if a precipitate will form when 350 mL of 5.5 x 10^{-2} *M* lead (II) nitrate is mixed with 250 mL of 4.8 x 10^{-2} *M* sodium iodide.

Find $[Pb^{2+}]$

$$0.35\text{ L solution} \times \frac{5.5 \times 10^{-2}\text{ mol } Pb^{2+}}{1\text{ L solution}} = 1.9 \times 10^{-2}\text{ mol } Pb^{2+}$$

$$[Pb^{2+}] = \frac{1.9 \times 10^{-2}\text{ mol } Pb^{2+}}{0.35\text{ L} + 0.25\text{ L}} = 3.2 \times 10^{-2} M$$

Find $[I^-]$

$$0.25\text{ L solution} \times \frac{4.8 \times 10^{-2}\text{ mol } I^-}{1\text{ L solution}} = 1.2 \times 10^{-2}\text{ mol } I^-$$

$$[I^-] = \frac{1.2 \times 10^{-2}\text{ mol } I^-}{0.35\text{ L} + 0.25\text{ L}} = 0.020 M$$

$$Q = [Pb^{2+}][I^-]^2 = (3.2 \times 10^{-2})(0.020)^2 = 1.3 \times 10^{-5} M$$

$Q > K_{sp}$, so a precipitate will form

11) Find the solubility product constant for copper (II) carbonate if the concentration of Cu^{2+} is found to be 1.5 x 10^{-6} *M* when the solution is saturated.

$$CuCO_{3(s)} \rightarrow Cu^{2+}{}_{(aq)} + CO_3{}^{2-}{}_{(aq)}$$

$$K_{sp} = [Cu^{2+}][CO_3{}^{2-}] = (1.5\text{ x }10^{-6})(1.5\text{ x }10^{-6}) = 2.3 \times 10^{-12} M$$

12) Use the K_{sp} value obtained in the previous question to determine if a precipitate will form when 325 mL of 0.50 x 10^{-6} *M* copper (II) nitrate is mixed with 325 mL of 0.50 x 10^{-6} *M* potassium carbonate.

Find $[Cu^{2+}]$ $$0.325\text{ L solution}\times\frac{0.50\times10^{-6}\text{ mol Cu}^{2+}}{1\text{ L solution}}=1.6\times10^{-7}\text{ mol Cu}^{2+}$$

$$[Cu^{2+}]=\frac{1.6\times10^{-7}\text{ mol Pb}^{2+}}{0.325\text{ L}+0.325\text{ L}}=2.5\times10^{-7}\,M$$

Find $[CO_3^{2-}]$ $$0.325\text{ L solution}\times\frac{0.50\times10^{-6}\text{ mol CO}_3^{2-}}{1\text{ L solution}}=1.6\times10^{-7}\text{ mol CO}_3^{2-}$$

$$[I^-]=\frac{1.6\times10^{-7}\text{ mol CO}_3^{2-}}{0.325\text{ L}+0.325\text{ L}}=2.5\times10^{-7}\,M$$

$Q=[Cu^{2+}][CO_3^{2-}]=(2.5\times10^{-7})(2.5\times10^{-7})=6.3\times10^{-14}\,M$

$Q<K_{sp}$, no precipitate will form

13) If a 0.50 *M* solution of K_2SO_4 is slowly poured into a beaker containing 0.25 *M* barium nitrate and 0.30 M lead (II) nitrate at 25°C, what will be the first precipitate that forms? K_{sp} for barium sulfate is 1.1 x 10^{-10} and K_{sp} for lead (II) sulfate is 1.6 x 10^{-8}.

$BaSO_{4(s)}\rightarrow Ba^{2+}{}_{(aq)}+SO_4{}^{2-}{}_{(aq)}$ $K_{sp}=[Ba^{2+}][SO_4^{2-}]$

$$[SO_4^{2-}]=\frac{K_{sp}}{[Ba^{2+}]}=\frac{1.1\times10^{-10}}{0.25M}=4.4\times10^{-10}\,M$$

$PbSO_{4(s)}\rightarrow Pb^{2+}{}_{(aq)}+SO_4{}^{2-}{}_{(aq)}$ $K_{sp}=[Pb^{2+}][SO_4^{2-}]$

$$[SO_4^{2-}]=\frac{K_{sp}}{[Pb^{2+}]}=\frac{1.6\times10^{-8}}{0.30M}=5.3\times10^{-8}\,M$$

$BaSO_{4(s)}$ will be the first precipitate that forms, as $[SO_4^{2-}]$ will reach $4.4\times10^{-10}\,M$ before it reaches $5.3\times10^{-8}\,M$.

14) If a 0.50 *M* solution of KOH is slowly poured into a beaker containing 0.35 *M* magnesium nitrate and 0.032 *M* calcium nitrate at 25°C, what will be the first precipitate that forms? K_{sp} for calcium hydroxide is 6.5 x 10^{-6} and K_{sp} for magnesium hydroxide is 6.3 x 10^{-10}.

$$Mg(OH)_{2(s)} \rightarrow Mg^{2+}{}_{(aq)} + 2\ OH^{-}{}_{(aq)} \qquad K_{sp} = [Mg^{2+}][OH^{-}]^2$$

$$[OH^{-}] = \sqrt{\frac{K_{sp}}{[Mg^{2+}]}} = \frac{6.3\times10^{-10}}{0.35M} = 4.2\times10^{-5}M$$

$$Ca(OH)_{2(s)} \rightarrow Ca^{2+}{}_{(aq)} + 2\ OH^{-}{}_{(aq)} \qquad K_{sp} = [Ca^{2+}][OH^{-}]^2$$

$$[OH^{-}] = \sqrt{\frac{K_{sp}}{[Ca^{2+}]}} = \frac{6.5\times10^{-6}}{0.032M} = 1.4\times10^{-2}M$$

$Mg(OH)_{2(s)}$ will be the first precipitate that forms, as $[OH^{-}]$ will reach $4.2\times10^{-5}M$ before it reaches $1.4\times10^{-2}M$.

15) Explain why the solubility of AgBr decreases when NaBr is added to the system.

When additional Br^{-} is added to the system, the equilibrium in the following reaction ($AgBr(s) \rightleftharpoons Ag^{+}(aq) + Br^{-}(aq)$) shifts to the left to relieve the stress (Le Chatelier's principle). LO 6.23

16) The solubility product constant, K_{sp}, for Ag_2SO_4 is 1.2 x 10^{-5} at 25°C. Find $[Ag^{+}]$ and $[SO_4{}^{2-}]$ after 0.755 g of $AgNO_3$ are added to a 500.0 mL saturated solution of Ag_2SO_4 and equilibrium is established. Assume that the total volume of the solution remains the same and the final temperate is 25°C.

Step 1) Find moles of Ag^{+} in original solution

$$K_{sp} = [Ag^{+}]^2[SO_4{}^{2-}]$$

$$K_{sp} = 4s^2s = 4s^3$$

$$s = \sqrt[3]{\frac{1.2\times10^{-5}}{4}}$$

$$s = [SO_4{}^{2-}] = 1.4\times10^{-2}M$$

$$[Ag^{+}] = 2s = 2(1.4\times10^{-2}M) = 2.8\times10^{-2}M$$

$$n_{Ag^+} = 0.5000\ L\times\frac{2.8\times10^{-2}\ mol\ Ag^{+}}{1\ L}$$

$$n_{Ag^+} = 1.4\times10^{-2}\ mol\ Ag^{+}$$

Step 2) Find moles of Ag^{+} that were added

$$0.755\ g\ AgNO_3\times\frac{1\ mol\ AgNO_3}{169.88\ g\ AgNO_3}\times\frac{1\ mol\ Ag^{+}}{1\ mol\ AgNO_3} = 4.44\times10^{-3} mol\ Ag^{+}$$

Step 3) Find $[Ag^+]$ in the final solution

$$[Ag^+] = \frac{1.4\times10^{-2}\text{ mol Ag}^+ + 0.444\times10^{-2}\text{ mol Ag}^+}{0.5000\text{ L}}$$

$$[Ag^+] = \frac{1.8\times10^{-2}\text{ mol Ag}^+}{0.5000\text{ L}}$$

$$[Ag^+] = 0.036M$$

Step 4) Find $[SO_4^{2-}]$ in the final solution

$$K_{sp} = [Ag^+]^2[SO_4^{2-}]$$

$$[SO_4^{2-}] = \frac{K_{sp}}{[Ag^+]^2}$$

$$[SO_4^{2-}] = \frac{1.2\times10^{-5}}{0.036^2}$$

$$[SO_4^{2-}] = 9.3\times10^{-3}M$$