Equilibrium Law and the Equilibrium Constant

7.2

In the mid-1800s, two Norwegian chemists, Cato Maximilian Guldberg and Peter Waage (**Figure 1**), conducted detailed studies of chemical equilibrium systems. Based on their observations, they were able to describe a chemical equilibrium in a closed system in mathematical terms. The mathematical description is known as the **equilibrium law**. The equilibrium law is sometimes called the law of mass action.

What is the equilibrium law and why is it useful? To start our discussion, suppose that A, B, C, and D are chemical entities in gas or aqueous states, and that *a*, *b*, *c*, and *d* are the coefficients in a balanced chemical equation. We can now represent a chemical equilibrium system by this general chemical equation:

$$aA + bB \rightleftharpoons cC + dD$$

The equilibrium law for this general reaction may be represented quantitatively by the following mathematical equation:

$$K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

The square brackets in the equation indicate the concentrations of the chemical entities at equilibrium. K is the symbol for the **equilibrium constant**, a constant numerical value that defines the equilibrium law for a given system. Since the units of the equilibrium constant vary according to the values of *a*, *b*, *c*, and *d*, equilibrium constant values are usually reported without units.

Figure 1 Cato Maximilian Guldberg (1836–1902) and Peter Waage (1833–1900) first proposed the equilibrium law in 1864.

equilibrium law the mathematical description of a chemical system at equilibrium

equilibrium constant (*K*) a constant numerical value defining the equilibrium law for a given system; units are not included when giving the value of K

Analyzing the Equilibrium Law

To see how the equilibrium law is useful, consider the formation of hydrogen iodide gas, HI(g), from its elements. The balanced chemical equation for this reaction is

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

Table 1 shows the results from an investigation to observe how changing the initial concentration of the reactants and products affected the equilibrium. The reaction was allowed to proceed to equilibrium in a sealed 2.00 L flask at 485 °C.

Table 1 Equilibria for Formation of Hydrogen Iodide from Its Elements at 485 °C

Trial	Initial concentration (mol/L)			Equilibrium concentration (mol/L)		
	$[H_2(g)]$	$[I_2(g)]$	[HI(g)]	$[H_2(g)]$	$[I_2(g)]$	[HI(g)]
1	2.00	2.00	0	0.442	0.442	3.119
2	0	0	2.000	0.221	0.221	1.560
3	0	0.010	0.350	0.035	0.045	0.280

In **Table 2**, the data from each trial were entered into the equilibrium law equation. What do you notice about the value of the equilibrium constants?

Table 2 Calculation of K from Experimental Results in Table 1

Trial 1	Trial 2	Trial 3
$K = \frac{[HI(g)]^2}{[H_2(g)][I_2(g)]}$	$K = \frac{[HI(g)]^2}{[H_2(g)][I_2(g)]}$	$K = \frac{[HI(g)]^2}{[H_2(g)][I_2(g)]}$
$K = \frac{(3.119)^2}{(0.442)(0.442)}$	$K = \frac{(1.560)^2}{(0.221)(0.221)}$	$K = \frac{(0.280)^2}{(0.035)(0.045)}$
$K = 49.8$	$K = 49.8$	$K = 49.8$

Investigation 7.2.1

The Equilibrium Law (page 473)
This activity provides an opportunity for you to use experimental data to find the mathematical relationship for the equilibrium law.

Table 3 Equilibrium Constant for the Production of Ammonia Gas from Elemental Nitrogen and Hydrogen at Various Temperatures

Temperature (°C)	*K*
25	4.26×10^{8}
300	1.02×10^{-5}
400	8.00×10^{-7}

The equilibrium constant value is the same, 49.8, for all three trials. The equilibrium constant will always have the same value of 49.8 regardless of the initial concentrations, providing that all other variables are kept the same. Chemists have found that this holds true for any specific chemical reaction at equilibrium at a given temperature: the value of the equilibrium constant remains the same, regardless of initial concentrations.

However, the value of the equilibrium constant depends on the reaction temperature. In **Table 3**, notice how the equilibrium constant value changes with temperature when gaseous nitrogen and hydrogen are combined in a closed vessel and react to form ammonia gas, $NH_3(g)$. The chemical equation for this reaction is

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

Writing Equilibrium Law Equations and Calculating *K*

In Section 7.1, you saw that the equilibrium position of a reaction describes the relative concentration of reactants and products in a chemical reaction system. By convention, the equilibrium position is expressed as being far to the right (toward products) or far to the left (toward reactants). The equilibrium constant, *K*, can be used to determine the equilibrium position of a chemical reaction system. You will calculate the equilibrium constant from the equilibrium law equation in Tutorial 1. Always use amount concentrations of the reactants and products at equilibrium in the equilibrium law expression calculation. Remember that, in this textbook, the concentrations of aqueous solutions and gaseous substances are provided in mol/L.

Tutorial 1 / Writing Equilibrium Law Equations and Calculating *K*

In this tutorial, you will first see how to write an equilibrium equation from balanced chemical equations. Then, you will use equilibrium law equations to calculate the equilibrium constant, *K*, for several chemical reaction systems.

Sample Problem 1: Writing an Equilibrium Law Equation from a Balanced Chemical Equation

Write the equilibrium equation for the reaction described by the following balanced equation:

$$4\,NH_3(g) + 7\,O_2(g) \rightleftharpoons 4\,NO_2(g) + 6\,H_2O(g)$$

Solution

Step 1. Identify the reactants and products and their coefficients.

Reactants: $4\,NH_3(g)$; $7\,O_2(g)$

Products: $4\,NO_2(g)$; $6\,H_2O(g)$

Step 2. Apply the equilibrium law equation.

Place the products in the numerator and the reactants in the denominator. Use the coefficients from the balanced equation as exponents.

$$K = \frac{[NO_2(g)]^4[H_2O(g)]^6}{[NH_3(g)]^4[O_2(g)]^7}$$

Sample Problem 2: Calculating *K* for a Synthesis Reaction

In a closed vessel at 500 °C, gaseous nitrogen, $N_2(g)$, and hydrogen, $H_2(g)$, combine in an equilibrium reaction to form ammonia gas, $NH_3(g)$:

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

The equilibrium concentrations of gaseous nitrogen, hydrogen, and ammonia, respectively, are 1.50×10^{-5} mol/L, 3.45×10^{-1} mol/L, and 2.00×10^{-4} mol/L.

Calculate the equilibrium constant, *K*, for this chemical reaction under these conditions.

Given: $[N_2(g)]_{equilibrium} = 1.50 \times 10^{-5}$ mol/L; $[H_2(g)]_{equilibrium} = 3.45 \times 10^{-1}$ mol/L; $[NH_3(g)]_{equilibrium} = 2.00 \times 10^{-4}$ mol/L

Required: *K*

Solution: Write the equilibrium law equation using the balanced chemical equation. Then, substitute the equilibrium concentrations into the equilibrium law equation and solve for *K*. Remember that *K* is reported without units.

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

$$= \frac{(2.00 \times 10^{-4})^2}{(1.50 \times 10^{-5})(3.45 \times 10^{-1})^3}$$

$$K = 6.49 \times 10^{-2}$$

Statement: The equilibrium constant, *K*, is 6.49×10^{-2} for the chemical reaction system at 500 °C.

Sample Problem 3: Calculating *K* for a Decomposition Reaction

In a closed vessel at 500 °C, ammonia gas, $NH_3(g)$, will decompose into its elements:

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

Calculate the equilibrium constant, *K*, for the decomposition of ammonia gas into its elements under these conditions. This reaction is the reverse of that in Sample Problem 2, so use the same equilibrium concentrations.

Given: $[N_2(g)]_{\text{equilibrium}} = 1.50 \times 10^{-5}$ mol/L; $[H_2(g)]_{\text{equilibrium}} = 3.45 \times 10^{-1}$ mol/L; $[NH_3(g)]_{\text{equilibrium}} = 2.00 \times 10^{-4}$ mol/L

Required: *K*

Solution:

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

$$= \frac{(1.50 \times 10^{-5})(3.45 \times 10^{-1})^3}{(2.00 \times 10^{-4})^2}$$

$$K = 15.4$$

Statement: The equilibrium constant, *K*, for the decomposition of ammonia at 500 °C is 15.4.

If you compare the two equilibrium constants that you calculated in Sample Problems 2 and 3, you can see that they are very different. However, these values are mathematically related.

Allow *K* to represent the equilibrium constant of the forward reaction or the formation of ammonia gas (Sample Problem 2). Then, let *K′* represent the equilibrium constant of the reverse reaction or the decomposition of ammonia gas (Sample Problem 3). The equilibrium law equations for the reactions are

$$K = \frac{[NH_3(g)]^2}{[N_2(g)][H_2(g)]^3} = 6.49 \times 10^{-2}$$

$$K' = \frac{[N_2(g)][H_2(g)]^3}{[NH_3(g)]^2} = 15.4$$

If you compare the concentrations in these two equations, you will see that *K* is the mathematical reciprocal of *K′*. To check this, you can calculate the reciprocal of *K′*:

$$K' = 15.4\text{, so}$$

$$\frac{1}{K'} = \frac{1}{15.4}$$

$$= 6.49 \times 10^{-2}$$

Practice

1. Write the equilibrium law equation for the reactions represented by the following balanced chemical equations:
   (a) $2\ CO_2(g) \rightleftharpoons 2\ CO(g) + O_2(g)$
   (b) $2\ Cl_2(g) + 2\ H_2O(g) \rightleftharpoons 4\ HCl(g) + O_2(g)$
   (c) $2\ O_3(g) \rightleftharpoons 3\ O_2(g)$
   (d) $4\ NH_3(g) + 3\ O_2(g) \rightleftharpoons 2\ N_2(g) + 6\ H_2O(g)$ T/I

2. In a closed vessel at 327 °C, methanol gas, $CH_3OH(g)$, is formed by reacting gaseous carbon monoxide with hydrogen gas. When the reaction reaches equilibrium, the concentration of carbon monoxide gas is 0.079 mol/L, of hydrogen gas is 0.158 mol/L, and of methanol gas is 0.021 mol/L. The balanced chemical equation for the chemical reaction system is

$$CO(g) + 2\,H_2(g) \rightleftharpoons CH_3OH(g)$$

Calculate the equilibrium constant for this reaction. T/I [ans: $K = 11$]

3. Write the balanced chemical equation for the reaction with the following equilibrium law equation:

$$K = \frac{[NO(g)]^2[Br_2(g)]}{[NOBr(g)]^2}$$

T/I [ans: $2\,NOBr(g) \rightleftharpoons 2\,NO(g) + Br_2(g)$]

Equilibrium Constant and Reaction Rate

When a reversible chemical reaction is at equilibrium, the rate of the forward reaction is equal to the rate of the reverse reaction. We can show why this is true. Consider a hypothetical reversible reaction in which reactants A and B form products C and D.

$$A + B \rightleftharpoons C + D$$

Since it is a reversible chemical reaction, dynamic equilibrium is achieved when the forward reaction proceeds at the same rate as the reverse reaction. The chemical equation for the forward reaction of the chemical equilibrium system is

$$A + B \rightarrow C + D$$

Assume that this reaction is an elementary process, which is a chemical process that is completed in a single step. Recall from Chapter 6 that, in a dilute solution, the rate of a chemical reaction is proportional to the amount concentrations of the reactants. If k_f is the rate constant for the forward reaction, the rate of the forward reaction is

$$\text{Forward reaction rate} = k_f[A][B]$$

Similarly, for the reverse reaction given by the chemical equation

$$C + D \rightarrow A + B$$

if k_r is the rate constant, then this equation represents the reverse reaction rate:

$$\text{Reverse reaction rate} = k_r[C][D]$$

At equilibrium, the forward reaction rate must equal the reverse reaction rate.

$$k_f[A][B] = k_r[C][D]$$

Rearranging this equation as a ratio gives

$$\frac{k_f}{k_r} = \frac{[C][D]}{[A][B]}$$

This ratio is the same as the equilibrium law equation for our hypothetical equilibrium system. Therefore, the ratio of the rate constants of the forward and reverse reactions is equal to the equilibrium constant, K.

$$\frac{k_f}{k_r} = \frac{[C][D]}{[A][B]} = K$$

For simple reversible reactions in which all the coefficients are 1, we would also arrive at this equation by placing the coefficients of 1 in the equation originally proposed by the chemists Guldberg and Waage:

$$K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

However, this mathematical relationship is true only for elementary processes. Most chemical reactions are not elementary processes.

Heterogeneous Equilibria

The chemical equilibrium systems that we have discussed so far have been examples of a **homogeneous equilibrium**, in which all reactants and products in the systems were in the same state of matter. When the reactants and products in a chemical equilibrium system are present in more than one state, the system is a **heterogeneous equilibrium**.

For example, calcium carbonate, $CaCO_3(s)$, can undergo a reversible decomposition reaction to form solid calcium oxide, $CaO(s)$, and carbon dioxide gas:

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$

This reaction is used in the commercial preparation of calcium oxide (also known as lime) from calcium carbonate sources, such as limestone. Lime has many uses, such as in the manufacture of steel and plastics. Straightforward application of the equilibrium law leads to the equation

$$K = \frac{[CO_2(g)][CaO(s)]}{[CaCO_3(s)]}$$

However, the equilibrium position of a heterogeneous equilibrium does not depend on the quantities of pure solids or liquids. The fundamental reason for this behaviour is that the concentrations of pure solids and liquids cannot change. In the reversible decomposition reaction of calcium carbonate, the concentrations of solid calcium carbonate and solid calcium oxide remain constant (**Figure 2**). CAREER LINK

If we represent these constant concentrations by the symbols C_1 and C_2 respectively, we can write the equilibrium law equation as

$$K = \frac{[CO_2(g)]C_1}{C_2}$$

Since C_1 and C_2 are constants, and K is also a constant, we can simplify the equilibrium law equation by including all the constants together. We can then write the equilibrium law equation for this reaction as

$$K = [CO_2(g)]$$

We can write the following generalization from this analysis:

> If pure solids or pure liquids are involved in a chemical equilibrium system, their concentrations are not included in the equilibrium law equation for the reaction system.

We cannot simplify the equilibrium law expression for solutions or gases, since the concentrations of substances in these states can vary. Therefore, carefully note the states of all substances in a reaction system. For example, consider the decomposition of liquid water to gaseous hydrogen and oxygen:

$$2\ H_2O(l) \rightleftharpoons 2\ H_2(g) + O_2(g)$$

The equilibrium law equation for this reaction is

$$K = [H_2(g)]^2[O_2(g)]$$

We did not include water in the equation, because it is a pure liquid. However, we could instead carry out the reaction under conditions where water is in the gas state.

$$2\ H_2O(g) \rightleftharpoons 2\ H_2(g) + O_2(g)$$

Since the concentration of water vapour can change, we must now include it the equilibrium law equation:

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

homogeneous equilibrium a chemical equilibrium system in which all reactants and products are in the same state of matter, such as the gas state

heterogeneous equilibrium a chemical equilibrium system in which the reactants and products are present in at least two different states, such as gases and solids

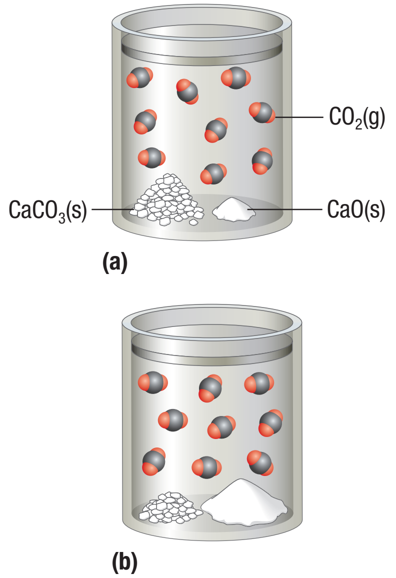

Figure 2 The position of the equilibrium $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$ does not depend on the quantities of solid calcium carbonate and calcium oxide that are present, but only on the concentration of carbon dioxide gas.

Tutorial 2 Equilibrium Expressions for Heterogeneous Equilibria

In this tutorial, you will write the equilibrium law equations for heterogeneous equilibrium reaction systems. Remember to simplify the equation by omitting the concentrations of any pure solids or liquids.

Sample Problem 1: A Reaction System with a Pure Solid and a Pure Liquid

Solid phosphorous pentachloride, $PCl_5(s)$, can undergo a reversible reaction to form liquid phosphorous trichloride, $PCl_3(l)$, and chlorine gas, $Cl_2(g)$. Write the equilibrium law expression for this heterogeneous equilibrium.

Solution

Write the balanced chemical equation for the reaction. Then, write the equilibrium law equation. Omit the concentrations of the pure solid, $PCl_5(s)$, and the pure liquid, $PCl_3(l)$.

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

$$K = [Cl_2(g)]$$

Sample Problem 2: A Reaction System with Two Pure Solids

When deep blue solid copper(II) sulfate pentahydrate, $CuSO_4{\cdot}5H_2O(s)$, is heated to drive off water vapour, $H_2O(g)$, white solid anhydrous copper(II) sulfate, $CuSO_4(s)$, is formed. The reaction is reversible. Write the balanced chemical equation and the equilibrium law equation for this reaction system.

Solution

$$CuSO_4{\cdot}5H_2O(s) \rightleftharpoons CuSO_4(s) + 5\,H_2O(g)$$

$$K = [H_2O(g)]^5$$

Practice

1. Write the equilibrium law equation for the following heterogeneous equilibrium reaction systems:
   (a) the decomposition of solid ammonium chloride, $NH_4Cl(s)$, to gaseous ammonia, $NH_3(g)$, and hydrogen chloride gas, $HCl(g)$, which is represented by the balanced chemical equation

   $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$ [ans: $K = [NH_3(g)][HCl(g)]$]

   (b) the production of solid sodium carbonate, $Na_2CO_3(s)$; carbon dioxide gas; and water vapour by heating solid sodium hydrogen carbonate, $NaHCO_3(s)$ [ans: $K = [CO_2(g)][H_2O(g)]$]

   (c) the reaction of elemental zinc, $Zn(s)$, in an aqueous solution of copper(II) chloride, $CuCl_2(aq)$ (**Figure 3**) T/I $\left[\text{ans: } K = \dfrac{[ZnCl_2(aq)]}{[CuCl_2(aq)]}\right]$

Figure 3 Stages in the reaction progress of elemental zinc in copper(II) chloride solution

The Magnitude of the Equilibrium Constant, *K*

We have seen that the magnitude of the equilibrium constant, *K*, can vary significantly. What does the magnitude of the equilibrium constant tell us about an equilibrium system? To answer this question, consider the reaction of carbon monoxide gas, $CO(g)$, with oxygen gas, $O_2(g)$, to produce carbon dioxide gas, $CO_2(g)$. This reaction contributes some of the carbon dioxide gas released into the atmosphere by internal combustion engines (**Figure 4**). The balanced equation for this reaction is

$$2\,CO(g) + O_2(g) \rightleftharpoons 2\,CO_2(g)$$

When this reaction is conducted in a sealed vessel at 25 °C, the equilibrium constant is 3.3×10^{91}. Therefore, the equilibrium law equation can be written as

$$K = \frac{[CO_2(g)]^2}{[CO(g)]^2[O_2(g)]} = 3.3 \times 10^{91}$$

Looking at the equilibrium law equation, the numerator includes the equilibrium concentration of the product, and the denominator includes the equilibrium concentrations of the reactants. The very large value of the equilibrium constant indicates that the equilibrium position is significantly toward the right. We can therefore predict that, at equilibrium, the concentration of carbon dioxide gas (the product) will be much greater than the concentrations of carbon monoxide gas and oxygen gas (the reactants).

In contrast, the reversible reaction of nitrogen dioxide gas, $NO_2(g)$, with nitric oxide gas, $NO(g)$, to form gaseous dinitrogen monoxide, $N_2O(g)$, and oxygen,

$$NO_2(g) + NO(g) \rightleftharpoons N_2O(g) + O_2(g)$$

has an equilibrium constant that is close to being equal to 1.

$$K = \frac{[N_2O(g)][O_2(g)]}{[NO_2(g)][NO(g)]} = 0.915$$

Since this equilibrium constant is close to 1, the concentration of the products at equilibrium is approximately equal to the concentration of the reactants.

Finally, let us look at the equilibrium constant for the reversible thermal decomposition of water at 1000 °C. The balanced chemical equation and the equilibrium law equation for this reaction are as follows:

$$2\,H_2O(g) \rightleftharpoons 2\,H_2(g) + O_2(g)$$

$$K = \frac{[H_2(g)]^2[O_2(g)]}{[H_2O(g)]^2} = 7.3 \times 10^{-10}$$

The value of the equilibrium constant for this chemical reaction system is extremely small. This means that the concentration of water vapour is much larger than the concentrations of hydrogen gas and oxygen gas at equilibrium. At this temperature, the equilibrium position is toward the left and highly favours the reactants.

Table 4 summarizes the relationship between the magnitude of the equilibrium constant value and the equilibrium position.

Figure 4 In the open system of this tailpipe, oxygen gas present in the air will react with any carbon monoxide gas emitted by the engine, converting it into carbon dioxide gas that enters the atmosphere.

Table 4 Relationship between the Magnitude of the Equilibrium Constant, *K*, and Equilibrium Position

Magnitude of *K*	Equilibrium position
K >>> 1	far to the right (favours products)
K ≈ 1	equilibrium concentration of the products similar to that of the reactants
K <<< 1	far to the left (favours reactants)

7.2 Review

Summary

- An equilibrium law equation is a mathematical representation of a reversible chemical reaction.
- The equilibrium constant, *K*, is the ratio of the concentrations of the products to the concentrations of the reactants at equilibrium. It is independent of initial concentration, but varies with the temperature at which the reaction occurs.
- In a homogeneous equilibrium, all substances have the same state. In a heterogeneous equilibrium, the substances are in at least two different states.
- The concentrations of solids and pure liquids are not included in the equilibrium expression because they are constant.
- The magnitude of the equilibrium constant, *K*, reflects the equilibrium position.

Questions

1. Write the equilibrium law equation, *K*, for the chemical reaction systems represented by each of the following chemical equations: T/I
   (a) $SiH_2(g) + 2\,Cl_2(g) \rightleftharpoons SiCl_4(g) + H_2(g)$
   (b) $2\,PBr_3(g) + 3\,Cl_2(g) \rightleftharpoons 2\,PCl_3(g) + 3\,Br_2(g)$
   (c) $H_2O(l) \rightleftharpoons H_2O(g)$
   (d) $2\,NaHCO_3(s) \rightleftharpoons Na_2CO_3(s) + CO2(g) + H_2O(g)$
2. At a particular temperature, a 3.0 L flask contains 2.4 mol of chlorine, $Cl_2(g)$; 1.0 mol of nitrosyl chloride, NOCl(g); and 4.5×10^{-3} mol of nitric oxide, NO(g). Calculate the equilibrium constant, given the balanced chemical equation
   $$2\,NOCl(g) \rightleftharpoons 2\,NO(g) + Cl_2(g)$$ T/I
3. An equilibrium mixture contains 1.0 mol of iron metal, Fe(s); 1.0×10^{-3} mol of oxygen gas; and 2.0 mol of solid iron(III) oxide, $Fe_2O_3(s)$, in a 2.0 L container. The reaction is represented by the balanced chemical equation
   $$4\,Fe(s) + 3\,O_2(g) \rightleftharpoons 2\,Fe_2O_3(s)$$
   Calculate the value of *K* for this reaction. T/I
4. Use a graphic organizer to summarize what you have learned about homogeneous and heterogeneous equilibria and how their equilibrium law equations are written. K/U C
5. Liquid methanol, $CH_3OH(l)$, is an important solvent in industry and can also be used as a fuel. Methanol may be synthesized in the gas state from carbon monoxide gas, CO(g), and hydrogen gas, $H_2(g)$, in a chemical reaction system that comes to equilibrium in a closed vessel at 225 °C. T/I A C
   (a) Write the balanced chemical equation for the formation of methanol gas from carbon monoxide gas and hydrogen gas at 225 °C.
   (b) The equilibrium constant, *K*, for this reaction at 225 °C is 6.3×10^{-3}. Predict the relative concentration of the methanol gas when equilibrium is established.
   (c) Based on your answer for (b), predict whether this equilibrium system could generate large quantities of methanol. Explain your answer.
   (d) What would be the equilibrium constant for the reverse reaction?
6. Coal, C(s), can be burned directly as a source of energy or can be converted to carbon monoxide gas, CO(g), which can also be used as a fuel. The conversion process involves reacting coal with carbon dioxide gas to produce carbon monoxide gas. T/I C A
   (a) Write the balanced chemical equation for the reaction.
   (b) Write the equilibrium constant equation.
   (c) The value of the equilibrium constant for this chemical reaction system is 1.6×10^{-2} at 25 °C and 1.7×10^{2} at 1000 °C. At which temperature will more carbon monoxide gas be produced? Explain your reasoning.
   (d) What are some advantages of a gaseous fuel over a solid fuel?
   (e) What safety issues are involved in working with carbon monoxide gas at high temperatures?