

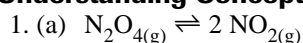
- (b) More nitrogen is fixed naturally (1.75×10^8 tonnes per year) than synthetically (8×10^7 tonnes per year).
- (c) Studies show that high levels of nitrogen fixation caused by nitrogen pollution are likely to cause severe declines in native plant communities by favouring weedy plants that favour high ammonia concentrations in soil. This problem can be reduced by decreasing the use of machines that burn fossil fuels such as automobiles, trucks, buses, planes, leaf blowers, lawn mowers, and jet skis, and by reducing the nitrogen oxide emissions of coal- and oil-burning electricity generating plants and large industries.

7.5 QUANTITATIVE CHANGES IN EQUILIBRIUM SYSTEMS

PRACTICE

(Page 465)

Understanding Concepts



$$[\text{N}_2\text{O}_{4(g)}] = 5.30 \text{ mol/L}$$

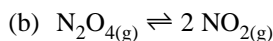
$$[\text{NO}_{2(g)}] = 2.15 \text{ mol/L}$$

$$Q = \frac{[\text{NO}_{2(g)}]^2}{[\text{N}_2\text{O}_{4(g)}]}$$

$$= \frac{[2.15]^2}{[5.30]}$$

$$Q = 0.87$$

Since the value of Q is equal to the value of K (0.87), the reaction is at equilibrium.



$$[\text{N}_2\text{O}_{4(g)}] = 0.80 \text{ mol/L}$$

$$[\text{NO}_{2(g)}] = 1.55 \text{ mol/L}$$

$$Q = \frac{[\text{NO}_{2(g)}]^2}{[\text{N}_2\text{O}_{4(g)}]}$$

$$= \frac{[1.15]^2}{[0.80]}$$

$$Q = 3.0$$

Since the value of Q is larger than the value of K (0.87), the reaction is not at equilibrium. In order to attain equilibrium, the reaction will shift to the left (as written). The concentration of the reactants will increase and the concentration of the products will decrease.

2. $[\text{Cl}_{2(g)}]_{\text{initial}} = \frac{3.20 \text{ mol}}{1.00 \text{ L}}$

$$[\text{Cl}_{2(g)}]_{\text{initial}} = 3.20 \text{ mol/L}$$

$$[\text{PCl}_{3(g)}]_{\text{initial}} = \frac{1.5 \text{ mol}}{1.0 \text{ L}}$$

$$[\text{PCl}_{3(g)}]_{\text{initial}} = 1.5 \text{ mol/L}$$

$$[\text{PCl}_{5(g)}]_{\text{initial}} = \frac{2.0 \text{ mol}}{1.0 \text{ L}}$$

$$[\text{PCl}_{5(g)}]_{\text{initial}} = 2.0 \text{ mol/L}$$

$$Q = \frac{[\text{PCl}_{3(g)}][\text{Cl}_{2(g)}]}{[\text{PCl}_{5(g)}]}$$

$$= \frac{[1.5][3.20]}{[2.0]}$$

$$Q = 2.4$$

Since the value of Q is smaller than the value of K (12.5), the reaction is not at equilibrium. In order to attain equilibrium, the reaction will shift to the right (as written). The concentration of the reactants will decrease and the concentration of the products will increase.

PRACTICE

(Page 466)

Understanding Concepts

$$3. \quad K = \frac{[\text{NH}_{3(g)}]^2}{[\text{N}_{2(g)}][\text{H}_{2(g)}]^3}$$

$$K = 626$$

$$[\text{N}_{2(g)}] = \frac{[\text{NH}_{3(g)}]^2}{626[\text{H}_{2(g)}]^3}$$

$$= \frac{(0.46)^2}{626(0.50)^3}$$

$$[\text{N}_{2(g)}] = 2.7 \times 10^{-3} \text{ mol/L}$$

The equilibrium concentration of nitrogen is $2.7 \times 10^{-3} \text{ mol/L}$.



$$K = \frac{[\text{PCl}_{3(g)}][\text{Cl}_{2(g)}]}{[\text{PCl}_{5(g)}]}$$

$$K = 32$$

$$[\text{PCl}_{5(g)}] = \frac{[\text{PCl}_{3(g)}][\text{Cl}_{2(g)}]}{32}$$

$$= \frac{[0.80][0.70]}{32}$$

$$[\text{PCl}_{5(g)}] = 1.80 \times 10^{-2} \text{ mol/L}$$

The equilibrium concentration of phosphorus pentachloride is $1.8 \times 10^{-3} \text{ mol/L}$.

PRACTICE

(Page 472)

Understanding Concepts

5. Initial concentrations...

$$[\text{CO}_{2(g)}] = [\text{H}_{2(g)}]$$

$$= \frac{1.00 \text{ mol}}{10.0 \text{ L}}$$

$$[\text{CO}_{2(g)}] = 0.100 \text{ mol/L}$$

ICE Table for the Reaction of $\text{CO}_{2(g)}$ with $\text{H}_{2(g)}$				
	$\text{CO}_{2(g)} +$	$\text{H}_{2(g)} \rightleftharpoons$	$\text{CO}_{(g)} +$	$\text{H}_2\text{O}_{(g)}$
Initial concentration (mol/L)	0.100	0.100	0.000	0.000
Change in concentration (mol/L)	$-x$	$-x$	$+x$	$+x$
Equilibrium concentration (mol/L)	$0.100 - x$	$0.100 - x$	x	x

$$K = \frac{[\text{CO}_{(\text{g})}][\text{H}_2\text{O}_{(\text{g})}]}{[\text{CO}_{2(\text{g})}][\text{H}_{2(\text{g})}]}$$

$$\frac{x^2}{(0.100 - x)^2} = 1.60$$

$$\frac{x}{(0.100 - x)} = \sqrt{1.60}$$

$$\frac{x}{(0.100 - x)} = 1.2649 \text{ (extra digits carried)}$$

$$1.2649(0.100 - x) = x$$

$$0.12649 - 1.2649x = x$$

$$0.12649 = 2.2649x$$

$$x = 5.58 \times 10^{-2}$$

$$[\text{CO}_{2(\text{g})}] = [\text{H}_{2(\text{g})}]$$

$$= 0.100 - (5.58 \times 10^{-2})$$

$$[\text{CO}_{2(\text{g})}] = 0.0442 \text{ mol/L}$$

The concentrations of hydrogen and carbon dioxide are both 0.0442 mol/L, and the concentrations of carbon monoxide and water are both 0.0558 mol/L.

6. Initial concentrations...

$$[\text{I}_{2(\text{g})}] = [\text{Cl}_{2(\text{g})}]$$

$$= \frac{0.50 \text{ mol}}{200.0 \text{ L}}$$

$$[\text{I}_{2(\text{g})}] = 0.25 \text{ mol/L}$$

ICE Table for the Reaction of $\text{I}_{2(\text{g})}$ with $\text{Cl}_{2(\text{g})}$			
	$\text{I}_{2(\text{g})} +$	$\text{Cl}_{2(\text{g})} \rightleftharpoons$	$2 \text{ ICl}_{(\text{g})}$
Initial concentration (mol/L)	0.25	0.25	
Change in concentration (mol/L)	$-x$	$-x$	$+2x$
Equilibrium concentration (mol/L)	$0.25 - x$	$0.25 - x$	$2x$

$$K = \frac{[\text{ICl}_{(\text{g})}]^2}{[\text{I}_{2(\text{g})}][\text{Cl}_{2(\text{g})}]}$$

$$\frac{(2x^2)}{(0.25 - x)^2} = 81.9$$

$$\frac{2x}{(0.25 - x)} = \sqrt{81.9}$$

$$\frac{2x}{(0.25 - x)} = 9.05$$

$$9.05(0.25 - x) = 2x$$

$$2.2625 - 9.05x = 2x \quad \text{(extra digits carried)}$$

$$2.2625 = 11.05x$$

$$x = 0.2048$$

$$[\text{I}_{2(\text{g})}] = [\text{Cl}_{2(\text{g})}]$$

$$= 0.25 - 0.2251$$

$$[\text{I}_{2(\text{g})}] = 0.045 \text{ mol/L}$$

$$[\text{ICl}_{(\text{g})}] = 2x$$

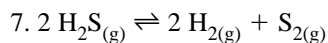
$$[\text{ICl}_{(\text{g})}] = 0.41 \text{ mol/L}$$

At equilibrium, the concentrations of iodine and chlorine are 0.045 mol/L and the concentration of the iodine chloride is 0.41 mol/L.

PRACTICE

(Page 476)

Understanding Concepts



ICE Table for the Decomposition of Hydrogen Sulfide			
	$2 \text{H}_2\text{S}_{(\text{g})} \rightleftharpoons$	$2 \text{H}_{2(\text{g})} +$	$\text{S}_{2(\text{g})}$
Initial concentration (mol/L)	0.200		
Change in concentration (mol/L)	$-2x$	$+2x$	$+x$
Equilibrium concentration (mol/L)	$0.200 - 2x$	$+2x$	$+x$

$$K = \frac{[\text{H}_{2(\text{g})}]^2[\text{S}_{2(\text{g})}]}{[\text{H}_2\text{S}_{(\text{g})}]^2}$$

$$K = 4.20 \times 10^{-6}$$

$$\frac{[2x]^2[x]}{[0.200 - 2x]^2} = 4.20 \times 10^{-6}$$

$$\frac{4x^3}{[0.200 - 2x]^2} = 4.20 \times 10^{-6}$$

If we assume $0.200 - 2x \doteq 0.200\dots$

$$\frac{4x^3}{[0.200]^2} \doteq 4.20 \times 10^{-6}$$

$$x^3 \doteq 4.2 \times 10^{-8}$$

$$x \doteq 3.48 \times 10^{-3}$$

Validate the assumption...

$$= \frac{0.200 - 2 \times 3.48 \times 10^{-3}}{0.200} \times 100$$

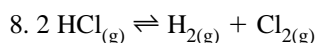
$$= 3.48\%$$

Since $3.48\% < 5\%$, the assumption is justified.

$$[\text{S}_{2(\text{g})}] = x$$

$$= 3.48 \times 10^{-3} \text{ mol/L}$$

The equilibrium concentration of $\text{S}_{2(\text{g})}$ is $3.48 \times 10^{-3} \text{ mol/L}$.



ICE Table for the Decomposition of Hydrogen Chloride			
	$2 \text{HCl}_{(\text{g})} \rightleftharpoons$	$\text{H}_{2(\text{g})} +$	$\text{Cl}_{2(\text{g})}$
Initial concentration (mol/L)	2.00		
Change in concentration (mol/L)	$-2x$	$+x$	$+x$
Equilibrium concentration (mol/L)	$2.00 - 2x$	x	x

$$K = \frac{[\text{H}_{2(\text{g})}][\text{Cl}_{2(\text{g})}]}{[\text{HCl}_{(\text{g})}]^2}$$

$$K = 3.2 \times 10^{-34}$$

$$\frac{x^2}{[2.00 - 2x]^2} = 3.2 \times 10^{-34}$$

$$\sqrt{\frac{x^2}{[2.00 - 2x]^2}} = \sqrt{3.2 \times 10^{-34}}$$

$$\frac{x}{[2.00 - 2x]} = 1.789 \times 10^{-17} \quad (\text{extra digits carried})$$

If we assume that $2.00 - x \doteq 2.00$

$$\frac{x}{2.00} \doteq 1.789 \times 10^{-17}$$

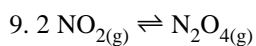
$$x \doteq 3.58 \times 10^{-17}$$

The equilibrium concentrations of hydrogen and chlorine are 3.58×10^{-17} and the equilibrium concentration of hydrogen chloride is 2.00 mol/L.

PRACTICE

(Page 480)

Understanding Concepts



Initial concentrations are

$$[\text{NO}_{2(\text{g})}] = 0.650 \text{ mol/L}$$

$$[\text{N}_2\text{O}_{4(\text{g})}] = 0.000 \text{ mol/L}$$

$$Q = \frac{[\text{N}_2\text{O}_{4(\text{g})}]}{[\text{NO}_{2(\text{g})}]^2}$$

$$= \frac{0.000}{(0.650)^2}$$

$$Q = 0.000$$

Since Q is less than K , the reaction will proceed to the right.

ICE Table for the Formation of $\text{N}_2\text{O}_{4(\text{g})}$		
	$2 \text{NO}_{2(\text{g})} \rightleftharpoons$	$\text{N}_2\text{O}_{4(\text{g})}$
Initial concentration (mol/L)	0.650	0.000
Change in concentration (mol/L)	$-2x$	$+x$
Equilibrium concentration (mol/L)	$0.650 - 2x$	$+x$

At equilibrium...

$$K = \frac{[\text{N}_2\text{O}_{4(\text{g})}]}{[\text{NO}_{2(\text{g})}]^2}$$

$$\frac{x}{(0.65 - 2x)^2} = 1.15$$

$$1.15(0.65 - 2x)^2 = x$$

$$1.15(0.4225 - 2.60x + 4x^2) = x \quad (\text{extra digits carried})$$

$$4.60x^2 - 3.99x + 0.486 = 0$$

$$x = \frac{3.99 \pm \sqrt{(-3.99)^2 - 4(4.60)(0.486)}}{2(4.60)}$$

$$x = 0.7209 \text{ or } 0.1465 \text{ (extra digits carried)}$$

x cannot be equal to 0.7209, since this value will give negative concentrations.

$$x = 0.1465$$

$$[\text{N}_2\text{O}_{4(\text{g})}] = x$$

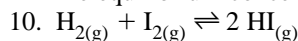
$$= 0.147 \text{ mol/L}$$

$$[\text{NO}_{2(\text{g})}] = 0.65 - 2x$$

$$= 0.65 - 2(0.1465)$$

$$[\text{NO}_{2(\text{g})}] = 0.357 \text{ mol/L}$$

The equilibrium concentrations of $\text{NO}_{2(\text{g})}$ and $\text{N}_2\text{O}_{4(\text{g})}$ are 0.357 mol/L and 0.147 mol/L, respectively.



Initial concentrations are

$$[\text{HI}_{(\text{g})}] = 0.00$$

$$[\text{H}_{2(\text{g})}] = [\text{I}_{2(\text{g})}]$$

$$= \frac{0.40 \text{ mol}}{0.50 \text{ L}}$$

$$[\text{H}_{2(\text{g})}] = 0.80 \text{ mol/L}$$

$$Q = \frac{[\text{HI}_{(\text{g})}]^2}{[\text{H}_{2(\text{g})}][\text{I}_{2(\text{g})}]}$$

$$Q = \frac{0.000}{0.80^2}$$

$$Q = 0.000$$

Since Q is less than K , the reaction will proceed to the right.

ICE Table for the Formation of $\text{HI}_{(\text{g})}$			
	$\text{H}_{2(\text{g})} +$	$\text{I}_{2(\text{g})} \rightleftharpoons$	$2 \text{HI}_{(\text{g})}$
Initial concentration (mol/L)	0.80	0.80	0.000
Change in concentration (mol/L)	$-x$	$-x$	$2x$
Equilibrium concentration (mol/L)	$0.80 - x$	$0.80 - x$	$2x$

At equilibrium...

$$Q = \frac{[\text{HI}_{(\text{g})}]^2}{[\text{H}_{2(\text{g})}][\text{I}_{2(\text{g})}]}$$

$$\frac{(2x)^2}{(0.80 - x)^2} = 46.0$$

$$\sqrt{\frac{(2x)^2}{(0.80 - x)^2}} = \sqrt{46.0}$$

$$\frac{2x}{0.80 - x} = 6.782 \quad (\text{extra digits carried})$$

$$2x = 6.782(0.80 - x)$$

$$x = 0.618$$

$$[\text{HI}_{(\text{g})}] = 2x$$

$$= 2(0.618)$$

$$[\text{HI}_{(\text{g})}] = 1.2 \text{ mol/L}$$

$$[\text{H}_{2(\text{g})}] = [\text{I}_{2(\text{g})}]$$

$$= 0.80 - x$$

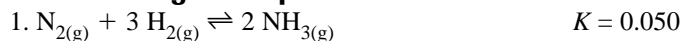
$$[\text{H}_{2(\text{g})}] = 0.18 \text{ mol/L}$$

The equilibrium concentrations of hydrogen, iodine, and hydrogen iodide are 0.18 mol/L, 0.18 mol/L, and 1.2 mol/L, respectively.

SECTION 7.5 QUESTIONS

(Page 481)

Understanding Concepts



$$[\text{N}_{2(\text{g})}] = 2.0 \times 10^{-4} \text{ mol/L}$$

$$[\text{H}_{2(\text{g})}] = 4.0 \times 10^{-3} \text{ mol/L}$$

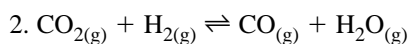
$$[\text{NH}_{3(\text{g})}] = 2.2 \times 10^{-3} \text{ mol/L}$$

$$Q = \frac{[\text{NH}_{3(\text{g})}]^2}{[\text{N}_{2(\text{g})}][\text{H}_{2(\text{g})}]^3}$$

$$= \frac{[2.2 \times 10^{-3}]^2}{[2.0 \times 10^{-4}][4.0 \times 10^{-3}]^3}$$

$$Q = 3.8 \times 10^5$$

Since the value of Q is larger than the value of K (0.050), the reaction is not at equilibrium. In order to attain equilibrium, the reaction will shift to the left (as written). The concentration of the reactants will increase and the concentration of the products will decrease.



$$[\text{H}_2\text{O}_{(\text{g})}]_{\text{initial}} = \frac{0.25 \text{ mol}}{1.00 \text{ L}}$$

$$[\text{H}_2\text{O}_{(\text{g})}]_{\text{initial}} = 0.25 \text{ mol/L}$$

$$[\text{CO}_{(\text{g})}]_{\text{initial}} = \frac{0.20 \text{ mol}}{1.00 \text{ L}}$$

$$[\text{CO}_{(\text{g})}]_{\text{initial}} = 0.20 \text{ mol/L}$$

$$\text{At equilibrium, } [\text{CO}_{2(\text{g})}] = \frac{0.10 \text{ mol}}{1.00 \text{ L}}$$

$$[\text{CO}_{2(\text{g})}] = 0.10 \text{ mol/L}$$

ICE Table for the Reaction of Carbon Dioxide and Hydrogen				
	$\text{CO}_{2(\text{g})}$	$+ \text{H}_{2(\text{g})} \rightleftharpoons$	$\text{CO}_{(\text{g})} +$	$\text{H}_2\text{O}_{(\text{g})}$
Initial concentration (mol/L)	0.00	0.00	0.20	0.25
Change in concentration (mol/L)	+0.10	+0.10	-0.10	-0.10
Equilibrium concentration (mol/L)	0.10	0.10	0.10	0.15

$$K = \frac{[\text{CO}_{(\text{g})}][\text{H}_2\text{O}_{(\text{g})}]}{[\text{CO}_{2(\text{g})}][\text{H}_{2(\text{g})}]}$$

$$= \frac{[0.10][0.15]}{[0.10][0.10]}$$

$$K = 1.5$$

3. (a)
$$K = \frac{[\text{H}_{2(\text{g})}][\text{Br}_{2(\text{g})}]}{[\text{HBr}_{(\text{g})}]^2}$$

$$[\text{H}_{2(\text{g})}]_{\text{initial}} = [\text{Br}_{2(\text{g})}]_{\text{initial}}$$

$$[\text{H}_{2(\text{g})}]_{\text{initial}} = \frac{0.25 \text{ mol}}{0.50 \text{ L}}$$

$$= 0.50 \text{ mol/L}$$

ICE Table for the Decomposition of Hydrogen Bromide			
	$2 \text{HBr}_{(\text{g})} \rightleftharpoons$	$\text{H}_{2(\text{g})} +$	$\text{Br}_{2(\text{g})}$
Initial concentration (mol/L)	0.00	0.50	0.50
Change in concentration (mol/L)	+2x	-x	-x
Equilibrium concentration (mol/L)	2x	0.50 - x	0.50 - x

$$K = \frac{[\text{H}_{2(\text{g})}][\text{Br}_{2(\text{g})}]}{[\text{HBr}_{(\text{g})}]^2}$$

$$\frac{[0.50 - x][0.50 - x]}{[2x]^2} = 0.020$$

$$\frac{[0.50 - x]^2}{[2x]^2} = 0.020$$

$$\sqrt{\frac{[0.50 - x]^2}{[2x]^2}} = \sqrt{0.020}$$

$$\frac{0.50 - x}{2x} = 0.1414 \quad (\text{extra digits carried})$$

$$0.2828x = 0.50 - x$$

$$x = 0.390$$

$$[\text{HBr}_{(\text{g})}] = 2x$$

$$= 2(0.390)$$

$$[\text{HBr}_{(\text{g})}] = 0.78 \text{ mol/L}$$

$$[\text{Br}_{2(\text{g})}] = [\text{H}_{2(\text{g})}]$$

$$= 0.50 - x$$

$$= 0.50 - 0.390$$

$$[\text{Br}_{2(\text{g})}] = 0.11 \text{ mol/L}$$

$$[\text{H}_{2(\text{g})}] = 0.11 \text{ mol/L}$$

The concentrations of hydrogen bromide, hydrogen, and bromine are 0.78 mol/L, 0.11 mol/L, and 0.11 mol/L, respectively.

$$(b) \quad n_{\text{HBr}} = (0.78 \text{ mol/L})(0.50 \text{ L})$$

$$n_{\text{HBr}} = 0.39 \text{ mol}$$

$$n_{\text{H}_2} = n_{\text{Br}_2}$$

$$= (0.11 \text{ mol/L})(0.50 \text{ L})$$

$$n_{\text{H}_2} = 0.055 \text{ mol}$$

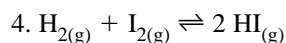
$$n_{\text{Br}_2} = 0.055 \text{ mol}$$

There is 0.39 mol of hydrogen bromide, 0.055 mol of hydrogen, and 0.55 mol of bromine present at equilibrium.

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

$$= \frac{0.39}{0.50} \times 100\%$$

$$\text{percent reaction} = 78\%$$



Initial concentrations are

$$K = \frac{[\text{H}_{2(\text{g})}][\text{I}_{2(\text{g})}]}{[\text{HI}_{(\text{g})}]^2}$$

$$K = \frac{[\text{HI}_{(\text{g})}]^2}{[\text{H}_{2(\text{g})}][\text{I}_{2(\text{g})}]}$$

$$[\text{H}_{2(\text{g})}]_{\text{initial}} = \frac{0.20 \text{ mol}}{0.500 \text{ L}}$$

$$[\text{H}_{2(\text{g})}]_{\text{initial}} = 0.40 \text{ mol/L}$$

$$[\text{I}_{2(\text{g})}]_{\text{initial}} = \frac{0.50 \text{ mol}}{0.500 \text{ L}}$$

$$[\text{I}_{2(\text{g})}]_{\text{initial}} = 1.0 \text{ mol/L}$$

$$[\text{HI}_{(\text{g})}]_{\text{initial}} = 0.0 \text{ mol/L}$$

$$Q = \frac{[\text{HI}_{(\text{g})}]^2}{[\text{H}_{2(\text{g})}][\text{I}_{2(\text{g})}]}$$

$$= \frac{0.000}{(0.40)(1.0)}$$

$$Q = 0.000$$

Since Q is less than K , the reaction will proceed to the right.

ICE Table for the Formation of $\text{HI}_{(\text{g})}$			
	$\text{H}_{2(\text{g})} +$	$\text{I}_{2(\text{g})} \rightleftharpoons$	$2 \text{HI}_{(\text{g})}$
Initial concentration (mol/L)	0.40	1.0	0.00
Change in concentration (mol/L)	$-x$	$-x$	$+2x$
Equilibrium concentration (mol/L)	$0.40 - x$	$1.0 - x$	$2x$

At equilibrium,

$$Q = \frac{[\text{HI}_{(\text{g})}]^2}{[\text{H}_{2(\text{g})}][\text{I}_{2(\text{g})}]}$$

$$\frac{(2x)^2}{(0.40 - x)(1.0 - x)} = 46$$

$$4x^2 = (46)(0.40 - x)(1 - x)$$

$$4x^2 = (46)(0.4 - 1.4x + x^2)$$

$$x^2 = (11.5)(0.40 - x)(1 - x)$$

$$x^2 = 4.6 - 16.1x + 11.5x^2$$

$$10.5x^2 - 16.1x + 4.6 = 0$$

$$x = \frac{16.1 \pm \sqrt{16.1^2 - 4(10.5)(4.6)}}{2(10.5)}$$

$$x = 0.3798 \text{ or } 1.15 \quad \text{extra digits carried}$$

The root 1.15 is rejected as it exceeds the initial amount of hydrogen.

$$K = \frac{[\text{HI}_{(\text{g})}]^2}{[\text{H}_{2(\text{g})}][\text{I}_{2(\text{g})}]}$$

$$[\text{HI}_{(\text{g})}] = 2x$$

$$[\text{HI}_{(\text{g})}] = 0.7596 \text{ mol/L}$$

$$[\text{H}_{2(\text{g})}] = 0.40 - x$$

$$= 0.40 - 0.3798$$

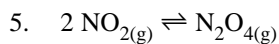
$$[\text{H}_{2(\text{g})}] = 0.020$$

$$[\text{I}_{2(\text{g})}] = 1.0 - x$$

$$= 1.0 - 0.3798$$

$$[\text{I}_{2(\text{g})}] = 0.62 \text{ mol/L}$$

The equilibrium concentrations of hydrogen, iodine, and hydrogen iodide are 0.020 mol/L, 0.62 mol/L, and 0.76 mol/L, respectively.



Initial concentrations are

$$[\text{N}_2\text{O}_{4(g)}] = \frac{4.00 \text{ mol}}{1.00 \text{ L}}$$

$$[\text{N}_2\text{O}_{4(g)}] = 4.00 \text{ mol/L}$$

$$[\text{NO}_{2(g)}] = 0.00 \text{ mol/L}$$

ICE Table for the Formation of $\text{N}_2\text{O}_{4(g)}$		
	$2 \text{NO}_{2(g)} \rightleftharpoons$	$\text{N}_2\text{O}_{4(g)}$
Initial concentration (mol/L)	0.00	4.00
Change in concentration (mol/L)	+2x	-x
Equilibrium concentration (mol/L)	2x	4.00 - x

At equilibrium,

$$K = \frac{[\text{N}_2\text{O}_{4(g)}]}{[\text{NO}_{2(g)}]^2}$$

$$K = 1.15$$

$$\frac{4.00 - x}{(2x)^2} = 1.15$$

$$\frac{4.00 - x}{4x^2} = 1.15$$

$$4 - x = 4.6x^2$$

$$4.6x^2 + x - 4 = 0$$

$$x = \frac{-1 \pm \sqrt{1^2 - 4(4.6)(-4)}}{2(4.6)}$$

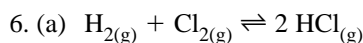
$$x = 0.830 \text{ or } -1.05$$

The negative root is rejected because negative concentrations are impossible.

$$[\text{NO}_{2(g)}] = 2x$$

$$[\text{NO}_{2(g)}] = 1.66 \text{ mol/L}$$

The equilibrium concentration of $\text{NO}_{2(g)}$ is 1.66 mol/L.



Initial concentrations are

$$[\text{HCl}_{(g)}] = 0.00$$

$$[\text{HCl}_{(g)}] = [\text{Cl}_{2(g)}]$$

$$= \frac{1.50 \text{ mol}}{0.750 \text{ L}}$$

$$[\text{H}_{2(g)}] = 2.00 \text{ mol/L}$$

$$[\text{Cl}_{2(g)}] = 2.00 \text{ mol/L}$$

ICE Table for the Formation of $\text{HCl}_{(g)}$			
	$\text{H}_{2(g)} +$	$\text{Cl}_{2(g)} \rightleftharpoons$	$2 \text{HCl}_{(g)}$
Initial concentration (mol/L)	2.00	2.00	0.000
Change in concentration (mol/L)	-x	-x	2x
Equilibrium concentration (mol/L)	2.00 - x	2.00 - x	2x

At equilibrium,

$$K = \frac{[\text{HCl}_{(\text{g})}]^2}{[\text{H}_{2(\text{g})}][\text{Cl}_{2(\text{g})}]}$$

$$\frac{(2x)^2}{(2.00 - x)^2} = 4.4 \times 10^{-2}$$

$$\frac{4x^2}{(2.00 - x)^2} = 4.4 \times 10^{-2}$$

$$\sqrt{\frac{4x^2}{(2.00 - x)^2}} = \sqrt{4.4 \times 10^{-2}}$$

$$\frac{2x}{2.00 - x} = 0.21$$

$$x = 0.19$$

$$[\text{HCl}_{(\text{g})}] = 2x$$

$$[\text{HCl}_{(\text{g})}] = 0.38 \text{ mol/L}$$

$$[\text{H}_{2(\text{g})}] = [\text{Cl}_{2(\text{g})}]$$

$$= 2.00 - x$$

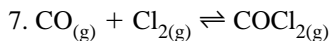
$$[\text{H}_{2(\text{g})}] = 1.81 \text{ mol/L}$$

$$[\text{Cl}_{2(\text{g})}] = 1.81 \text{ mol/L}$$

The equilibrium concentrations of hydrogen, chlorine, and hydrogen chloride are 1.81 mol/L, 1.81 mol/L, and 0.38 mol/L, respectively.

- (b) $n_{\text{H}_2} = (1.81 \text{ mol/L})(0.750 \text{ L})$
 $n_{\text{H}_2} = 1.36 \text{ mol}$
 $n_{\text{Cl}_2} = n_{\text{H}_2}$
 $n_{\text{Cl}_2} = 1.36 \text{ mol}$
 $n_{\text{HCl}} = (0.38 \text{ mol/L})(0.750 \text{ L})$
 $n_{\text{HCl}} = 0.285 \text{ mol/L}$
- (c) $\text{percent reaction} = \frac{0.285}{3.00} \times 100\%$
 $= 9.50\%$

The reaction extent is 9.50%.



Initial concentrations are

$$[\text{COCl}_{2(\text{g})}] = \frac{2.5 \text{ mol}}{10.00 \text{ L}}$$

$$= 0.25 \text{ mol/L}$$

ICE Table for the Formation of $\text{COCl}_{2(\text{g})}$			
	$\text{CO}_{(\text{g})} +$	$\text{Cl}_{2(\text{g})} \rightleftharpoons$	$\text{COCl}_{2(\text{g})}$
Initial concentration (mol/L)	0.00	0.00	0.25
Change in concentration (mol/L)	+x	+x	-x
Equilibrium concentration (mol/L)	x	x	0.25 - x

At equilibrium,

$$K = \frac{[\text{COCl}_{2(g)}]}{[\text{CO}_{(g)}][\text{Cl}_{2(g)}]}$$

$$= 8.2 \times 10^{-2}$$

$$\frac{0.25 - x}{x^2} = 8.2 \times 10^{-2}$$

$$8.2 \times 10^{-2} x^2 = 0.25 - x$$

$$8.2 \times 10^{-2} x^2 + x - 0.25 = 0$$

$$x = \frac{-1 \pm \sqrt{1^2 - 4(8.2 \times 10^{-2})(-0.25)}}{2(8.2 \times 10^{-2})}$$

$$= 0.25 \text{ or } -12.4$$

The negative root is rejected because negative concentrations are impossible.

$$[\text{CO}_{(g)}] = [\text{Cl}_{2(g)}] = x$$

$$= 0.25 \text{ mol/L}$$

The equilibrium concentrations of carbon monoxide and chlorine are 0.25 mol/L.



Initial concentrations are

$$[\text{PCl}_{3(g)}] = \frac{0.500 \text{ mol}}{1.00 \text{ L}}$$

$$[\text{PCl}_{3(g)}] = 0.500 \text{ mol/L}$$

$$[\text{Cl}_{2(g)}] = \frac{0.500 \text{ mol}}{1.00 \text{ L}}$$

$$[\text{Cl}_{2(g)}] = 0.500 \text{ mol/L}$$

ICE Table for the Decomposition of $\text{PCl}_{5(g)}$			
	$\text{PCl}_{5(g)} \rightleftharpoons$	$\text{PCl}_{3(g)} +$	$\text{Cl}_{2(g)}$
Initial concentration (mol/L)	0.00	0.500	0.500
Change in concentration (mol/L)	x	$-x$	$-x$
Equilibrium concentration (mol/L)	x	$0.500 - x$	$0.500 - x$

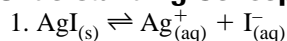
The equilibrium concentrations of phosphorus pentachloride, phosphorus trichloride, and chlorine are 0.0185 mol/L, 0.482 mol/L, and 0.482 mol/L, respectively.

7.6 THE SOLUBILITY PRODUCT CONSTANT

PRACTICE

(Page 486)

Understanding Concepts



$$K_{sp} = [\text{Ag}^+][\text{I}^-]$$

$$K_{sp} = 1.5 \times 10^{-16}$$