Calculating and Using a Reaction Quotient
(Student textbook page 461)

81. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. At 300 K, K_{eq} = 35.5 for the reaction below:

$3C(s) + 3H_2(g) \rightleftharpoons CH_4(g) + C_2H_2(g)$

The following concentrations are found to be present at a particular point in time:
$[H_2(g)]$ = 0.34 mol/L, $[CH_4(g)]$ = 2.13 mol/L, and$[C_2H_2(g)]$ = 1.77 mol/L.

What Is Required?
You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the value of the equilibrium constant, K_{eq}: 35.5
You know the concentration of the components:
$[H_2(g)]$ = 0.34 mol/L
$[CH_4(g)]$ = 2.13 mol/L
$[C_2H_2(g)]$ = 1.77 mol/L

Plan Your Strategy	Act on Your Strategy
Write the expression for the reaction quotient, Q_{eq}.	$Q_{eq} = \frac{[CH_4][C_2H_2]}{[H_2]^3}$
Substitute the given concentrations into this reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_{eq}.	$Q_{eq} = \frac{[CH_4][C_2H_2]}{[H_2]^3}$ $= \frac{(2.13)(1.77)}{(0.34)^3}$ $= 95.9$ $= 96$
Compare the calculated value of Q_{eq} and the given K_{eq} to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_{eq} > K_{eq}$ the system is not at equilibrium. The ratio of products to reactants is too large. The reaction will shift to the left to lower the concentration of products and increase the concentration of reactants.

Check Your Solution
The expression for the reaction quotient has the products in the numerator and the reactant in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

82. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. A 3.00 L reaction vessel contains 9.00 mol of hydrogen iodide gas, 6.00 mol of hydrogen gas, and 4.50 mol of iodine gas. At a given temperature, the K_{eq} value for the reaction below is 50.0.

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

What Is Required?

You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?

You know the balanced chemical equation for the reaction.
You know the volume of reaction vessel: 3.00 L
You know the value of the equilibrium constant, K_{eq}: 50.0.
You know the amounts in moles of the components:
$H_2(g)$ = 6.00 mol
$HI(g)$ = 9.00 mol
$I_2(g)$ = 4.50 mol

Plan Your Strategy	Act on Your Strategy
Express the concentration of each component in mol/L.	$[H_2] = \frac{6.0\text{ mol}}{3.0\text{ L}} = 2.0\text{ mol/L}$ $[HI] = \frac{9.0\text{ mol}}{3.0\text{ L}} = 3.0\text{ mol/L}$ $[I_2] = \frac{4.5\text{ mol}}{3.0\text{ L}} = 1.5\text{ mol/L}$
Write the expression for the reaction quotient, Q_{eq}.	$Q_{eq} = \frac{[HI]^2}{[H_2][I_2]}$
Substitute the given concentrations into this reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_{eq}.	$Q_{eq} = \frac{[HI]^2}{[H_2][I_2]} = \frac{(3.0)^2}{(2.0)(1.5)} = 3.0$

Compare the calculated value of Q_{eq} and the given K_{eq} to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_{eq} < K_{eq}$ the system is not at equilibrium. The ratio of products to reactants is too small. The reaction will shift to the right to increase the concentration of products and decrease the concentration of reactants.

Check Your Solution

The expression for the reaction quotient has the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

83. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. At 525 K, $K_{eq} = 0.041$ for the following reaction:

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$$

At one point in time, measurements were taken and the following concentrations were measured: $[PCl_5(g)] = 0.77$ mol/L, $[PCl_3(g)] = 1.21$ mol/L, and $[Cl_2(g)] = 0.49$ mol/L.

What Is Required?
You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know that $K_{eq} = 0.041$.
You know the concentration of the components:
$PCl_5(g)] = 0.77$ mol/L
$PCl_3(g)] = 1.21$ mol/L
$[Cl_2(g)] = 0.49$ mol/L

Plan Your Strategy	Act on Your Strategy
Write the expression for the reaction quotient, Q_{eq}.	$Q_{eq} = \frac{[PCl_3][Cl_2]}{[PCl_5]}$
Substitute the given concentrations into this reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_{eq}.	$Q_{eq} = \frac{[PCl_3][Cl_2]}{[PCl_5]}$ $= \frac{(1.21)(0.49)}{(0.77)}$ $= 0.77$
Compare the calculated value of Q_{eq} and the given K_{eq} to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_{eq} > K_{eq}$ the system is not at equilibrium. The ratio of products to reactants is too large. The reaction will shift to the left to lower the concentration of products and increase the concentration of reactants.

Check Your Solution
The expression for the reaction quotient has the products in the numerator and the reactant in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

84. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. Consider the following reaction:

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

$K_{eq} = 6.0 \times 10^{-2}$ at 875 K. At a point just after nitrogen gas and hydrogen gas are injected into a container, the concentrations of each reactant and product are $[N_2(g)] = 1.50 \times 10^{-5}$ mol/L , $[H_2(g)] = 0.354$ mol/L, and$[NH_3(g)] = 2.00 \times 10^{-4}$ mol/L.

What Is Required?
You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know that $K_{eq} = 6.0 \times 10^{-2}$.
You know the concentration of components:
$[N_2(g)] = 1.50 \times 10^{-5}$ mol/L
$[H_2(g)] = 0.354$ mol/L
$[NH_3(g)] = 2.00 \times 10^{-4}$ mol/L

Plan Your Strategy	Act on Your Strategy
Write the expression for the reaction quotient, Q_{eq}.	$Q_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$
Substitute the given concentrations into this reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_{eq}.	$Q_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$ $= \frac{(2.00\times10^{-4})^2}{(1.50\times10^{-5})(0.354)^3}$ $= 0.060$
Compare the calculated value of Q_{eq} and the given K_{eq} to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_{eq} = K_{eq}$ the system is at equilibrium. There is no shift in the position of equilibrium.

Check Your Solution
The expression for the reaction quotient has the products in the numerator and the reactant in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

85. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. In the reaction of phosphorus trichloride, $PCl_3(g)$, and ammonia, $NH_3(g)$, hydrogen chloride gas and $P(NH_2)_3(g)$ forms by this reaction:

$$PCl_3(g) + 3NH_3(g) \rightleftharpoons P(NH_2)_3(g) + 3HCl(g)$$

At 600 K, this reaction has an equilibrium constant, K_{eq}, that is 142.3. A system of these gases was analyzed and the following concentrations were determined: $[PCl_3(g)] = 2.15$ mol/L , $[NH_3(g)] = 1.21$ mol/L, $[P(NH_2)_3(g)] = 5.74$ mol/L and $[HCl(g)] = 2.04$ mol/L.

What Is Required?
You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?
You know the balanced chemical equation.
You know that $K_{eq} = 142.3$.
You know the concentration of the components:
$[PCl_3(g)] = 2.15$ mol/L
$[NH_3(g)] = 1.21$ mol/L
$[P(NH_2)_3(g)] = 5.74$ mol/L
$[HCl(g)] = 2.04$ mol/L

Plan Your Strategy	Act on Your Strategy
Write the expression for the reaction quotient, Q_{eq}.	$Q_{eq} = \frac{[P(NH_2)_3][HCl]^3}{[PCl_3][NH_3]^3}$
Substitute the given concentrations into the reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_{eq}.	$Q_{eq} = \frac{[P(NH_2)_3][HCl]^3}{[PCl_3][NH_3]^3}$ $= \frac{(5.74)(2.04)^3}{(2.15)(1.21)^3}$ $= 12.8$
Compare the calculated value of Q_{eq} and the given K_{eq} to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_{eq} < K_{eq}$, the system is not at equilibrium. The ratio of products to reactants is too small. The reaction will shift to the right to increase the concentration of products and decrease the concentration of reactants.

Check Your Solution
The expression for the reaction quotient has the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

86. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. At 865 K, $K_p = 2.11$ for the reaction below:

$$P_4(s) + 6Cl_2(g) \rightleftharpoons 4PCl_3(g)$$

At one point in time, measurements were taken and the following partial pressures were measured: $P_{Cl_2} = 1.21\text{atm}$, $P_{PCl_3} = 3.04$ atm

What Is Required?
You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?
You know the balanced chemical equation.
You know the value of K_p: 2.11
You know the partial pressures of the components:
$P_{Cl_2} = 1.21$ atm
$P_{PCl_3} = 3.04$ atm

Plan Your Strategy	Act on Your Strategy
Write the expression for the reaction quotient, Q_p.	$Q_p = \frac{P_{PCl_3}^4}{P_{Cl_2}^6}$
Substitute the given concentrations into this reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_p.	$Q_p = \frac{P_{PCl_3}^4}{P_{Cl_2}^6}$ $= \frac{(3.04)^4}{(1.21)^6}$ $= 27.2$
Compare the calculated value of Q_p and the given K_p to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_p > K_p$ the system is not at equilibrium. The ratio of products to reactants is too large. The reaction will shift to the left to lower the concentration of products and increase the concentration of reactants.

Check Your Solution
The expression for the reaction quotient has the product in the numerator and the reactant in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

87. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. At 660 K, $K_p = 25.2$ for the reaction below:

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$$

At one point in time, measurements were taken and the following partial pressures were measured:
$P_{N_2} = 2.37$ atm, $P_{O_2} = 2.27$ atm, and $P_{NO} = 3.74$ atm.

What Is Required?
You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know that $K_p = 25.2$.
You know the partial pressures of the components:
$P_{NO} = 3.74$ atm
$P_{N_2} = 2.37$ atm
$P_{O_2} = 2.27$ atm

Plan Your Strategy	Act on Your Strategy
Write the expression for the reaction quotient, Q_p.	$Q_p = \frac{P_{NO}^2}{P_{N_2}P_{O_2}}$
Substitute the given concentrations into the reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_p.	$Q_p = \frac{P_{NO}^2}{P_{N_2}P_{O_2}}$ $= \frac{(3.74)^2}{(2.37)(2.27)}$ $= 2.60$
Compare the calculated value of Q_p and the given K_p to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_p < K_p$ the system is not at equilibrium. The ratio of products to reactants is too small. The reaction will shift to the right to increase the concentration of products and decrease the concentration of reactants.

Check Your Solution
The expression for the reaction quotient has the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.