Section 7.5: Quantitative Changes in Equilibrium Systems
Tutorial 1 Practice, page 452
1. (a) Given: $[HI(g)] = 0.14$ mol/L; $[H_2(g)] = 0.040$ mol/L; $[I_2(g)] = 0.010$ mol/L;
$K = 0.020$
Required: Q; direction of reaction to reach equilibrium

Solution: $Q = \dfrac{[H_2(g)][I_2(g)]}{[HI(g)]^2}$

$= \dfrac{(0.040)(0.010)}{(0.14)^2}$

$Q = 0.020$

Statement: The value of Q is 0.020. Q is equal to K, so the reaction is at equilibrium.
(b) Given: $[HI(g)] = 0.20$ mol/L; $[H_2(g)] = 0.15$ mol/L; $[I_2(g)] = 0.090$ mol/L; $K = 0.020$
Required: Q; direction of reaction to reach equilibrium

Solution: $Q = \dfrac{[H_2(g)][I_2(g)]}{[HI(g)]^2}$

$= \dfrac{(0.15)(0.090)}{(0.20)^2}$

$Q = 0.34$

Statement: The value of Q is 0.34. Q is greater than K, so the product concentrations are greater than at equilibrium; the reaction will shift toward the left, more HI.
2. Time 1:
Given: $[N_2O_4(g)] = 0.80$ mol/L; $[NO_2(g)] = 1.55$ mol/L; $K = 0.87$
Required: Q; direction of reaction to reach equilibrium

Solution: $N_2O_4(g) \rightleftharpoons 2\,NO_2(g)$

$Q = \dfrac{[NO_2(g)]^2}{[N_2O_4(g)]}$

$= \dfrac{(1.55)^2}{0.80}$

$Q = 3.0$

Statement: The value of Q is 3.0. Q is greater than K, so the product concentrations are greater than at equilibrium; the reaction will shift toward the left, more N_2O_4.
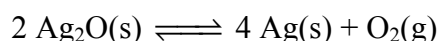
Time 2:
Given: $[N_2O_4(g)] = 1.66$ mol/L; $[NO_2(g)] = 1.20$ mol/L; $K = 0.87$
Required: Q; direction of reaction to reach equilibrium

Solution: $Q = \dfrac{[NO_2(g)]^2}{[N_2O_4(g)]}$

$= \dfrac{(1.20)^2}{1.66}$

$Q = 0.867$

Statement: The value of Q is 0.867. Q is approximately equal to K, so the reaction is at equilibrium.

3. (a) The equilibrium constant equation for the reaction represented by

$$2\ Ag_2O(s) \rightleftharpoons 4\ Ag(s) + O_2(g)$$

is $K = [O_2(g)]$, because the other substances are solids.

(b) Given: $K = 2.5 \times 10^{-3}$; $[O_2(g)]_{instantaneous} = 5.0 \times 10^{-2}$

Required: Q; direction of reaction to reach equilibrium

Solution: $Q = [O_2(g)]$

$$Q = 5.0 \times 10^{-2}$$

Statement: The value of Q is greater than K, so the reaction will shift to the left, toward the reactant.

Tutorial 2 Practice, page 454

1. Given: Volume, $V = 250$ mL $= 0.25$ L; $n_{initial\ I_2\ (g)} = 0.50$ mol; $n_{initial\ Br_2\ (g)} = 0.50$ mol; $K = 1.2 \times 10^2$

Required: $[I_2(g)]_{equilibrium}$; $[Br_2(g)]_{equilibrium}$

Analysis: $c = \dfrac{n}{V}$

Solution:

Step 1. Calculate concentrations, c, for in mol/L from the given amounts of all entities. Calculate $[I_2(g)]_{initial}$:

$$c = \dfrac{n_{initial}}{V}$$
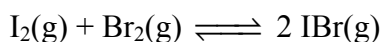
$$= \dfrac{0.50\ \text{mol}}{0.25\ \text{L}}$$

$$c = 2.0\ \text{mol/L}$$

Using the same formula,

$[Br_2(g)]_{initial} = 2.0$ mol/L

Since there is no iodine monobromide gas initially, $[IBr(g)]_{initial}$ is 0.0 mol/L.

Step 2. Write the balanced equation for the equilibrium reaction system.

$$I_2(g) + Br_2(g) \rightleftharpoons 2\ IBr(g)$$

Step 3. Determine the equilibrium law equation.

$$K = \dfrac{[IBr(g)]^2}{[I_2(g)][Br_2(g)]}$$

Step 4. Use an ICE table to determine the relationship between the equilibrium concentrations of the reactant and the products.

	$I_2(g)$ +	$Br_2(g)$ $\rightleftharpoons$	2 $IBr(g)$
I	2.0	2.0	0.0
C	$-x$	$-x$	$+2x$
E	$2.0 - x$	$2.0 - x$	$2x$

Step 5. Calculate the value of x by substituting the given value of K and the concentrations from the ICE table into the equilibrium constant equation.

$$K = \frac{[\text{IBr(g)}]^2}{[\text{I}_2(\text{g})][\text{Br}_2(\text{g})]}$$

$$1.2 \times 10^2 = \frac{(2x)^2}{(2.0-x)(2.0-x)}$$

$$\sqrt{1.2 \times 10^2} = \frac{\sqrt{(2x)^2}}{\sqrt{(2.0-x)(2.0-x)}}$$

$$\pm 10.95 = \frac{2x}{(2.0-x)}$$

$$(\pm 10.95)(2.0-x) = 2x$$

$$\pm 21.9 - 10.95x = 2x$$

$$\pm 21.9 = 12.95x$$

$$x = \pm 1.69$$

Substituting the negative value of x into the expressions for equilibrium concentrations in the ICE table results in a negative concentration value for $\text{IBr}_2(\text{g})$, which is not possible in the real world. Therefore, use $x = 1.69$ mol/L in the equilibrium concentration calculations.

Step 6. Calculate the equilibrium concentrations.

$[\text{I}_2(\text{g})]_{\text{equilibrium}}$ and $[\text{Br}_2(\text{g})]_{\text{equilibrium}} = 2.0 - x$

$$= 2.0 - 1.69$$

$[\text{I}_2(\text{g})]_{\text{equilibrium}}$ and $[\text{Br}_2(\text{g})]_{\text{equilibrium}} = 0.3$ mol/L

Statement: The equilibrium concentration for both iodine gas and bromine gas is 0.3 mol/L.

2. Given: Volume, $V = 2.00$ L; $n_{\text{initial H}_2\text{ (g)}} = 4.00$ mol; $n_{\text{initial I}_2\text{ (g)}} = 1.99$ mol; $K = 49.7$;

$[\text{H}_2(\text{g})]_{\text{equilibrium}} = 1.07$ mol/L

Required: $[\text{I}_2(\text{g})]_{\text{equilibrium}}$; $[\text{HI(g)}]_{\text{equilibrium}}$

Analysis: $c = \dfrac{n}{V}$

Solution:

Step 1. Convert the given amounts to concentrations in mol/L by dividing by the volume.

Calculate $[\text{H}_2(\text{g})]_{\text{initial}}$:   Calculate $[\text{H}_2(\text{g})]_{\text{initial}}$:

$$c = \frac{n_{\text{initial}}}{V} \qquad\qquad\qquad c = \frac{n_{\text{initial}}}{V}$$

$$= \frac{4.00 \text{ mol}}{2.00 \text{ L}} \qquad\qquad\qquad = \frac{1.99 \text{ mol}}{2.00 \text{ L}}$$

$c = 2.00$ mol/L $c = 0.995$ mol/L

$[\text{H}_2(\text{g})]_{\text{initial}} = 2.00$ mol/L $[\text{I}_2(\text{g})]_{\text{initial}} = 0.995$ mol/L

Since there is no hydrogen iodide gas initially, $[\text{HI(g)}]_{\text{initial}}$ is 0.00 mol/L.

Step 2. Use an ICE table to determine the relationship between the equilibrium concentrations of the reactant and the products.

	$H_2(g)$ +	$I_2(g)$	$\rightleftharpoons$ 2 HI(g)
I	2.00	0.995	0.00
C	$-x$	$-x$	$+2x$
E	$2.00 - x$	$0.995 - x$	$2x$

Step 3. Use $[H_2(g)]_{equilibrium}$ and $[H_2(g)]_{initial}$ to calculate the value of x.
$$2.00 - x = 1.07$$

$$2.00 - 1.07 = x$$

$$x = 0.93 \text{ mol/L}$$

Step 4. Calculate the equilibrium concentrations.
$$[I_2(g)]_{equilibrium} = 0.995 - x$$

$$= 0.995 - 0.93$$

$$[I_2(g)]_{equilibrium} = 0.065 \text{ mol/L}$$

$$HI(g)]_{equilibrium} = 2x$$

$$= 2(0.93)$$

$$HI(g)]_{equilibrium} = 1.86 \text{ mol/L}$$

Statement: The equilibrium concentration for hydrogen gas is 1.07 mol/L, the equilibrium concentration for iodine gas is 0.065 mol/L, and the equilibrium concentration for hydrogen iodide gas is 1.86 mol/L.

3. Given: $CO_2(g) + C(s) \rightleftharpoons 2\ CO(g)$; Volume, $V = 0.500$ L; $n_{initial\ CO_2\ (g)} = 0.250$ mol;

$[CO(g)]_{equilibrium} = 0.0157$ mol/L
Required: $[CO_2(g)]_{equilibrium}$; K
Solution:
Step 1. Determine the equilibrium law equation.
$$K = \frac{[CO(g)]^2}{[CO_2(g)]}$$
Step 2. Convert the given initial amount of carbon dioxide gas to concentration in mol/L by dividing by the volume.
$$c = \frac{n_{initial}}{V}$$

$$= \frac{0.250 \text{ mol}}{0.500 \text{ L}}$$

$$c = 0.500 \text{ mol/L}$$
$[CO_2(g)]_{initial} = 0.500$ mol/L
Since there is no carbon monoxide gas initially, $[CO(g)]_{initial}$ is 0.000 mol/L.

Step 3. Use an ICE table to determine the relationship between the equilibrium concentrations of the carbon dioxide gas and carbon monoxide gas.

	$CO_2(g)$ + $C(s)$ $\rightleftharpoons$ 2 $CO(g)$	
I	0.500	0.00
C	$-x$	$+2x$
E	$0.500 - x$	$2x$

Step 4. Use $[CO(g)]_{equilibrium}$ and $[CO(g)]_{initial}$ to calculate the value of x.
$$2x = 0.0157$$
$$x = 0.0078 \text{ mol/L}$$

Step 5: Calculate the equilibrium concentration of carbon dioxide gas.
$$[CO_2(g)]_{equilibrium} = 0.500 - x$$
$$= 0.500 - 0.0078$$
$$[CO_2(g)]_{equilibrium} = 0.492 \text{ mol/L}$$

Step 6: Determine the equilibrium constant, K.
$$K = \frac{[CO(g)]^2}{[CO_2(g)]}$$
$$= \frac{(0.0157)^2}{0.492}$$
$$K = 5.01 \times 10^{-4}$$

Statement: The concentration of carbon dioxide gas at equilibrium is 0.492 mol/L and the equilibrium constant, K, is 5.01×10^{-4}.

Tutorial 3 Practice, page 458
1. Given: $V = 0.500$ L; $n_{cyclopropane} = 2.50$ mol; $K = 5.6$

Required: $[cyclopropane(g)]_{equilibrium}$; $[propene(g)]_{equilibrium}$
Solution:
Step 1. Convert the given initial amount of cyclopropane gas to concentration in mol/L by dividing by the volume.
$$c = \frac{n_{initial}}{V}$$
$$= \frac{2.50 \text{ mol}}{0.500 \text{ L}}$$
$$c = 5.00 \text{ mol/L}$$
$[cyclopropane(g)]_{initial} = 5.00$ mol/L
Since there is no propene gas initially, $[propene(g)]_{initial}$ is 0.000 mol/L.
Step 2. Since initially there are no products, Q is 0, which is smaller than K. The equilibrium system will shift to the right, in the direction of products.

Step 3. Use an ICE table to determine the equilibrium concentrations.

	cyclopropane(g) $\rightleftharpoons$ propene(g)	
I	5.00	0.00
C	$-x$	$+x$
E	$5.00 - x$	x

Step 4. Substitute the equilibrium concentrations into the equilibrium constant equation, and solve for the unknown.

$$K = \frac{[\text{propene(g)}]}{[\text{cyclopropane(g)}]}$$

$$5.6 = \frac{x}{(5.00 - x)}$$

$$5.6(5.00 - x) = x$$

$$28 - 5.6x = x$$

$$-6.6x = -28$$

$$x = 4.24 \text{ mol/L}$$

Step 5: Calculate the equilibrium concentrations.

$$[\text{cyclopropane(g)}]_{\text{equilibrium}} = 5.00 - x$$

$$= 5.00 - 4.24$$

$$[\text{cyclopropane(g)}]_{\text{equilibrium}} = 0.76 \text{ mol/L}$$

$$[\text{propene (g)}]_{\text{equilibrium}} = x, \text{ or } 4.24 \text{ mol/L}$$

Statement: The concentration of cyclopropane gas at equilibrium is 0.76 mol/L and the concentration of propene gas at equilibrium is 4.24 mol/L.

2. Given: $V = 1.00$ L; $[\text{H}_2\text{(g)}]_{\text{initial}} = 0.200$ mol/L; $[\text{I}_2\text{(g)}]_{\text{initial}} = 0.200$ mol/L; $[\text{HI(g)}]_{\text{initial}} = 0.00$ mol/L; $K = 49.5$

Required: $[\text{HI(g)}]_{\text{equilibrium}}$; $[\text{H}_2\text{(g)}]_{\text{equilibrium}}$; $[\text{I}_2\text{(g)}]_{\text{equilibrium}}$

Solution:

Step 1. Since initially there are no products, Q is 0, which is smaller than K. The equilibrium system will shift to the right, in the direction of products.

Step 2. Using the balanced chemical equation, construct an ICE table for calculating equilibrium concentrations.

	$\text{H}_2\text{(g)}$ +	$\text{I}_2\text{(g)} \rightleftharpoons$	2 HI(g)
I	0.200	0.200	0.000
C	$-x$	$-x$	$+2x$
E	$0.200 - x$	$0.200 - x$	$2x$

Step 3. Substitute the expressions from the ICE table into the equilibrium constant equation and solve the equation. The equilibrium law for this reaction is

$$49.5 = \frac{[HI(g)]^2}{[H_2(g)][I_2(g)]}$$

$$49.5 = \frac{(2x)^2}{(0.200 - x)(0.200 - x)}$$

$$\sqrt{49.5} = \sqrt{\frac{(2x)^2}{(0.200 - x)(0.200 - x)}}$$

$$\pm 7.04 = \frac{2x}{0.200 - x}$$

$$\pm 7.04(0.200 - x) = 2x$$

$$\pm 1.407 - 7.04x = 2x$$

$$x = \pm 0.156 \text{ mol/L}$$

Substituting the negative value of x into the expressions for equilibrium concentrations in the ICE table results in a negative concentration value for HI(g), which is not possible in the real world. Therefore, use $x = 0.156$ mol/L in the equilibrium concentration calculations.

Step 4. Calculate the equilibrium concentrations.

$$[H_2(g)]_{equilibrium} = 0.200 - x$$

$$= 0.200 - 0.156$$

$$[H_2(g)]_{equilibrium} = 0.044 \text{ mol/L}$$

$$[I_2(g)]_{equilibrium} = [H_2(g)]_{equilibrium} = 0.044 \text{ mol/L}$$

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

$$= 2(0.156)$$

$$[HI(g)]_{equilibrium} = 0.312 \text{ mol/L}$$

Statement: The equilibrium concentration for hydrogen iodide gas is 0.312 mol/L, and the equilibrium concentration for both hydrogen gas and iodine gas is 0.044 mol/L.

3. Given: $2 \, HCl(g) \rightleftharpoons H_2(g) + Cl_2(g)$; $V = 250.0 \, mL = 0.2500 \, L$;
$n_{HCl(g)} = 0.500 \, mol$; $K = 3.2 \times 10^{-34}$

Required: $[H_2(g)]_{equilibrium}$; $[Cl_2(g)]_{equilibrium}$; $[HCl(g)]_{equilibrium}$

Solution:

Step 1. Convert the given initial amount of hydrogen chloride gas to concentration in mol/L by dividing by the volume.

$$c = \frac{n_{initial}}{V}$$

$$= \frac{0.500 \, mol}{0.2500 \, L}$$

$c = 2.00 \, mol/L$

$[HCl(g)]_{initial} = 2.00 \, mol/L$

Since there is no hydrogen gas or chlorine gas initially,

$[H_2(g)]_{initial} = 0.00 \, mol/L$

$[Cl_2(g)]_{initial} = 0.00 \, mol/L$

Step 2. Calculate Q and compare it to K.

At the start of the reaction Q is 0, which is less than K. The chemical reaction system must shift to the right, toward the products.

Step 3. Use an ICE table to determine the required equilibrium concentrations.

	$2 \, HCl(g) \rightleftharpoons$	$H_2(g)$ +	$Cl_2(g)$
I	2.00	0.00	0.00
C	$-x$	$+0.5x$	$+0.5x$
E	$2.00 - x$	$0.5x$	$0.5x$

Step 4. Substitute the equilibrium concentrations into the equilibrium constant equation, and solve for the unknown.

$$K = \frac{[H_2(g)][Cl_2(g)]}{[HCl(g)]^2}$$

$$3.2 \times 10^{-34} = \frac{(0.5x)(0.5x)}{(2.00-x)^2}$$

$$\sqrt{3.2 \times 10^{-34}} = \sqrt{\frac{(0.5x)(0.5x)}{(2.00-x)^2}}$$

$$\pm 1.79 \times 10^{-17} = \frac{0.5x}{2.00-x}$$

$$(\pm 1.79 \times 10^{-17})(2.00-x) = 0.5x$$

$$(\pm 3.58 \times 10^{-17}) - (1.79 \times 10^{-17})x = 0.5x$$

$$\pm 3.58 \times 10^{-17} = 0.5x + (1.79 \times 10^{-17})x$$

$$\pm 3.58 \times 10^{-17} = [0.5 + (1.79 \times 10^{-17})]x$$

$$x = \frac{\pm 3.58 \times 10^{-17}}{0.5 + (1.79 \times 10^{-17})}$$

$$x = \pm 7.2 \times 10^{-17} \text{ mol/L}$$

Substituting the negative value of x into the expressions for equilibrium concentrations in the ICE table results in a negative concentration value for $H_2(g)$ and $Cl_2(g)$, which is not possible in the real world. Therefore, use $x = 7.2 \times 10^{-17}$ mol/L in the equilibrium concentration calculations.

Step 4. Calculate the equilibrium concentrations.

$$[HCl(g)]_{equilibrium} = 0.200 - x$$

$$= 0.200 - (7.2 \times 10^{-17})$$

$$[HCl(g)]_{equilibrium} = 0.200 \text{ mol/L}$$

$$[H_2(g)]_{equilibrium} = 0.5x$$

$$= 0.5(7.2 \times 10^{-17})$$

$$[H_2(g)]_{equilibrium} = 3.6 \times 10^{-17} \text{ mol/L}$$

$$[Cl_2(g)]_{equilibrium} = [H_2(g)]_{equilibrium} = 3.6 \times 10^{-17} \text{ mol/L}$$

Statement: The equilibrium concentration for hydrogen chloride gas is 0.200 mol/L, and the equilibrium concentration for both hydrogen gas and chlorine gas is 3.6×10^{-17}.

Section 7.5 Questions, page 459

1. (a) $N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$
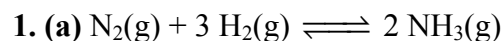
(b) Given: $V = 1.00$ L; $n_{\text{initial } H_2\,(g)} = 2.00$ mol; $n_{\text{initial } N_2\,(g)} = 1.00$ mol;

$n_{\text{initial } NH_3\,(g)} = 2.00$ mol; $K = 1.2 \times 10^2$

Required: Q

Solution:

Step 1. Write concentrations, c, in mol/L from the given amounts of all entities, using

$c = \dfrac{n}{V}$.

$[H_2(g)]_{\text{initial}} = 2.00$ mol/L

$[N_2(g)]_{\text{initial}} = 1.00$ mol/L

$[NH_3(g)]_{\text{initial}} = 2.00$ mol/L

Step 2. Use the balanced equation in (a) to write the equilibrium constant equation for the reaction system and calculate Q from the concentrations determined in Step 1.

$$Q = \frac{[NH_3(g)]^2}{[H_2(g)]^3[N_2(g)]}$$

$$= \frac{(2.00)^2}{(2.00)^3(1.00)}$$

$Q = 0.500$

Statement: The reaction quotient, Q, for the reaction system is 0.500.

(c) The chemical system is not at equilibrium. We have been given $K = 1.2 \times 10^2$, and have determined that $Q = 0.500$ in (b). The value of Q is less than the value of K. To reach equilibrium, the equilibrium system must shift to the right, toward the product.

2. For the chemical reaction system represented by

$$H_2O(g) + Cl_2O(g) \rightleftharpoons 2\,HOCl(g)$$
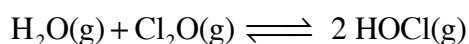
(a) Given: $V = 1.0$ L; $n_{\text{initial } HOCl(g)} = 1.0$ mol; $n_{\text{initial } Cl_2O(g)} = n_{\text{initial } H_2O(g)} = 0.10$ mol;

$K = 0.0900$

Required: Q

Solution:

Step 1. Write concentrations, c, in mol/L from the given amounts of all entities, using

$c = \dfrac{n}{V}$.

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

$[H_2O(g)]_{\text{initial}} = [Cl_2O(g)]_{\text{initial}} = 0.10$ mol/L

Step 2. Write the equilibrium constant equation for the reaction system and calculate Q from the calculated concentrations:

$$Q = \frac{[HOCl(g)]^2}{[H_2O(g)][Cl_2O(g)]}$$

$$= \frac{(1.0)^2}{(0.10)(0.10)}$$

$$Q = 100$$

Since $K = 0.0900$, Q is greater than K, and the system is not at equilibrium.

Step 3. Since Q is greater than K, the system has more of the product now than it would have at equilibrium. To reach equilibrium, the equilibrium system must shift to the left (toward the reactants).

Statement: At the given initial conditions, this chemical equilibrium system will shift to the left (toward the reactants) to achieve equilibrium.

(b) Given: $V = 2.0$ L; $n_{\text{initial HOCl(g)}} = 0.084$ mol; $n_{\text{initial Cl}_2\text{O(g)}} = 0.080$ mol;

$n_{\text{initial H}_2\text{O(g)}} = 0.98$ mol; $K = 0.0900$

Required: Q

Solution:

Step 1. Write concentrations, c, in mol/L from the given amounts of all entities, using $c = \dfrac{n}{V}$.

$$[HOCl(g)]_{\text{initial}} = \frac{0.084 \text{ mol}}{2.0 \text{ L}}$$

$[HOCl(g)]_{\text{initial}} = 0.042$ mol/L (one extra digit carried)

$$[H_2O(g)]_{\text{initial}} = \frac{0.98 \text{ mol}}{2.0 \text{ L}}$$

$[H_2O(g)]_{\text{initial}} = 0.49$ mol/L

$$[Cl_2O(g)]_{\text{initial}} = \frac{0.080 \text{ mol}}{2.0 \text{ L}}$$

$[Cl_2O(g)]_{\text{initial}} = 0.04$ mol/L

Step 2. Write the equilibrium constant equation for the reaction system and calculate Q from the calculated concentrations:

$$Q = \frac{[HOCl(g)]^2}{[H_2O(g)][Cl_2O(g)]}$$

$$= \frac{(0.042)^2}{(0.49)(0.04)}$$

$$Q = 0.09$$

Since $K = 0.0900$, Q is equal to K, so the reaction is at equilibrium.

Statement: At the given initial conditions, this reaction system is at equilibrium.

(c) Given: $V = 3.0$ L; $n_{\text{initial HOCl(g)}} = 0.25$ mol; $n_{\text{initial Cl}_2\text{O(g)}} = 0.0010$ mol; ;

$n_{\text{initial H}_2\text{O(g)}} = 0.56$ mol; $K = 0.0900$

Required: Q

Solution:

Step 1. Write concentrations, c, in mol/L from the given amounts of all entities, using

$c = \dfrac{n}{V}$.

$[\text{HOCl(g)}]_{\text{initial}} = \dfrac{0.25 \text{ mol}}{3.0 \text{ L}}$

$[\text{HOCl(g)}]_{\text{initial}} = 0.08$ mol/L

$[\text{H}_2\text{O(g)}]_{\text{initial}} = \dfrac{0.56 \text{ mol}}{3.0 \text{ L}}$

$[\text{H}_2\text{O(g)}]_{\text{initial}} = 0.19$ mol/L

$[\text{Cl}_2\text{O(g)}]_{\text{initial}} = \dfrac{0.0010 \text{ mol}}{3.0 \text{ L}}$

$[\text{Cl}_2\text{O(g)}]_{\text{initial}} = 3.3 \times 10^{-4}$ mol/L

Step 2. Write the equilibrium constant equation for the reaction system and calculate Q from the calculated concentrations:

$Q = \dfrac{[\text{HOCl(g)}]^2}{[\text{H}_2\text{O(g)}][\text{Cl}_2\text{O(g)}]}$

$= \dfrac{(0.08)^2}{(0.19)(3.3 \times 10^{-4})}$

$Q = 102$

Since $K = 0.0900$, Q is greater than K, and the system is not at equilibrium.

Step 3. Since Q is greater than K, the system has more of the product now than it would have at equilibrium. To reach equilibrium, the equilibrium system must shift to the left (toward the reactants).

Statement: At the given initial conditions, this chemical equilibrium system will shift to the left (toward the reactants) to achieve equilibrium.

3. Given: $[\text{NO}_2\text{(g)}]_{\text{initial}} = 0.850$ mol/L; $[\text{N}_2\text{O}_4\text{(g)}]_{\text{initial}} = 0.00$ mol/L

$K = 1.15$

Required: $[\text{NO}_2\text{(g)}]_{\text{equilibrium}}$; $[\text{N}_2\text{O}_4\text{(g)}]_{\text{equilibrium}}$

Solution:

Step 1. Since initially there are no products, Q is 0, which is smaller than K. The equilibrium system will shift to the right, in the direction of products.

Step 2. Use an ICE table to determine the relationship between the equilibrium concentrations of the reactant and the products.

	2 NO$_2$(g) $\rightleftharpoons$	N$_2$O$_4$(g)
I	0.850	0.000
C	$-x$	$+0.5x$
E	$0.850 - x$	$0.5x$

Step 3. Substitute the expressions from the ICE table into the equilibrium constant equation. The equilibrium law for this reaction is

$$K = \frac{[\text{N}_2\text{O}_4(\text{g})]}{[\text{NO}_2(\text{g})]^2}$$

$$1.15 = \frac{0.5x}{(0.850 - x)^2}$$

This equation cannot be solved by taking the square root of both sides. Apply the hundred rule to see if x can be assumed to be negligible.

$$\frac{[\text{NO}_2(\text{g})]_{\text{initial}}}{K} = \frac{0.850}{1.15}$$

$$\frac{[\text{NO}_2(\text{g})]_{\text{initial}}}{K} = 0.739$$

Since 0.110 is much less than 100, the assumption that $0.850 + x \approx 0.850$ is not warranted. Therefore, you must solve the equation.

$$1.15(x^2 - 1.70x + 0.7225) = 0.5x$$

$$1.15x^2 - 2.455x + 0.831 = 0$$

Step 4. Solve the equation by using the quadratic formula.

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$= \frac{-(-2.455) \pm \sqrt{(-2.455)^2 - 4(1.15)(0.831)}}{2(1.15)}$$

$$= \frac{2.455 \pm 1.485}{2.3}$$

$$x = 1.71$$

or

$$x = 0.422$$

The value of x is 0.422 mol/L because 1.71 mol/L would give a negative value for $[NO_2(g)]_{equilibrium}$. Use this value to calculate the equilibrium concentrations for $NO_2(g)$ and $N_2O_4(g)$ from the expressions in the ICE table.

$[NO_2(g)]_{equilibrium} = 0.850 - x$

$\qquad\qquad = 0.850 - 0.422$

$[NO_2(g)]_{equilibrium} = 0.428$ mol/L

$[N_2O_4(g)]_{equilibrium} = 0.5x$

$\qquad\qquad = 0.5(0.422)$

$[N_2O_4(g)]_{equilibrium} = 0.211$ mol/L

Step 5. Check your results by calculating K to see if it equals the given value of 1.15.

$K = \dfrac{[N_2O_4(g)]}{[NO_2(g)]^2}$

$\quad = \dfrac{0.211}{(0.428)^2}$

$K = 1.15$

The calculated value of 1.15 for K equals the given value of 1.15.

Statement: The solution is at equilibrium when the concentration of nitrogen dioxide is 0.428 mol/L and the concentration of dinitrogen tetroxide is 0.211 mol/L.
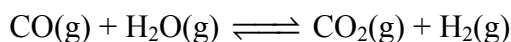
4. Given: Reactants: $CO(g) + H_2O(g)$; $V = 1.00$ L; $n_{initial\ CO(g)} = n_{initial\ H_2O(g)} = 2.00$ mol;

$n_{equilibrium\ CO_2(g)} = n_{equilibrium\ H_2(g)} = 1.30$ mol; all substances are in gas phase

Required: K

Solution:

Step 1. Write the balanced equation for the equilibrium reaction system.

$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$$

Step 2. Write the concentrations of the reactants using the given amounts.

$[CO(g)]_{initial} = [H_2O(g)]_{initial} = 2.00$ mol/L
$[CO_2(g)]_{initial} = [H_2(g)]_{initial} = 0.00$ mol/L
$[CO_2(g)]_{equilibrium} = [H_2(g)]_{equilibrium} = 1.30$ mol/L

Step 3. Use an ICE table to determine the relationship between the equilibrium concentrations of the reactant and the products.

	CO(g) +	H₂O(g) ⇌	CO₂(g) +	H₂(g)
I	2.00	2.00	0.00	0.00
C	$-x$	$-x$	$+x$	$+x$
E	$2.00 - x$	$2.00 - x$	x	x

Step 4. Calculate the value of x.

$x = [CO_2(g)]_{equilibrium}$

$\quad = [H_2(g)]_{equilibrium}$

$x = 1.30$ mol/L

Step 5. Calculate the equilibrium concentrations.

$[CO(g)]_{equilibrium}$ $= 2.00 - x$

$= 2.00 - 1.30$

$[CO(g)]_{equilibrium}$ $= 0.70$ mol/L

and

$[H_2O(g)]_{equilibrium} = 0.70$ mol/L

Step 6: Determine the equilibrium constant, K.

$$K = \frac{[CO_2(g)][H_2(g)]}{[CO(g)][H_2O(g)]}$$

$$= \frac{(1.30)(1.30)}{(0.70)(0.70)}$$

$K = 3.4$

Statement: The equilibrium constant, K, for the reaction is 3.4, assuming all substances are in the gas phase.

5. Given: $2\ SO_3(g) \rightleftharpoons 2\ SO_2(g) + O_2(g)$; $V = 3.0$ L; $n_{initial\ SO_3\ (g)} = 12.0$ mol;

$n_{equilibrium\ SO_2\ (g)} = 3.0$ mol

Required: K

Solution:

Step 1. Write concentrations, c, in mol/L from the given amounts of all entities, using

$c = \dfrac{n}{V}$.

$[SO_3(g)]_{initial} = \dfrac{12.0\ mol}{3.0\ L}$

$[SO_3(g)]_{initial} = 4.0$ mol/L

$[SO_2(g)]_{initial} = [O_2(g)]_{initial} = 0.00$ mol/L

$[SO_2(g)]_{equilibrium} = \dfrac{3.0\ mol}{3.0\ L}$

$[SO_2(g)]_{equilibrium} = 1.0$ mol/L

Step 2. Using the balanced chemical equation, construct an ICE table to determine the relationship between the equilibrium concentrations of the reactant and the products.

	2 SO₃(g) $\rightleftharpoons$	2 SO₂(g) +	O₂(g)
I	4.0	0.0	0.0
C	$-x$	$+x$	$+0.5x$
E	$4.0 - x$	1.0	$+0.5x$

Step 3. Calculate the value of x.

$x = [SO_2(g)]_{equilibrium}$

$x = 1.0$ mol/L

$[SO_3(g)]_{equilibrium} = 4.0 - x$

$\qquad\qquad\qquad = 4.0 - 1.0$

$[SO_3(g)]_{equilibrium} = 3.0$ mol/L

$[O_2(g)]_{equilibrium} = 0.5x$

$\qquad\qquad\qquad = 0.5(1.0)$

$[O_2(g)]_{equilibrium} = 0.50$ mol/L

Step 4: Determine the equilibrium constant, K.

$$K = \frac{[SO_2(g)]^2[O_2(g)]}{[SO_3(g)]^2}$$

$$= \frac{(1.0)^2(0.50)}{(3.0)^2}$$

$K = 5.6 \times 10^{-2}$

Statement: The equilibrium constant, K, for the reaction is 5.6×10^{-2}.

6. Given: $SO_2(g) + NO_2(g) \rightleftharpoons SO_3(g) + NO(g)$; $K = 3.75$;

$[SO_2]_{initial} = [NO_2(g)]_{initial} = [SO_3(g)]_{initial} = [NO(g)]_{initial} = 0.800$ mol/L

Required: $[SO_2]_{equilibrium}$; $[NO_2(g)]_{equilibrium}$; $[SO_3(g)]_{equilibrium}$; $[NO(g)]_{equilibrium}$

Solution:

Step 1. Use an ICE table to determine the relationship between the equilibrium concentrations of the reactant and the products.

	SO₂(g) +	NO₂(g) ⇌	SO₃(g) +	NO(g)
I	0.800	0.800	0.800	0.800
C	$-x$	$-x$	$+x$	$+x$
E	$0.800 - x$	$0.800 - x$	$0.800 + x$	$0.800 + x$

Step 2. Calculate the value of x by substituting the given value of K and the concentrations from the ICE table into the equilibrium constant equation.

$$K = \frac{[SO_3(g)][NO(g)]}{[SO_2(g)][NO_2(g)]}$$

$$3.75 = \frac{(0.800+x)(0.800+x)}{(0.800-x)(0.800-x)}$$

$$\sqrt{3.75} = \sqrt{\frac{(0.800+x)(0.800+x)}{(0.800-x)(0.800-x)}}$$

$$\pm 1.94 = \frac{(0.800+x)}{(0.800-x)}$$

$$\pm 1.94(0.800-x) = 0.800+x$$

$$\pm 1.552 - 1.94x = 0.800+x$$

$$\pm 1.552 - 0.800 = 2.94x$$

$$x = 0.256 \text{ mol/L}$$

$$\text{or}$$

$$x = -0.713 \text{ mol/L}$$

Substituting the negative value of x into the expressions for equilibrium concentrations in the ICE table results in positive concentration values for all entities. To determine which value of x to use, substitute each value for x in the algebraic equation for the equilibrium constant:

$$K = \frac{(0.800+x)(0.800+x)}{(0.800-x)(0.800-x)} \qquad\qquad K = \frac{(0.800+x)(0.800+x)}{(0.800-x)(0.800-x)}$$

$$= \frac{(0.800+0.256)(0.800+0.256)}{(0.800-0.256)(0.800-0.256)} \qquad = \frac{(0.800-0.713)(0.800-0.713)}{(0.800+0.713)(0.800+0.713)}$$

$$K = 3.77 \qquad\qquad\qquad\qquad K = 0.003$$

$K = 3.77$ is close to he equilibrium constant of 3.75; therefore, use $x = 2.56$ mol/L in the equilibrium concentration calculations.

Step 3. Calculate the equilibrium concentrations.

$$[SO_2]_{\text{equilibrium}} = 0.800 - x \qquad\qquad [SO_3]_{\text{equilibrium}} = 0.800 + x$$

$$= 0.800 - 0.256 \qquad\qquad\qquad = 0.800 + 0.256$$

$$[SO_2]_{\text{equilibrium}} = 0.544 \text{ mol/L} \qquad\qquad [SO_2]_{\text{equilibrium}} = 1.06 \text{ mol/L}$$

$$[NO_2(g)]_{\text{equilibrium}} = [SO_2]_{\text{equilibrium}} \qquad\qquad [NO(g)]_{\text{equilibrium}} = [SO_3]_{\text{equilibrium}}$$

$$[NO_2(g)]_{\text{equilibrium}} = 0.544 \text{ mol/L} \qquad\qquad [NO(g)]_{\text{equilibrium}} = 1.06 \text{ mol/L}$$

Statement: The equilibrium concentration for both sulfur dioxide gas and nitrogen dioxide gas is 0.544 mol/L. The equilibrium concentration for both sulfur trioxide gas and nitric oxide gas is 1.06 mol/L.

7. Given: $2\ H_2O(g) \rightleftharpoons 2\ H_2(g) + O_2(g)$; $[H_2O(g)]_{equilibrium} = 1.1 \times 10^{-1}$ mol/L; $[H_2(g)]_{equilibrium} = 1.2 \times 10^{-2}$ mol/L; $K = 2.4 \times 10^{-3}$

Required: $[O_2(g)]_{equilibrium}$

Analysis: $K = \dfrac{[H_2(g)]^2[O_2(g)]}{[H_2O(g)]^2}$

Solution:

$$K = \frac{[H_2(g)]^2[O_2(g)]}{[H_2O(g)]^2}$$

$$2.4 \times 10^{-3} = \frac{(1.2 \times 10^{-2})^2[O_2(g)]}{(1.1 \times 10^{-1})^2}$$

$$[O_2(g)] = \frac{(2.4 \times 10^{-3})(1.1 \times 10^{-1})^2}{(1.2 \times 10^{-2})^2}$$

$$[O_2(g)]_{equilibrium} = 0.20\ \text{mol/L}$$

Statement: The equilibrium concentration of oxygen gas is 0.20 mol/L for the given reaction.

8. Given: $HCO_3^-(aq) \rightleftharpoons H^+(aq) + CO_3^{2-}(aq)$; $[HCO_3^-(aq)]_{initial} = 0.16$ mol/L; $[CO_3^{2-}(aq)]_{initial} = 0.00$ mol/L; $K = 5.6 \times 10^{-11}$

Required: $[CO_3^{2-}(aq)]_{equilibrium}$

Analysis:

	$HCO_3^-(aq) \rightleftharpoons$	$H^+(aq)$ +	$CO_3^{2-}(aq)$
I	0.16	0.00	0.00
C	$-x$	$+x$	$+x$
E	$0.16 - x$	x	x

$$K = \frac{[H^+(aq)][CO_3^{2-}(aq)]}{[HCO_3^-(aq)]}$$

Solution:

$$K = \frac{[H^+(aq)][CO_3^{2-}(aq)]}{[HCO_3^-(aq)]}$$

$$5.6 \times 10^{-11} = \frac{(x)(x)}{0.16 - x}$$

$$(5.6 \times 10^{-11})(0.16 - x) = x^2$$

$$x^2 + (5.6 \times 10^{-11})(x) - (9.0 \times 10^{-12}) = 0$$

Solve the equation by using the quadratic formula.

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$= \frac{-(5.6 \times 10^{-11}) \pm \sqrt{(5.6 \times 10^{-11})^2 - 4(1)(-9.0 \times 10^{-12})}}{2(1)}$$

$$= \frac{-(5.6 \times 10^{-11}) \pm (6 \times 10^{-6})}{2}$$

$$x = \pm 3.00 \times 10^{-6}$$

The value of x is 3.0×10^{-6} mol/L because -3.0×10^{-6} mol/L would give negative values for the reactants and products. Use this value to calculate the equilibrium concentrations for the carbonate ion from the expression in the ICE table.

$$x = 3.0 \times 10^{-6}$$

$$[CO_3^{2-}(aq)]_{equilibrium} = 3.0 \times 10^{-6} \text{ mol/L}$$

Check your results by calculating K to see if it equals the given value of 5.6×10^{-11}.

$$K = \frac{(x)(x)}{0.16 - x}$$

$$= \frac{(3.0 \times 10^{-6})^2}{0.16 - (3.0 \times 10^{-6})}$$

$$K = 5.6 \times 10^{-11}$$

The calculated value of 5.6×10^{-11} for K equals the given value of 5.6×10^{-11}.

Statement: The equilibrium concentration of carbonate ion is of 3.0×10^{-6}.