The concentrations of $PCl_{3(g)}$ and $Cl_{2(g)}$ are both 0.100 mol/L.

9. (a)

ICE Table for the Production of Methanol			
	$CO_{(g)} +$	$2\ H_{2(g)} \rightleftharpoons$	$CH_3OH_{(g)}$
Initial concentration (mol/L)	0.100	0.200	0
Change in concentration (mol/L)	–0.040	–0.080	+0.040
Equilibrium concentration (mol/L)	0.060	0.120	0.040

The concentrations of carbon monoxide and methanol are 0.060 mol/L and 0.0400 mol/L, respectively.

(b) $$\text{percent reaction} = \frac{\text{actual product yield}}{\text{theoretical product yield}} \times 100\%$$

$$= \frac{0.0400}{0.100} \times 100\%$$

percent reaction = 40%

10. (a) The saturated solution is in a state of dynamic equilibrium. Even though the appearance of the system did not change, the evidence suggests that some radioactive iodine became incorporated into the original sample of iodine. Furthermore, the radioactivity of the solution implies that some radioactive iodine had dissolved.

Making Connections

11. (a) Immediately following a fructose-rich meal, a net amount of diffusion of fructose will occur into the (surrounding) intestinal cells because the concentration of fructose in the intestinal tract is higher than the concentration of fructose in the intestinal cells. Diffusion into the cells continues until the concentration of fructose inside the intestinal cells and the intestinal tract are equal. At this point, a dynamic equilibrium is reached; the rate of fructose diffusing into the cells equals the rate of fructose diffusing out of the cells. Thus, a certain amount of fructose remains in the intestinal tract (unabsorbed).

(b) Cells use active transport mechanisms to pump nutrient molecules such as glucose and amino acids (digested protein). These systems require the use of cell membrane proteins and cell energy in the form of adenosine triphosphate (ATP). Using energy to actively pump certain nutrients against a concentration gradient ensures that intestinal cells absorb the majority of nutrient molecules.

7.2 EQUILIBRIUM LAW IN CHEMICAL REACTIONS

PRACTICE

(Page 442)

Understanding Concepts

1. (a) $K = \dfrac{[SO_{3(g)}]^2}{[SO_{2(g)}]^2[O_{2(g)}]}$

(b) $4\ NH_{3(g)} + 3\ O_{2(g)} \rightleftharpoons 2\ N_{2(g)} + 6\ H_2O_{(g)}$

$K = \dfrac{[N_{2(g)}]^2[H_2O_{(g)}]^6}{[NH_{3(g)}]^4[O_{2(g)}]^3}$

(c) $K = \dfrac{[NO_{(g)}]^2[Br_{2(g)}]}{[NOBr_{(g)}]^2}$

PRACTICE

(Page 444)

Understanding Concepts

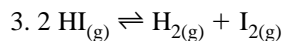2. $CO_{(g)} + 2\ H_{2(g)} \rightleftharpoons CH_3OH_{(g)}$

$$K = \frac{[CH_3OH_{(g)}]}{[CO_{(g)}][H_{2(g)}]^2}$$

$$= \frac{[0.021]}{[0.079][0.158]^2}$$

$$K = 10.6$$

The value of the equilibrium constant for this reaction at 327°C is 10.6.

3. $2\ HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)}$

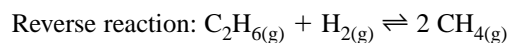$$K = \frac{[H_{2(g)}][I_{2(g)}]}{[HI_{(g)}]^2}$$
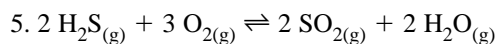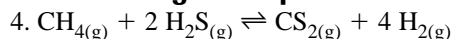
$$= \frac{[1.065][0.065]}{[1.870]^2}$$

$$K = 0.020$$

The value of the equilibrium constant for this reaction at 440°C is 0.020.

PRACTICE

(Page 444)

Understanding Concepts

4. $CH_{4(g)} + 2\ H_2S_{(g)} \rightleftharpoons CS_{2(g)} + 4\ H_{2(g)}$

5. $2\ H_2S_{(g)} + 3\ O_{2(g)} \rightleftharpoons 2\ SO_{2(g)} + 2\ H_2O_{(g)}$

6. Forward reaction: $2\ CH_{4(g)} \rightleftharpoons C_2H_{6(g)} + H_{2(g)}$

   Reverse reaction: $C_2H_{6(g)} + H_{2(g)} \rightleftharpoons 2\ CH_{4(g)}$

$$K' = \frac{[CH_{4(g)}]^2}{[H_{2(g)}][C_2H_{6(g)}]}$$

$$= \frac{1}{K}$$
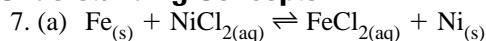
$$= \frac{1}{9.5 \times 10^{-13}}$$

$$K' = 1.1 \times 10^{12}$$

The value of the equilibrium constant for the reverse reaction is 1.1×10^{12}.
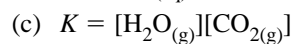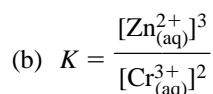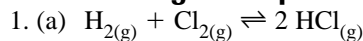
PRACTICE

(Page 447)

Understanding Concepts

7. (a) $Fe_{(s)} + NiCl_{2(aq)} \rightleftharpoons FeCl_{2(aq)} + Ni_{(s)}$

$$K = \frac{[Fe^{2+}_{(aq)}]}{[Ni^{2+}_{(aq)}]}$$

   (b) $K = \dfrac{[Zn^{2+}_{(aq)}]^3}{[Cr^{3+}_{(aq)}]^2}$

   (c) $K = [H_2O_{(g)}][CO_{2(g)}]$

SECTION 7.2 QUESTIONS

(Page 448)

Understanding Concepts

1. (a) $H_{2(g)} + Cl_{2(g)} \rightleftharpoons 2\ HCl_{(g)}$

$$K = \frac{[HCl_{(g)}]^2}{[H_{2(g)}][Cl_{2(g)}]}$$

(b) $N_{2(g)} + 3\ H_{2(g)} \rightleftharpoons 2\ NH_{3(g)}$

$$K = \frac{[NH_{3(g)}]^2}{[N_{2(g)}][H_{2(g)}]^3}$$

(c) $2\ H_{2(g)} + O_{2(g)} \rightleftharpoons 2\ H_2O_{(g)}$

$$K = \frac{[H_2O_{(g)}]^2}{[H_{2(g)}]^2[O_{2(g)}]}$$

(d) $Ni^{2+}_{(aq)} + 6\ NH_{3(aq)} \rightleftharpoons Ni(NH_3)^{2+}_{6(aq)}$

$$K = \frac{[Ni(NH_3)_6{}^{2+}{}_{(aq)}]}{[Ni^{2+}_{(aq)}][NH_{3(aq)}]^6}$$

(e) $CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$
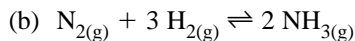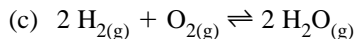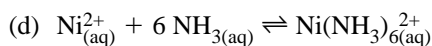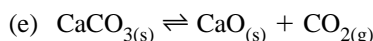
$$K = [CO_{2(g)}]$$

(f) $H_2CO_{3(aq)} \rightleftharpoons H_2O_{(l)} + CO_{2(g)}$
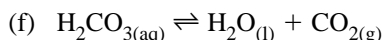
$$K = \frac{[CO_{2(g)}]}{[H_2CO_{3(aq)}]}$$
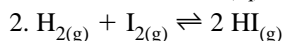
2. $H_{2(g)} + I_{2(g)} \rightleftharpoons 2\ HI_{(g)}$

$$K = \frac{[HI_{2(g)}]^2}{[H_{2(g)}][I_{2(g)}]}$$

$$= \frac{[3.116]^2}{[0.442][0.442]}$$

$$K = 49.7$$

The value of the equilibrium constant at 485°C is 49.7.

3. $PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$

$$K = \frac{[PCl_{3(g)}][Cl_{2(g)}]}{[PCl_{5(g)}]}$$

$$= \frac{[0.014][0.014]}{[4.3 \times 10^{-4}]}$$

$$K = 0.46$$

The value of the equilibrium constant at 200°C is 0.46.

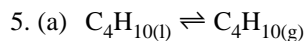4. $N_{2(g)} + 3\ H_{2(g)} \rightleftharpoons 2\ NH_{3(g)}$
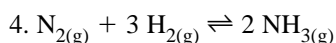
$$K = \frac{[NH_{3(g)}]^2}{[N_{2(g)}][H_{2(g)}]^3}$$

$$K = 8.00 \times 10^{-7}$$

$$[NH_{3(g)}]^2 = 8.00 \times 10^{-7} \times [N_{2(g)}][H_{2(g)}]^3$$
$$= 8.00 \times 10^{-7} [1.50][0.50]^3$$
$$[NH_{3(g)}]^2 = 1.5 \times 10^{-7}$$
$$[NH_{3(g)}] = 3.9 \times 10^{-4} \text{ mol/L}$$

The equilibrium concentration of ammonia is 3.9×10^{-4} mol/L.

5. (a) $C_4H_{10(l)} \rightleftharpoons C_4H_{10(g)}$

$$K = [C_4H_{10(g)}]$$

(b) The amount of butane in the gas phase varies directly with the temperature. As the temperature increases, for example, more butane will be found in the gas phase. However, the butane concentration at a particular temperature is a constant. Since the value of the equilibrium constant depends only on $[C_4H_{10(g)}]$, it also is constant.

(c) The concentration of butane gas depends only on the temperature and not on the amount of liquid butane. If the volume of liquid butane decreases, there's more space above the liquid for gaseous butane to occupy. The number of moles of butane may have increased but the mol/L remains the same.

6. (a) $2 \text{ HBr}_{(g)} \rightleftharpoons H_{2(g)} + Br_{2(g)}$
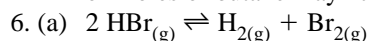
(b) $K = \dfrac{[H_{2(g)}][Br_{2(g)}]}{[HBr_{(g)}]^2}$

(c) $n_{HBr_{(g)}} = C_{HBr_{(g)}} \times 2.00 \text{ L}$
$$= 0.100 \text{ mol/L} \times 2.00 \text{ L}$$
$$n_{HBr_{(g)}} = 0.200 \text{ mol}$$
0.200 mol of $HBr_{(g)}$ is present at equilibrium.

(d) $n_{HBr} = 0.100 \text{ mol/L} \times 2.00 \text{ L}$
$$n_{HBr} = 0.200 \text{ mol}$$
0.200 mol of hydrogen bromide is present at equilibrium.

$$n_{HBr_{reacted}} = n_{HBr_{initial}} - n_{HBr_{equilibrium}}$$
$$= 1.00 \text{ mol} - 0.200 \text{ mol}$$
$$n_{HBr_{reacted}} = 0.800 \text{ mol}$$

0.800 mol of hydrogen bromide reacted.

(e) Given the 2:1 ratio, 0.400 mol of both hydrogen and bromine are produced.

(f) $[H_2] = \dfrac{0.400 \text{ mol}}{2.00 \text{ L}}$
$$= 0.200 \text{ mol/L}$$

$$[H_2] = [Br_2]$$

The concentrations of hydrogen and bromine are both 0.200 mol/L, while the hydrogen bromide concentration is 0.100 mol/L.

(g) $K = \dfrac{[H_{2(g)}][Br_{2(g)}]}{[HBr_{(g)}]^2}$
$$= \dfrac{[0.20][0.20]}{[0.100]^2}$$
$$K = 4.0$$

The value of the equilibrium constant is 4.00.

7. (a) One could specify either the value of the equilibrium constant or the extent of reaction (as percent reaction).
   (b) Both methods measure the extent to which reactants become products in a closed system. However, percent reaction describes the extent of the reaction as the yield of product at equilibrium compared to the maximum possible yield. The value of the equilibrium constant gives the ratio of products to reactants at equilibrium. Furthermore, the percent reaction is variable, for a given reaction, depending on the concentration of reactants; the equilibrium constant is independent of concentration.
   (c) Both methods are effective in describing the extent of the reaction. However, they also have their limitations. The value of the equilibrium constant is valid only for a given temperature. Percent reaction values depend not only on temperature but on concentration as well.

Applying Inquiry Skills
8. **Analysis**

(a) $$K = \frac{[SO_{3(g)}]^2}{[SO_{2(g)}]^2[O_{2(g)}]}$$

For trial 1…

$$K = \frac{[3.50 \times 10^{-2}]^2}{[1.50 \times 10^{-1}]^2[1.26 \times 10^{-2}]}$$

$K = 4.32$

For trial 2…

$$K = \frac{[2.60 \times 10^{-3}]^2}{[5.90 \times 10^{-2}]^2[4.50 \times 10^{-4}]}$$

$K = 4.32$

For trial 3…

$$K = \frac{[3.6 \times 10^{-3}]^2}{[1.00 \times 10^{-2}]^2[3.0 \times 10^{-2}]}$$

$K = 4.32$

Making Connections
9. The small value of the equilibrium constant suggests that the equilibrium strongly favours oxygen. Only a negligible amount of ozone is produced by this reaction.

7.3   QUALITATIVE CHANGES IN EQUILIBRIUM SYSTEMS

PRACTICE

(Page 457)

Understanding Concepts
1. The equilibrium will shift to the
   (a) left
   (b) right
   (c) left
   (d) right
   (e) right