

## CHAPTER EIGHT

### Equilibrium Constant

**8.1** Use  $\text{BaSO}_4$  to distinguish between solubility, molar solubility, and solubility product.

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

**8.3** Write balanced equations and solubility product expressions for the solubility equilibria of the following compounds: (a)  $\text{CuBr}$ , (b)  $\text{ZnC}_2\text{O}_4$ , (c)  $\text{Ag}_2\text{CrO}_4$ , (d)  $\text{Hg}_2\text{Cl}_2$ , (e)  $\text{AuCl}_3$ , (f)  $\text{Mn}_3(\text{PO}_4)_2$ .

**8.4** Write the solubility product expression for the ionic compound  $\text{A}_x\text{B}_y$ .

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

**8.6** Silver chloride has a larger  $K_{\text{sp}}$  than silver carbonate (see Table 16.2). Does this mean that  $\text{AgCl}$  also has a larger molar solubility than  $\text{Ag}_2\text{CO}_3$ ?

**8.7** Calculate the concentration of ions in the following saturated solutions: (a)  $[\text{I}^-]$  in  $\text{AgI}$  solution with  $[\text{Ag}^+] = 9.1 \times 10^{-9} \text{ M}$ , (b)  $[\text{Al}^{3+}]$  in  $\text{Al}(\text{OH})_3$  solution with  $[\text{OH}^-] = 2.9 \times 10^{-9} \text{ M}$ .

**8.8** From the solubility data given, calculate the solubility products for the following compounds: (a)  $\text{SrF}_2$ ,  $7.3 \times 10^{-2} \text{ g/L}$ , (b)  $\text{Ag}_3\text{PO}_4$ ,  $6.7 \times 10^{-3} \text{ g/L}$ .

**8.9** The molar solubility of  $\text{MnCO}_3$  is  $4.2 \times 10^{-6} \text{ M}$ . What is  $K_{\text{sp}}$  for this compound?

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

**8.11** The solubility of an ionic compound  $\text{M}_2\text{X}_3$  (molar mass = 288 g) is  $3.6 \times 10^{-17} \text{ g/L}$ . What is  $K_{\text{sp}}$  for the compound?

**8.12** Using data from Table 16.2, calculate the molar solubility of  $\text{CaF}_2$ .

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

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

**8.15** If 20.0 mL of 0.10  $\text{M}$   $\text{Ba}(\text{NO}_3)_2$  are added to 50.0 mL of 0.10  $\text{M}$   $\text{Na}_2\text{CO}_3$ , will  $\text{BaCO}_3$  precipitate?

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

**8.17** How does the common ion effect influence solubility equilibria? Use Le Châtelier's principle to explain the decrease in solubility of  $\text{CaCO}_3$  in a  $\text{Na}_2\text{CO}_3$  solution.

**8.18** The molar solubility of  $\text{AgCl}$  in  $6.5 \times 10^{-3} \text{ M}$   $\text{AgNO}_3$  is  $2.5 \times 10^{-8} \text{ M}$ . In deriving  $K_{\text{sp}}$  from these data, which of the following assumptions are reasonable?

(a)  $K_{\text{sp}}$  is the same as solubility.

(b)  $K_{\text{sp}}$  of  $\text{AgCl}$  is the same in  $6.5 \times 10^{-3} \text{ M}$   $\text{AgNO}_3$  as in pure water.

(c) Solubility of  $\text{AgCl}$  is independent of the concentration of  $\text{AgNO}_3$ .

(d)  $[\text{Ag}^+]$  in solution does not change significantly upon the addition of  $\text{AgCl}$  to  $6.5 \times 10^{-3} \text{ M}$   $\text{AgNO}_3$ .

(e)  $[\text{Ag}^+]$  in solution after the addition of  $\text{AgCl}$  to  $6.5 \times 10^{-3} \text{ M}$   $\text{AgNO}_3$  is the same as it would be in pure water.

**8.19** How many grams of  $\text{CaCO}_3$  will dissolve in  $3.0 \times 10^2 \text{ mL}$  of 0.050  $\text{M}$   $\text{Ca}(\text{NO}_3)_2$ ?

**8.20** The solubility product of  $\text{PbBr}_2$  is  $8.9 \times 10^{-6}$ . Determine the molar solubility (a) in pure water, (b) in 0.20  $\text{M}$   $\text{KBr}$  solution, (c) in 0.20  $\text{M}$   $\text{Pb}(\text{NO}_3)_2$  solution.

**8.21** Calculate the molar solubility of  $\text{AgCl}$  in a solution made by dissolving 10.0 g of  $\text{CaCl}_2$  in 1.00 L of solution.

**8.22** Calculate the molar solubility of  $\text{BaSO}_4$  (a) in water, (b) in a solution containing 1.0  $\text{M}$   $\text{SO}_4^{2-}$  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  $\text{CuSO}_4$  are dissolved in  $9.0 \times 10^2$  mL of  $0.30 \text{ M NH}_3$ , what are the concentrations of  $\text{Cu}^{2+}$ ,  $\text{Cu}(\text{NH}_3)_4^{2+}$ , and  $\text{NH}_3$  at equilibrium?

**8.26** Calculate the concentrations of  $\text{Cd}^{2+}$ ,  $\text{Cd}(\text{CN})_4^{2-}$ , and  $\text{CN}^-$  at equilibrium when 0.50 g of  $\text{Cd}(\text{NO}_3)_2$  dissolves in  $5.0 \times 10^2$  mL of  $0.50 \text{ M NaCN}$ .

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

**8.28** Calculate the molar solubility of  $\text{AgI}$  in a  $1.0 \text{ M NH}_3$  solution.

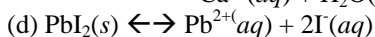
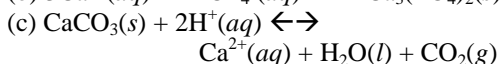
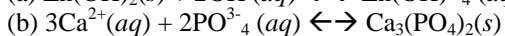
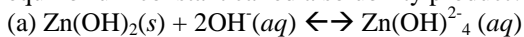
**8.29** Both  $\text{Ag}^+$  and  $\text{Zn}^{2+}$  form complex ions with  $\text{NH}_3$ . Write balanced equations for the reactions. However,  $\text{Zn}(\text{OH})_2$  is soluble in  $6 \text{ M NaOH}$ , and  $\text{AgOH}$  is not. Explain.

**8.30** Explain, with balanced ionic equations, why (a)  $\text{CuI}_2$  dissolves in ammonia solution, (b)  $\text{AgBr}$  dissolves in  $\text{NaCN}$  solution, (c)  $\text{HgCl}_2$  dissolves in  $\text{KCl}$  solution.

**8.31** A 200-mL volume of  $\text{NaOH}$  solution was added to 400 mL of a  $2.00 \text{ M HNO}_2$  solution. The pH of the mixed solution was 1.50 units greater than that of the original acid solution. Calculate the molarity of the  $\text{NaOH}$  solution.

**8.32** A solution is made by mixing exactly 500 mL of  $0.167 \text{ M NaOH}$  with exactly 500 mL  $0.100 \text{ M CH}_3\text{COOH}$ . Calculate the equilibrium concentrations of  $\text{H}^+$ ,  $\text{CH}_3\text{COOH}$ ,  $\text{CH}_3\text{COO}^-$ ,  $\text{OH}^-$ , and  $\text{Na}^+$ .

**8.33** For which of the following reactions is the equilibrium constant called a solubility product?



**8.34** Equal volumes of  $0.12 \text{ M AgNO}_3$  and  $0.14 \text{ M ZnCl}_2$  solution are mixed. Calculate the equilibrium concentrations of  $\text{Ag}^+$ ,  $\text{Cl}^-$ ,  $\text{Zn}^{2+}$ , and  $\text{NO}_3^-$ .

**8.35** Calculate the solubility (in g/L) of  $\text{Ag}_2\text{CO}_3$ .

**8.36** The molar solubility of  $\text{Pb}(\text{IO}_3)_2$  in a  $0.10 \text{ M NaIO}_3$  solution is  $2.4 \times 10^{-11} \text{ mol/L}$ . What is  $K_{\text{sp}}$  for  $\text{Pb}(\text{IO}_3)_2$ ?

**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  $\text{BaSO}_4$ . If this substance were to equilibrate with the 5.0 L of the blood in the patient's body, what would be  $[\text{Ba}^{2+}]$ ? For a good estimate, we may assume that the temperature is at  $25^\circ\text{C}$ . Why is  $\text{Ba}(\text{NO}_3)_2$  not chosen for this procedure?

**8.38** Acid-base reactions usually go to completion. Confirm 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 ( $\text{NH}_3$ ). (c) A weak acid ( $\text{CH}_3\text{COOH}$ ) reacting with a strong base. (d) A weak acid ( $\text{CH}_3\text{COOH}$ ) reacting with a weak base ( $\text{NH}_3$ ). (*Hint*: Strong acids exist as  $\text{H}^+$  ions and strong bases exist as  $\text{OH}^-$  ions in solution. You need to look up  $K_a$ ,  $K_b$ , and  $K_w$ .)

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

**8.40**  $\text{CaSO}_4$  ( $K_{\text{sp}} = 2.4 \times 10^{-5}$ ) has a larger  $K_{\text{sp}}$  value than that of  $\text{Ag}_2\text{SO}_4$  ( $K_{\text{sp}} = 1.4 \times 10^{-5}$ ). Does it follow that  $\text{CaSO}_4$  also has greater solubility (g/L)?

**8.41** Water containing  $\text{Ca}^{2+}$  and  $\text{Mg}^{2+}$  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  $\text{Ca}^{2+}$  ions from hard water is by adding washing soda ( $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ ). (a) The molar solubility of  $\text{CaCO}_3$  is  $9.3 \times 10^{-5} \text{ M}$ . What is its molar solubility in a  $0.050 \text{ M Na}_2\text{CO}_3$  solution? (b) Why are  $\text{Mg}^{2+}$  ions not removed by this procedure? (c) The  $\text{Mg}^{2+}$  ions are removed as  $\text{Mg}(\text{OH})_2$  by adding slaked lime [ $\text{Ca}(\text{OH})_2$ ] to the water to produce a saturated solution. Calculate the pH of a saturated  $\text{Ca}(\text{OH})_2$  solution. (d) What is the concentration of  $\text{Mg}^{2+}$  ions at this pH? (e) In general, which ion ( $\text{Ca}^{2+}$  or  $\text{Mg}^{2+}$ ) would you remove first? Why?