

7.5 on next page

PRACTICE

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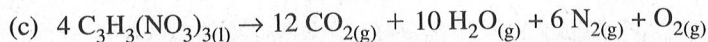
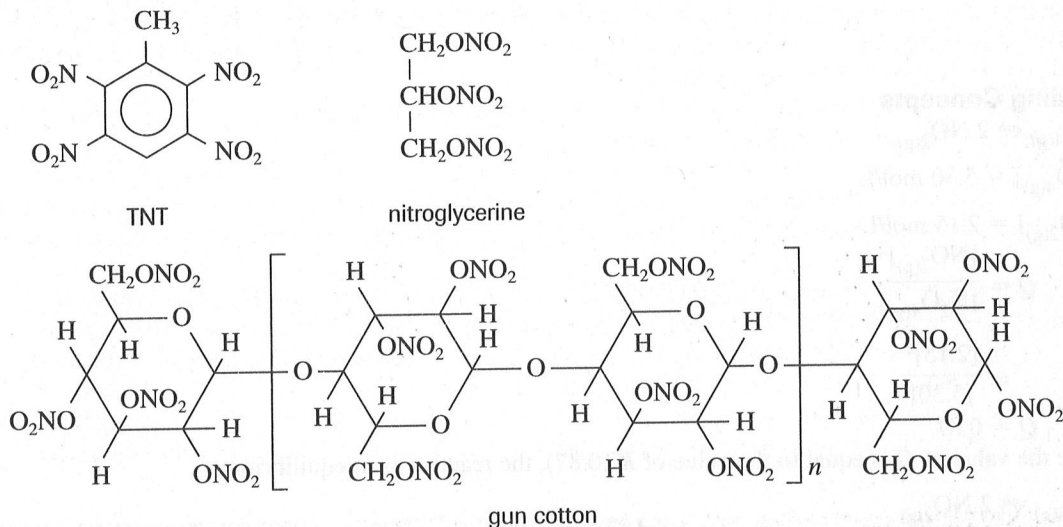
Understanding Concepts

- Lowering the temperature would have the effect of slowing down the reaction.
- The catalyst does not affect the position of the equilibrium since it lowers the activation energy for both the forward and reverse reactions. However, using a catalyst allows the manufacturer to run the process at a lower temperature and still maintain yield.

Making Connections

5. (a) TNT, nitroglycerine, and gun cotton

(b)



This reaction is explosive because it is highly exothermic ($\Delta H = 1415 \text{ kJ/mol}$ nitroglycerine) and converts four moles of liquid nitroglycerine into 29 moles of gaseous products.

- (d) Gun cotton (nitrocellulose) is a highly explosive nitrated cellulose made by soaking cotton (cellulose) in nitric acid and sulfuric acid; used to make smokeless gunpowder.

SECTION 7.4 QUESTIONS

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Making Connections

- (a) Decrease the temperature, increase the pressure, increase $[\text{N}_2(g)]$, increase $[\text{H}_2(g)]$, and remove ammonia from the system.

(b) Decreasing the temperature will shift the equilibrium to the right but makes the reaction uneconomical because the yield of ammonia decreases.

(c) The heat released by the forward reaction can be used to help maintain the temperature of the reaction chamber. The company should also ensure that the equipment is well maintained so that gas leaks in the system do not reduce pressure. They should keep the system insulated to prevent heat loss, and they should use high-efficiency pumps and heaters to save on energy costs. The system should separate and recycle unreacted $\text{H}_2(g)$ and $\text{N}_2(g)$.
- (a) Possible sources of these gases are: nitrogen—extraction from the atmosphere, and hydrogen—electrolysis of water.

(b) Hydrogen is produced by reacting methane ($\text{CH}_4(g)$, natural gas) with steam ($\text{H}_2\text{O}(g)$), producing carbon dioxide ($\text{CO}_2(g)$) and hydrogen ($\text{H}_2(g)$). Nitrogen ($\text{N}_2(g)$) is obtained by the fractional distillation of air (~78% $\text{N}_2(g)$). Fractional distillation is a process by which the components in a chemical mixture are separated according to their different boiling points. Vapours from a boiling solution are passed along a cooled column. The temperature of the column gradually decreases along its length. Components with a higher boiling point condense on the column and return to the solution; components with a lower boiling point pass through the column, are condensed, and are collected in a suitable collecting vessel.
- (a) Bacteria such as *Rhizobium* and *Azotobacter* absorb nitrogen from air and fix it into ammonia through a series of reactions that use an enzyme complex called nitrogenase as a catalyst.

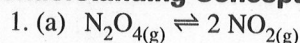
- (b) More nitrogen is fixed naturally (1.75×10^8 tonnes per year) than synthetically (8×10^7 tonnes per year).
- (c) Studies show that high levels of nitrogen fixation caused by nitrogen pollution are likely to cause severe declines in native plant communities by favouring weedy plants that favour high 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

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Understanding Concepts



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

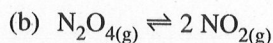
$$[\text{NO}_{2(g)}] = 2.15 \text{ mol/L}$$

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

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

$$Q = 0.87$$

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



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

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

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

$$= \frac{[1.55]^2}{[0.80]}$$

$$Q = 3.0$$

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.

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Understanding Concepts

$$3. \quad K = \frac{[\text{NH}_3(\text{g})]^2}{[\text{N}_2(\text{g})][\text{H}_2(\text{g})]^3}$$

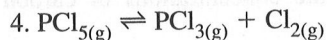
$$K = 626$$

$$[\text{N}_2(\text{g})] = \frac{[\text{NH}_3(\text{g})]^2}{626[\text{H}_2(\text{g})]^3}$$

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

$$[\text{N}_2(\text{g})] = 2.7 \times 10^{-3} \text{ mol/L}$$

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



$$K = \frac{[\text{PCl}_3(\text{g})][\text{Cl}_2(\text{g})]}{[\text{PCl}_5(\text{g})]}$$

$$K = 32$$

$$[\text{PCl}_5(\text{g})] = \frac{[\text{PCl}_3(\text{g})][\text{Cl}_2(\text{g})]}{32}$$

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

$$[\text{PCl}_5(\text{g})] = 1.80 \times 10^{-2} \text{ mol/L}$$

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

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Understanding Concepts

5. Initial concentrations...

$$[\text{CO}_2(\text{g})] = [\text{H}_2(\text{g})]$$

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

$$[\text{CO}_2(\text{g})] = 0.100 \text{ mol/L}$$

ICE Table for the Reaction of $\text{CO}_2(\text{g})$ with $\text{H}_2(\text{g})$				
	$\text{CO}_2(\text{g}) +$	$\text{H}_2(\text{g}) \rightleftharpoons$	$\text{CO}(\text{g}) +$	$\text{H}_2\text{O}(\text{g})$
Initial concentration (mol/L)	0.100	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

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

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

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

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

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

$$0.12649 - 1.2649x = x$$

$$0.12649 = 2.2649x$$

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

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

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

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

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...

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

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

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

ICE Table for the Reaction of $\text{I}_{2(g)}$ with $\text{Cl}_{2(g)}$			
	$\text{I}_{2(g)}$ +	$\text{Cl}_{2(g)} \rightleftharpoons$	$2 \text{ICl}_{(g)}$
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

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

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

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

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

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

$$2.2625 - 9.05x = 2x$$

$$2.2625 = 11.05x \text{ (extra digits carried)}$$

2.2625 = 11.05x			
x = 0.2048			
$[\text{I}_{2(g)}] = [\text{Cl}_{2(g)}]$			

$$= 0.25 - 0.2251$$

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

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

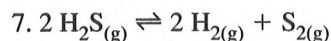
$$[\text{ICl}_{(\text{g})}] = 0.41 \text{ mol/L}$$

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.

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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...

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

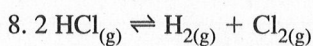
$$= 3.48\%$$

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			
	$2 \text{HCl}_{(g)} \rightleftharpoons$	$\text{H}_{2(g)} +$	$\text{Cl}_{2(g)}$
Initial concentration (mol/L)	2.00		
Change in concentration (mol/L)	-2x	+x	+x
Equilibrium concentration (mol/L)	$2.00 - 2x$	x	x

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

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

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

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

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

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

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

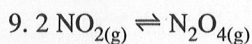
$$x \approx 3.58 \times 10^{-17}$$

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

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Understanding Concepts



Initial concentrations are

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

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

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

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

$$Q = 0.000$$

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

ICE Table for the Formation of $\text{N}_2\text{O}_{4(g)}$		
	$2 \text{NO}_{2(g)} \rightleftharpoons$	$\text{N}_2\text{O}_{4(g)}$
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...

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

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

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

$$(0.65 - 2x)(0.65 - 2x) = 0.4225 - 1.3x - 1.3x + 4x^2$$

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

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

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

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

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

$$x = 0.1465$$

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

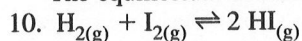
$$= 0.147 \text{ mol/L}$$

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

$$= 0.65 - 2(0.1465)$$

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

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



Initial concentrations are

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

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

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

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

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

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

$$Q = 0.000$$

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

ICE Table for the Formation of $\text{HI}_{(g)}$			
	$\text{H}_{2(g)}$	$\text{I}_{2(g)}$	$2 \text{HI}_{(g)}$
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...

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

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

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

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

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

$$x = 0.618$$

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

$$= 2(0.618)$$

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

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

$$= 0.80 - x$$

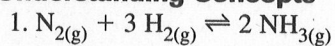
$$[\text{H}_{2(g)}] = 0.18 \text{ mol/L}$$

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

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Understanding Concepts



$$K = 0.050$$

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

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

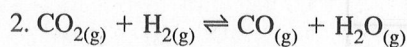
$$[\text{NH}_{3(g)}] = 2.2 \times 10^{-4} \text{ mol/L}$$

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

$$= \frac{[2.2 \times 10^{-4}]^2}{[2.0 \times 10^{-4}][4.0 \times 10^{-3}]^3} = \frac{4.84 \times 10^{-8}}{(2 \times 10^{-4})(64 \times 10^{-9})} = \frac{4.84 \times 10^{-8}}{128 \times 10^{-13}} = 0.0378 \times 10^5 = 3.78 \times 10^3$$

$$Q = 3.8 \times 10^3$$

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.



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

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

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

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

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

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

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

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

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

$$K = 1.5$$

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

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

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

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

ICE Table for the Decomposition of Hydrogen Bromide			
	$2 \text{HBr}_{(g)} \rightleftharpoons$	$\text{H}_{2(g)} +$	$\text{Br}_{2(g)}$
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

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

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

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

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

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

$$0.2828x = 0.50 - x$$

$$x = 0.390$$

$$\begin{aligned} [\text{HBr}_{(g)}] &= 2x \\ &= 2(0.390) \end{aligned}$$

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

$$\begin{aligned} [\text{Br}_{2(g)}] &= [\text{H}_{2(g)}] \\ &= 0.50 - x \end{aligned}$$

$$= 0.50 - 0.390$$

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

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

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

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

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

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

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

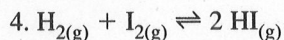
$$n_{\text{H}_2} = 0.055 \text{ mol}$$

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

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

$$\begin{aligned} (c) \text{ percent reaction} &= \frac{\text{actual product yield}}{\text{theoretical product yield}} \times 100\% \\ &= \frac{0.39}{0.50} \times 100\% \end{aligned}$$

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



Initial concentrations are

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

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

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

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

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

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

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

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

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

$$Q = 0.000$$

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

ICE Table for the Formation of $\text{HI}_{(g)}$			
	$\text{H}_{2(g)} +$	$\text{I}_{2(g)} \rightleftharpoons$	$2 \text{HI}_{(g)}$
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,

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

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

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

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

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

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

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

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

$$x = 0.3798 \text{ or } 1.15$$

extra digits carried

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

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

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

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

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

$$= 0.40 - 0.3798$$

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

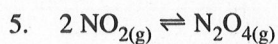
$$[\text{I}_{2(g)}] = 1.0 - x$$

$$= 1.0 - 0.3798$$

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

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

Doesn't match back of book.



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,

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

$$K = 1.15$$

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

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

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

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

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

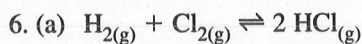
$$x = 0.830 \text{ or } -1.05$$

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,

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

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

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

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

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

$$x = 0.19$$

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

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

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

$$= 2.00 - x$$

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

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

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) $n_{\text{H}_2} = (1.81 \text{ mol/L})(0.750 \text{ L})$

$$n_{\text{H}_2} = 1.36 \text{ mol}$$

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

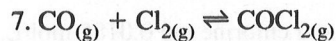
$$n_{\text{Cl}_2} = 1.36 \text{ mol}$$

$$n_{\text{HCl}} = (0.38 \text{ mol/L})(0.750 \text{ L})$$

$$n_{\text{HCl}} = 0.285 \text{ mol/L}$$

(c) percent reaction = $\frac{0.285}{3.00} \times 100\%$
 $= 9.50\%$

The reaction extent is 9.50%.



Initial concentrations are

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

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

ICE Table for the Formation of $\text{COCl}_{2(g)}$			
	$\text{CO}_{(g)}$	$\text{Cl}_{2(g)}$	$\text{COCl}_{2(g)}$
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,

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

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

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

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

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

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

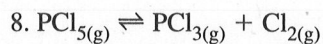
$$= 0.25 \text{ or } -12.4$$

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

full solution not shown. →

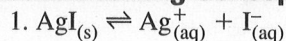
The equilibrium concentrations of phosphorus pentachloride, phosphorus trichloride, and chlorine are 0.0185 mol/L, 0.482 mol/L, and 0.482 mol/L, respectively.

7.6 THE SOLUBILITY PRODUCT CONSTANT

PRACTICE

(Page 486)

Understanding Concepts



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

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