CHAPTER EIGHT
Equilibrium Constant

8.1 Use $BaSO_4$ to distinguish between solubility, molar solubility, and solubility product.

8.2 Why do we usually not quote the K_{sp} values for soluble ionic compounds?

8.3 Write balanced equations and solubility product expressions for the solubility equilibria of the following compounds: (a) CuBr, (b) ZnC_2O_4, (c)Ag_2CrO_4, (d) Hg_2Cl_2, (e) $AuCl_3$, (f) $Mn_3(PO_4)_2$.

8.4 Write the solubility product expression for the ionic compound A_xB_y.

8.5 How can we predict whether a precipitate will form when two solutions are mixed?

8.6 Silver chloride has a larger K_{sp} than silver carbonate (see Table 16.2). Does this mean that AgCl also has a larger molar solubility than Ag_2CO_3?

8.7 Calculate the concentration of ions in the following saturated solutions: (a) $[I^-]$ in AgI solution with $[Ag^+] = 9.1 \times 10^{-9}$ *M*, (b) $[Al3^+]$ in $Al(OH)_3$ solution with $[OH^-] = 2.9 \times 10^{-9}$ *M*.

8.8 From the solubility data given, calculate the solubility products for the following compounds: (a) SrF_2, 7.3×10^{-2} g/L, (b) Ag_3PO_4, 6.7×10^{-3} g/L.

8.9 The molar solubility of $MnCO_3$ is 4.2×10^{-6} *M*. What is K_{sp} for this compound?

8.10 The solubility of an ionic compound MX (molar mass = 346 g) is 4.63×10^{-3} g/L. What is K_{sp} for the compound?

8.11 The solubility of an ionic compound M_2X_3 (molar mass = 288 g) is 3.6×10^{-17} g/L. What is K_{sp} for the compound?

8.12 Using data from Table 16.2, calculate the molar solubility of CaF_2.

8.13 What is the pH of a saturated zinc hydroxide solution?

8.14 The pH of a saturated solution of a metal hydroxide MOH is 9.68. Calculate the K_{sp} for the compound.

8.15 If 20.0 mL of 0.10 *M* $Ba(NO_3)_2$ are added to 50.0 mL of 0.10 *M* Na_2CO_3, will $BaCO_3$ precipitate?

8.16 A volume of 75 mL of 0.060 *M* NaF is mixed with 25 mL of 0.15 *M* $Sr(NO_3)_2$. Calculate the concentrations in the final solution of NO^-_3, Na^+, Sr^{2+}, and F^-. (K_{sp} for $SrF_2 = 2.0 \times 10^{-10}$.)

8.17 How does the common ion effect influence solubility equilibria? Use Le Châtelier's principle to explain the decrease in solubility of $CaCO_3$ in a Na_2CO_3 solution.

8.18 The molar solubility of AgCl in 6.5×10^{-3} *M* $AgNO_3$ is 2.5×10^{-8} *M*. In deriving K_{sp} from these data, which of the following assumptions are reasonable?
(a) K_{sp} is the same as solubility.
(b) K_{sp} of AgCl is the same in 6.5×10^{-3} *M* $AgNO_3$ as in pure water.
(c) Solubility of AgCl is independent of the concentration of $AgNO_3$.
(d) [Ag_] in solution does not change significantly upon the addition of AgCl to 6.5×10^{-3} *M* $AgNO_3$.
(e) $[Ag^+]$ in solution after the addition of AgCl to 6.5×10^{-3} *M* $AgNO_3$ is the same as it would be in pure water.

8.19 How many grams of $CaCO_3$ will dissolve in 3.0×10^2 mL of 0.050 *M* $Ca(NO_3)_2$?

8.20 The solubility product of $PbBr_2$ is 8.9×10^{-6}. Determine the molar solubility (a) in pure water, (b) in 0.20 *M* KBr solution, (c) in 0.20 *M* $Pb(NO_3)_2$ solution.

8.21 Calculate the molar solubility of AgCl in a solution made by dissolving 10.0 g of $CaCl_2$ in 1.00 L of solution.

8.22 Calculate the molar solubility of $BaSO_4$ (a) in water, (b) in a solution containing 1.0 *M* SO^{2-}_4 ions.

8.23 Explain the formation of complexes in Table 16.3 in terms of Lewis acid-base theory.

8.24 Give an example to illustrate the general effect of complex ion formation on solubility.

8.25 If 2.50 g of $CuSO_4$ are dissolved in 9.0 x 10^2 mLof 0.30 *M* NH_3, what are the concentrations of Cu^{2+}, $Cu(NH_3)^{2+}_4$, and NH_3 at equilibrium?

8.26 Calculate the concentrations of Cd^{2+}, $Cd(CN)^{2-}_4$, and CN^- at equilibrium when 0.50 g of $Cd(NO_3)_2$ dissolves in 5.0 x 10^2 mL of 0.50 *M* NaCN.

8.27 If NaOH is added to 0.010 *M* Al^{3+}, which will be the predominant species at equilibrium: $Al(OH)_3$ or $Al(OH)^-_4$? The pH of the solution is 14.00. [K_f for $Al(OH)^-_4$ = 2.0 x 10^{33}.]

8.28 Calculate the molar solubility of AgI in a 1.0 *M* NH_3 solution.

8.29 Both Ag^+ and Zn^{2+} form complex ions with NH_3. Write balanced equations for the reactions. However, $Zn(OH)_2$ is soluble in 6 *M* NaOH, and AgOH is not. Explain.

8.30 Explain, with balanced ionic equations, why (a) CuI_2 dissolves in ammonia solution, (b) AgBr dissolves in NaCN solution, (c) $HgCl_2$ dissolves in KCl solution.

8.31 A 200-mL volume of NaOH solution was added to 400 mL of a 2.00 *M* HNO_2 solution. The pH of the mixed solution was 1.50 units greater than that of the original acid solution. Calculate the molarity of the NaOH solution.

8.32 A solution is made by mixing exactly 500 mL of 0.167 *M* NaOH with exactly 500 mL 0.100 *M* CH_3COOH. Calculate the equilibrium concentrations of H^+, CH_3COOH, CH_3COO^-, OH^-, and Na^+.

8.33 For which of the following reactions is the equilibrium constant called a solubility product?
(a) $Zn(OH)_2(s) + 2OH^-(aq) \leftrightarrow Zn(OH)^{2-}_4(aq)$
(b) $3Ca^{2+}(aq) + 2PO^{3-}_4(aq) \leftrightarrow Ca_3(PO_4)_2(s)$
(c) $CaCO_3(s) + 2H^+(aq) \leftrightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$
(d) $PbI_2(s) \leftrightarrow Pb^{2+}(aq) + 2I^-(aq)$

8.34 Equal volumes of 0.12 *M*$AgNO_3$ and 0.14 *M* $ZnCl_2$ solution are mixed. Calculate the equilibrium concentrations of Ag^+, Cl^-, Zn^{2+}, and NO^-_3.

8.35 Calculate the solubility (in g/L) of Ag_2CO_3.

8.36 The molar solubility of $Pb(IO_3)_2$ in a 0.10 *M* $NaIO_3$ solution is 2.4 x 10^{-11} mol/L. What is K_{sp} for $Pb(IO_3)_2$?

8.37 Barium is a toxic substance that can seriously impair heart function. For an X ray of the gastrointestinal tract, a patient drinks an aqueous suspension of 20 g $BaSO_4$. If this substance were to equilibrate with the 5.0 L of the blood in the patient's body, what would be [Ba_2_]? For a good estimate, we may assume that the temperature is at 25°C. Why is $Ba(NO_3)_2$ not chosen for this procedure?

8.38 Acid-base reactions usually go to completion. Con- firm this statement by calculating the equilibrium constant for each of the following cases: (a) A strong acid reacting with a strong base. (b) A strong acid reacting with a weak base (NH_3). (c) A weak acid (CH_3COOH) reacting with a strong base. (d) A weak acid (CH_3COOH) reacting with a weak base (NH_3). (*Hint:* Strong acids exist as H^+ ions and strong bases exist as OH^- ions in solution. You need to look up K_a, K_b, and K_w.)

8.39 Look up the K_{sp} values for $BaSO_4$ and $SrSO_4$ in Table 16.2. Calculate the concentrations of Ba^{2+}, Sr^{2+}, and SO^{2-}_4 in a solution that is saturated with both compounds.

8.40 $CaSO_4$ (K_{sp} = 2.4 x 10^{-5}) has a larger K_{sp} value than that of Ag_2SO_4 (K_{sp} = 1.4 x 10^{-5}). Does it follow that $CaSO_4$ also has greater solubility (g/L)?

8.41 Water containing Ca^{2+} and Mg^{2+} ions is called *hard water* and is unsuitable for some household and industrial use because these ions react with soap to form insoluble salts, or curds. One way to remove the Ca^{2+} ions from hard water is by adding washing soda ($Na_2CO_3 \cdot 10H_2O$). (a) The molar solubility of $CaCO_3$ is 9.3 x 10^{-5} *M*. What is its molar solubility in a 0.050 *M* Na_2CO_3 solution? (b) Why are Mg^{2+} ions not removed by this procedure? (c) The Mg^{2+} ions are removed as $Mg(OH)_2$ by adding slaked lime [$Ca(OH)_2$] to the water to produce a saturated solution. Calculate the pH of a saturated $Ca(OH)_2$ solution. (d) What is the concentration of Mg^{2+} ions at this pH? (e) In general, which ion (Ca^{2+} or Mg^{2+}) would you remove first? Why?