Figure 18-13

The briney taste of a pickle is a great accompaniment for a deli sandwich. But the main purpose of the salt is to preserve the pickle. Can you think of other foods that are preserved by salting?

Solubility Equilibria

Some ionic compounds dissolve readily in water and some barely dissolve at all. Sodium chloride, or table salt, is typical of the soluble ionic compounds. On dissolving, all ionic compounds dissociate into ions.

$$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$$

Approximately 36 g NaCl dissolves in 100 mL of water at 273 K. Without this high solubility, sodium chloride couldn't flavor and preserve foods like the pickles shown in **Figure 18-13.** Sodium chloride's vital role as an electrolyte in blood chemistry also depends upon its high solubility.

Although high solubility in water is often beneficial, low solubility also is important in many applications. For example, although barium ions are toxic to humans, patients are required to ingest barium sulfate prior to having an X ray of the digestive tract taken. X rays taken without barium sulfate in the digestive system are not well defined. Why can patients safely ingest barium sulfate?

In water solution, barium sulfate dissociates according to this equation.

$$BaSO_4(s) \rightarrow Ba^{2+}(aq) + SO_4^{2-}(aq)$$

As soon as the first product ions form, the reverse reaction begins to re-form the reactants according to this equation.

$$BaSO_4(s) \leftarrow Ba^{2+}(aq) + SO_4^{2-}(aq)$$

With time, the rate of the reverse reaction becomes equal to the rate of the forward reaction and equilibrium is established.

$$BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$$

For sparingly soluble compounds, such as $BaSO_4$, the rates become equal when the concentrations of the aqueous ions are exceedingly small. Nevertheless, the solution at equilibrium is a saturated solution.

The equilibrium constant expression for the dissolving of $BaSO_4$ is

$$K_{eq} = \frac{[Ba^{2+}][SO_4^{2-}]}{[BaSO_4]}$$

In the equilibrium expression, $[BaSO_4]$ is constant because barium sulfate is a solid. This constant value is combined with K_{eq} by multiplying both sides of the equation by $[BaSO_4]$.

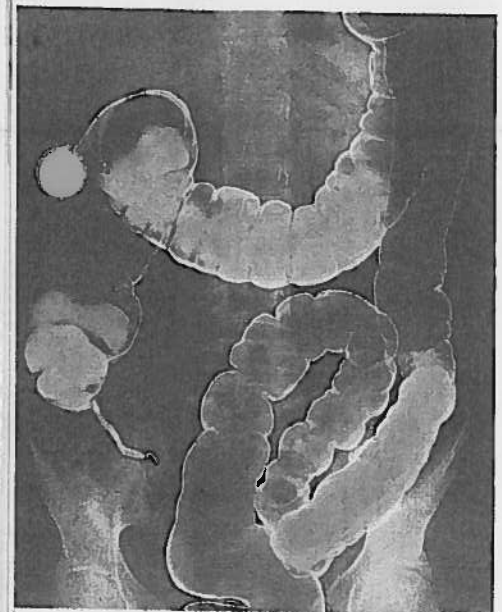


Figure 18-14

The presence of barium ions in the gastrointestinal system made the sharp definition of this X ray possible.

$$K_{eq} \times [BaSO_4] = [Ba^{2+}][SO_4^{2-}]$$

The product of K_{eq} and the concentration of the undissolved solid creates a new constant called the solubility product constant, K_{sp}. The **solubility product constant** is an equilibrium constant for the dissolving of a sparingly soluble ionic compound in water. The solubility product constant expression is

$$K_{sp} = [Ba^{2+}][SO_4^{2-}] = 1.1 \times 10^{-10} \text{ at } 298 \text{ K}$$

The solubility product constant expression is the product of the concentrations of the ions each raised to the power equal to the coefficient of the ion in the chemical equation. The small value of K_{sp} indicates that products are not favored at equilibrium. Thus, few barium ions are present at equilibrium ($1.0 \times 10^{-5}M$) and a patient can safely ingest a barium sulfate solution to obtain a clear X ray like the one shown in **Figure 18-14.**

Here is another example.

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$$
$$K_{sp} = [Mg^{2+}][OH^-]^2$$

K_{sp} depends only on the concentrations of the ions in the saturated solution. However, to establish an equilibrium system, some undissolved solid, no matter how small the amount, must be present in the equilibrium mixture.

The solubility product constants for some ionic compounds are listed in **Table 18-3.** Note that they are all small numbers. Solubility product constants are measured and recorded only for sparingly soluble compounds.

Using solubility product constants The solubility product constants in **Table 18-3** have been determined through careful experiments. K_{sp} values are important because they can be used to determine the solubility of a sparingly soluble compound. Recall that the solubility of a compound in water is the amount of the substance that will dissolve in a given volume of water.

Table 18-3

Solubility Product Constants at 298 K

Compound	K_{sp}	Compound	K_{sp}	Compound	K_{sp}
Carbonates		**Halides**		**Hydroxides**	
$BaCO_3$	2.6×10^{-9}	CaF_2	3.5×10^{-11}	$Al(OH)_3$	4.6×10^{-33}
$CaCO_3$	3.4×10^{-9}	$PbBr_2$	6.6×10^{-6}	$Ca(OH)_2$	5.0×10^{-6}
$CuCO_3$	2.5×10^{-10}	$PbCl_2$	1.7×10^{-5}	$Cu(OH)_2$	2.2×10^{-20}
$PbCO_3$	7.4×10^{-14}	PbF_2	3.3×10^{-8}	$Fe(OH)_2$	4.9×10^{-17}
$MgCO_3$	6.8×10^{-6}	PbI_2	9.8×10^{-9}	$Fe(OH)_3$	2.8×10^{-39}
Ag_2CO_3	8.5×10^{-12}	$AgCl$	1.8×10^{-10}	$Mg(OH)_2$	5.6×10^{-12}
$ZnCO_3$	1.5×10^{-10}	$AgBr$	5.4×10^{-13}	$Zn(OH)_2$	3×10^{-17}
Hg_2CO_3	3.6×10^{-17}	AgI	8.5×10^{-17}	**Sulfates**	
Chromates		**Phosphates**		$BaSO_4$	1.1×10^{-10}
$BaCrO_4$	1.2×10^{-10}	$AlPO_4$	9.8×10^{-21}	$CaSO_4$	4.9×10^{-5}
$PbCrO_4$	2.3×10^{-13}	$Ca_3(PO_4)_2$	2.1×10^{-33}	$PbSO_4$	2.5×10^{-8}
Ag_2CrO_4	1.1×10^{-12}	$Mg_3(PO_4)_2$	1.0×10^{-24}	Ag_2SO_4	1.2×10^{-5}

Suppose you wish to determine the solubility of silver iodide (AgI) in mol/L at 298 K. The equilibrium equation and solubility product constant expression are

$$AgI(s) \rightleftharpoons Ag^+(aq) + I^-(aq)$$
$$K_{sp} = [Ag^+][I^-] = 8.5 \times 10^{-17} \text{ at } 298 \text{ K}$$

The first thing you should do is let the symbol s represent the solubility of AgI; that is, the number of moles of AgI that dissolves in a liter of solution. The equation indicates that for every mole of AgI that dissolves, an equal number of moles of Ag^+ ions forms in solution. Therefore, $[Ag^+]$ equals s. Every Ag^+ has an accompanying I^- ion, so $[I^-]$ also equals s. Substituting s for $[Ag^+]$ and $[I^-]$, the K_{sp} expression becomes

$$[Ag^+][I^-] = (s)(s) = s^2 = 8.5 \times 10^{-17}$$
$$s = \sqrt{8.5 \times 10^{-17}} = 9.2 \times 10^{-9} \text{ mol/L.}$$

The solubility of AgI is 9.2×10^{-9} mol/L at 298 K.

EXAMPLE PROBLEM 18-5

Calculating Molar Solubility from K_{sp}

Use the K_{sp} value from **Table 18-3** to calculate the solubility in mol/L of copper(II) carbonate ($CuCO_3$) at 298 K.

1. Analyze the Problem

You have been given the solubility product constant for $CuCO_3$. The copper and carbonate ion concentrations are in a one-to-one relationship with the molar solubility of $CuCO_3$. Use the solubility product constant expression to solve for the solubility. Because K_{sp} is of the order of 10^{-10}, you can predict that the solubility will be the square root of K_{sp}, or about 10^{-5}.

Known	**Unknown**
K_{sp} ($CuCO_3$) = 2.5×10^{-10}	solubility of $CuCO_3$ = ? mol/L

2. Solve for the Unknown

Write the balanced chemical equation for the solubility equilibrium and the solubility product constant expression.

$CuCO_3(s) \rightleftharpoons Cu^{2+}(aq) + CO_3^{2-}(aq)$

$K_{sp} = [Cu^{2+}][CO_3^{2-}] = 2.5 \times 10^{-10}$

Relate the solubility to $[Cu^{2+}]$ and $[CO_3^{2-}]$.

$s = [Cu^{2+}] = [CO_3^{2-}]$

Substitute s for $[Cu^{2+}]$ and $[CO_3^{2-}]$ and solve for s.

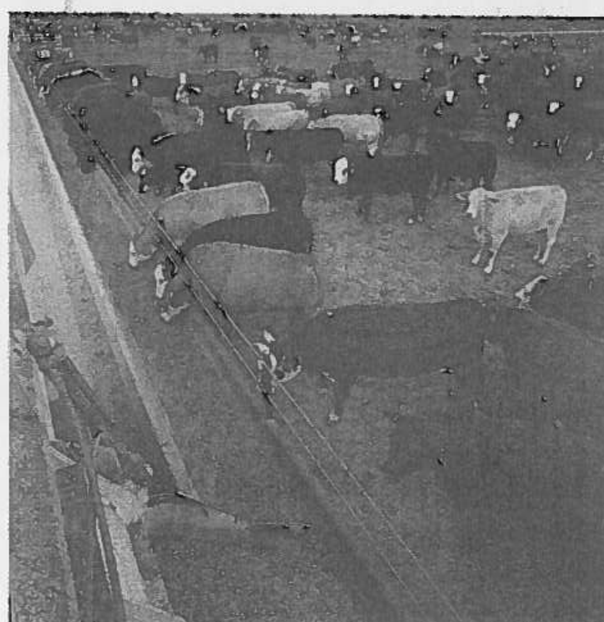
$(s)(s) = s^2 = 2.5 \times 10^{-10}$

$s = \sqrt{2.5 \times 10^{-10}} = 1.6 \times 10^{-5}$ mol/L

The molar solubility of $CuCO_3$ in water at 298 K is 1.6×10^{-5} mol/L.

3. Evaluate the Answer

The K_{sp} value has two significant figures, so the answer is correctly expressed with two digits. As predicted, the molar solubility of $CuCO_3$ is approximately 10^{-5} mol/L.



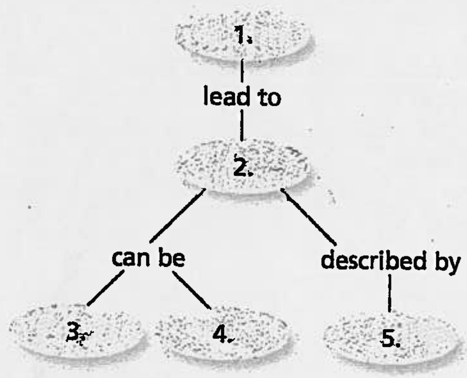
Finely ground copper carbonate is added to cattle and poultry feed to supply the necessary element, copper, to animal diets.

CHEMISTRY Online

Go to the Chemistry Web site at **science.glencoe.com** *or use the* **Chemistry CD-ROM** *for additional Chapter 18 Assessment.*

Concept Mapping

25. Fill in the spaces on the concept map with the following phrases: equilibrium constant expressions, reversible reactions, heterogeneous equilibria, homogeneous equilibria, chemical equilibria.



Mastering Concepts

26. Describe an equilibrium in everyday life that illustrates a state of balance between two opposing processes. (18.1)

27. Given the fact that the concentrations of reactants and products are not changing, why is the word *dynamic* used for describing chemical equilibrium? (18.1)

28. How can you indicate in a chemical equation that a reaction is reversible? (18.1)

29. Although the general equation for a chemical reaction is reactants $\rightarrow$ products, explain why this equation is not complete for a system at equilibrium. (18.1)

30. Explain the difference between a homogeneous equilibrium and a heterogeneous equilibrium. (18.1)

31. What is an equilibrium position? (18.1)

32. Explain how to use the law of chemical equilibrium in writing an equilibrium constant expression. (18.1)

33. Why does a numerically large K_{eq} mean that the products are favored in an equilibrium system? (18.1)

34. Why should you pay attention to the physical states of all reactants and products when writing equilibrium constant expressions? (18.1)

35. How can an equilibrium system contain small and unchanging amounts of products yet have large amounts of reactants? What can you say about the relative size of K_{eq} for such an equilibrium? (18.1)

36. Describe the opposing processes in the physical equilibrium that exists in a closed container half-filled with liquid ethanol. (18.1)

37. What is meant by a stress on a reaction at equilibrium (18.2)

38. How does Le Châtelier's principle describe an equilibrium's response to a stress? (18.2)

39. Why does removing a product cause an equilibrium to shift in the direction of the products? (18.2)

40. When an equilibrium shifts toward the reactants in response to a stress, how is the equilibrium position changed? (18.2)

41. Use Le Châtelier's principle to explain how a shift in the equilibrium $H_2CO_3(aq) \rightleftharpoons H_2O(l) + CO_2(g)$ causes a soft drink to go flat when its container is left open to the atmosphere. (18.2)

42. How is K_{eq} changed when heat is added to an equilibrium in which the forward reaction is exothermic? Explain using Le Châtelier's principle. (18.2)

43. Changing the volume of the system alters the equilibrium position of this equilibrium.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

But a similar change has no effect on this equilibrium.

$$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$$

Explain. (18.2)

44. How might the addition of a noble gas to the reaction vessel affect this equilibrium?

$$2N_2H_4(g) + 2NO_2(g) \rightleftharpoons 3N_2(g) + 4H_2O(g)$$

Assume that the volume of the reaction vessel does not change. (18.2)

45. When an equilibrium shifts to the right, what happens to the following? (18.2)

a. the concentrations of the reactants
b. the concentrations of the products

46. How would each of the following changes affect the equilibrium position of the system used to produce methanol from carbon monoxide and hydrogen? (18.2)

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g) + \text{heat}$$

a. adding CO to the system
b. cooling the system
c. adding a catalyst to the system
d. removing CH_3OH from the system
e. decreasing the volume of the system

47. Why is the concentration of a solid not included as part of the solubility product constant? (18.3)

48. What does it mean to say that two solutions have a common ion? Give an example that supports your answer. (18.3)

49. Explain the difference between Q_{sp} and K_{sp}. (18.3)

50. Explain why a common ion lowers the solubility of an ionic compound. (18.3)

51. Describe the solution that results when two solutions are mixed and Q_{sp} is found to equal K_{sp}. Does a precipitate form?

Mastering Problems

The Equilibrium Constant Expression (18.1)

52. Write equilibrium constant expressions for these homogeneous equilibria.

a. $2N_2H_4(g) + 2NO_2(g) \rightleftharpoons 3N_2(g) + 4H_2O(g)$
b. $2NbCl_4(g) \rightleftharpoons NbCl_3(g) + NbCl_5(g)$
c. $I_2(g) \rightleftharpoons 2I(g)$
d. $2SO_3(g) + CO_2(g) \rightleftharpoons CS_2(g) + 4O_2(g)$

53. Write equilibrium constant expressions for these heterogeneous equilibria.

a. $2NaHCO_3(s) \rightleftharpoons Na_2CO_3(s) + H_2O(g) + CO_2(g)$
b. $C_6H_6(l) \rightleftharpoons C_6H_6(g)$
c. $Fe_3O_4(s) + 4H_2(g) \rightleftharpoons 3Fe(s) + 4H_2O(g)$

54. Pure water has a density of 1.00 g/mL at 297 K. Calculate the molar concentration of pure water at this temperature.

55. Calculate K_{eq} for the following equilibrium when $[SO_3] = 0.0160$ mol/L, $[SO_2] = 0.00560$ mol/L, and $[O_2] = 0.00210$ mol/L.

$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$

56. K_{eq} for this reaction is 3.63.

$A + 2B \rightleftharpoons C$

The data in the table shows the concentrations of the reactants and product in two different reaction mixtures at the same temperature. Does the data provide evidence that both reactions are at equilibrium?

Table 18-5

Concentrations of A, B, and C

A (mol/L)	B (mol/L)	C (mol/L)
0.500	0.621	0.700
0.250	0.525	0.250

57. When solid ammonium chloride is put in a reaction vessel at 323 K, the equilibrium concentrations of both ammonia and hydrogen chloride are found to be 0.0660 mol/L. $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$. Calculate K_{eq}.

58. Suppose you have a cube of pure manganese metal measuring 5.25 cm on each side. You find that the mass of the cube is 1076.6 g. What is the molar concentration of manganese in the cube?

Le Châtelier's Principle (18.2)

59. Use Le Châtelier's principle to predict how each of the following changes would affect this equilibrium.
$H_2(g) + CO_2(g) \rightleftharpoons H_2O(g) + CO(g)$

a. adding $H_2O(g)$ to the system
b. removing $CO(g)$ from the system
c. adding $H_2(g)$ to the system
d. adding something to the system to absorb $CO_2(g)$

60. How would increasing the volume of the reaction vessel affect these equilibria?

a. $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$
b. $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$

61. How would decreasing the volume of the reaction vessel affect these equilibria?

a. $2N_2H_4(g) + 2NO_2(g) \rightleftharpoons 3N_2(g) + 4H_2O(g)$
b. $2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g)$

62. How would these equilibria be affected by increasing the temperature?

a. $4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g) + \text{heat}$
b. $\text{heat} + NaCl(s) \rightleftharpoons Na^+(aq) + Cl^-(aq)$

63. Ethylene (C_2H_4) reacts with hydrogen to form ethane (C_2H_6).
$C_2H_4(g) + H_2(g) \rightleftharpoons C_2H_6(g) + \text{heat}$
How would you regulate the temperature of this equilibrium in order to do the following?

a. increase the yield of ethane
b. decrease the concentration of ethylene
c. increase the amount of hydrogen in the system

64. How would simultaneously decreasing the temperature and volume of the system affect these equilibria?

a. $\text{heat} + CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$
b. $4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g) + \text{heat}$

Calculations Using K_{eq} (18.3)

65. K_{eq} is 1.60 at 933 K for this reaction.
$H_2(g) + CO_2(g) \rightleftharpoons H_2O(g) + CO(g)$
Calculate the equilibrium concentration of hydrogen when $[CO_2] = 0.320$ mol/L, $[H_2O] = 0.240$ mol/L and $[CO] = 0.280$ mol/L.

66. At 2273 K, $K_{eq} = 6.2 \times 10^{-4}$ for the reaction
$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$
If $[N_2] = 0.05200$ mol/L and $[O_2] = 0.00120$ mol/L, what is the concentration of NO at equilibrium?

Calculations Using K_{sp} (18.3)

67. Calculate the ion product to determine if a precipitate will form when 125 mL 0.00500*M* sodium chloride is mixed with 125 mL 0.00100*M* silver nitrate solution.

68. Calculate the molar solubility of strontium chromate in water at 298 K if $K_{sp} = 3.5 \times 10^{-5}$.

69. Will a precipitate form when 1.00 L of 0.150*M* iron(II) chloride solution is mixed with 2.00 L of 0.0333*M* sodium hydroxide solution? Explain your reasoning and show your calculations.

Mixed Review

Sharpen your problem-solving skills by answering the following.

70. How many moles per liter of silver chloride will be in a saturated solution of AgCl? $K_{sp} = 1.8 \times 10^{-10}$

71. A 6.00-L vessel contains an equilibrium mixture of 0.0222 mol PCl_3, 0.0189 mol PCl_5, and 0.1044 mol Cl_2. Calculate K_{eq} for the following reaction.
$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$

72. How would simultaneously increasing the temperature and volume of the system affect these equilibria?
a. $2O_3(g) \rightleftharpoons 3O_2(g) + \text{heat}$
b. $\text{heat} + N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$

73. The solubility product constant for lead(II) arsenate ($Pb_3(AsO_4)_2$), is 4.0×10^{-36} at 298 K. Calculate the molar solubility of the compound at this temperature.

74. How would these equilibria be affected by decreasing the temperature?
a. $2O_3(g) \rightleftharpoons 3O_2(g) + \text{heat}$
b. $\text{heat} + H_2(g) + F_2(g) \rightleftharpoons 2HF(g)$

Thinking Critically

75. **Predicting** Suppose you're thinking about using the following reaction to produce hydrogen from hydrogen sulfide.
$2H_2S(g) + \text{heat} \rightleftharpoons 2H_2(g) + S_2(g)$
Given that K_{eq} for the equilibrium is 2.27×10^{-4}, would you expect a high yield of hydrogen? Explain how you could regulate the volume of the reaction vessel and the temperature to increase the yield.

76. **Applying Concepts** Smelling salts, sometimes used to revive a groggy or unconscious person, are made of ammonium carbonate. The equation for the endothermic decomposition of ammonium carbonate is
$(NH_4)_2CO_3(s) \rightleftharpoons 2NH_3(g) + CO_2(g) + H_2O(g)$
Would you expect smelling salts to work as well on a cold winter day as on a warm summer day? Explain your answer.

77. **Comparing and Contrasting** Which of the two solids, calcium phosphate or iron(III) phosphate, has the greater molar solubility? K_{sp} ($Ca_3(PO_4)_2$) = 1.2×10^{-29}; K_{sp} ($FePO_4$) = 1.0×10^{-22}. Which compound has the greater solubility expressed in grams per liter?

78. **Recognizing Cause and Effect** You have 12.56 g of a mixture made up of sodium chloride and barium chloride. Explain how you could use a precipitation reaction to determine how much of each compound the mixture contains.

Writing in Chemistry

79. Research the role that solubility plays in the formation of kidney stones. Find out what compounds are found in kidney stones and their K_{sp} values. Summarize your findings in a report.

80. The presence of magnesium and calcium ions in water makes the water "hard." Explain in terms of solubility why the presence of these ions is often undesirable. Find out what measures can be taken to eliminate them.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

81. How are elctrons shared differently in H_2, O_2, and N_2? (Chapter 9)

82. How can you tell if a chemical equation is balanced? (Chapter 10)

83. What mass of carbon must burn to produce 4.56 L CO_2 gas at STP? (Chapter 14)
$C(s) + O_2(g) \rightarrow CO_2(g)$

84. When you reverse a thermochemical equation, why must you change the sign of ΔH? (Chapter 16)

Use these questions and the test-taking tip to prepare for your standardized test.

1. A system reaches chemical equilibrium when _____ .
   a. no new product is formed by the forward reaction
   b. the reverse reaction no longer occurs in the system
   c. the concentration of reactants in the system is equal to the concentration of products
   d. the rate at which the forward reaction occurs equals the rate of the reverse reaction

2. A value of K_{eq} greater than 1 means that _____ .
   a. more reactants than products exist at equilibrium
   b. more products than reactants exist at equilibrium
   c. the rate of the forward reaction is high at equilibrium
   d. the rate of the reverse reaction is high at equilibrium

3. The hydrogen sulfide produced as a byproduct of petroleum refinement can be used to produce elemental sulfur: $2H_2S(g) + SO_2(g) \rightarrow 3S(l) + 2H_2O(g)$
   The equilibrium constant expression for this reaction is _____ .
   a. $K_{eq} = \dfrac{[H_2O]}{[H_2S][SO_2]}$
   b. $K_{eq} = \dfrac{[H_2S]^2[SO_2]}{[H_2S]^2}$
   c. $K_{eq} = \dfrac{[H_2O]^2}{[H_2S]^2[SO_2]}$
   d. $K_{eq} = \dfrac{[S]^3[H_2O]^2}{[H_2S]^2[SO_2]}$

4. The following system is in equilibrium:
   $2S(s) + 5F_2(g) \rightleftharpoons SF_4(g) + SF_6(g)$
   The equilibrium will shift to the right if _____ .
   a. the concentration of SF_4 is increased
   b. the concentration of SF_6 is increased
   c. the pressure on the system is increased
   d. the pressure on the system is decreased

Interpreting Tables Use the table to answer questions 5–7.

Concentration Data for the Equilibrium System $MnCO_3(s) \rightleftharpoons Mn^{2+}(aq) + CO_3^{2-}(aq)$ at 298 K

Trial	$[Mn^{2+}]_0$ (*M*)	$[CO_3^{2-}]_0$ (*M*)	$[Mn^{2+}]_{eq}$ (*M*)	$[CO_3^{2-}]_{eq}$ (*M*)
1	0.0000	0.00400	5.60×10^{-9}	4.00×10^{-3}
2	0.0100	0.0000	1.00×10^{-2}	2.24×10^{-9}
3	0.0000	0.0200	1.12×10^{-9}	2.00×10^{-2}

5. The K_{sp} for $MnCO_3$ is _____ .
   a. 2.24×10^{-11}
   b. 4.00×10^{-11}
   c. 1.12×10^{-9}
   d. 5.60×10^{-9}

6. What is the molar solubility of $MnCO_3$ at 298 K?
   a. $4.73 \times 10^{-6}M$
   b. $6.32 \times 10^{-2}M$
   c. $7.48 \times 10^{-5}M$
   d. $3.35 \times 10^{-5}M$

7. A 50.0-mL volume of $3.00 \times 10^{-6}M$ K_2CO_3 is mixed with 50.0 mL of $MnCl_2$. A precipitate of $MnCO_3$ will form only when the concentration of the $MnCl_2$ solution is greater than _____ .
   a. $7.47 \times 10^{-6}M$
   b. $1.49 \times 10^{-5}M$
   c. $2.99 \times 10^{-5}M$
   d. $1.02 \times 10^{-5}M$

8. Which of the following statements about the common ion effect is NOT true?
   a. The effects of common ions on an equilibrium system can be explained by Le Châtelier's principle.
   b. The decreased solubility of an ionic compound due to the presence of a common ion is called the common ion effect.
   c. The addition of NaCl to a saturated solution of AgCl will produce the common ion effect.
   d. The common ion effect is due to a shift in equilibrium towards the aqueous products of a system.

9. If the forward reaction of a system in equilibrium is endothermic, increasing the temperature of the system will _____ .
   a. shift the equilibrium to the left
   b. shift the equilibrium to the right
   c. decrease the rate of the forward reaction
   d. decrease the rate of the reverse reaction

10. $Cl_2(g) + 3O_2(g) + F_2(g) \rightleftharpoons 2ClO_3F(g)$
    The formation of perchloryl fluoride (ClO_3F) from its elements has an equilibrium constant of 3.42×10^{-9} at 298 K. If $[Cl_2] = 0.563M$, $[O_2] = 1.01M$, and $[ClO_3F] = 1.47 \times 10^{-5}M$ at equilibrium, what is the concentration of F_2?
    a. $9.18 \times 10^{0}M$
    b. $3.73 \times 10^{-10}M$
    c. $1.09 \times 10^{-1}M$
    d. $6.32 \times 10^{-2}M$

THE PRINCETON REVIEW **TEST-TAKING TIP**

Maximize Your Score If possible, find out how your standardized test will be scored. In order to do your best, you need to know if there is a penalty for guessing, and if so, what the penalty is. If there is no random-guessing penalty at all, you should always fill in an answer, even if you haven't read the question!