Learning Objectives

• To solve quantitative problems involving chemical equilibriums.

There are two fundamental kinds of equilibrium problems:

1. those in which we are given the concentrations of the reactants and the products at equilibrium (or, more often, information that allows us to calculate these concentrations), and we are asked to calculate the equilibrium constant for the reaction; and

2. those in which we are given the equilibrium constant and the initial concentrations of reactants, and we are asked to calculate the concentration of one or more substances at equilibrium. In this section, we describe methods for solving both kinds of problems.

Calculating an Equilibrium Constant from Equilibrium Concentrations

We saw previously the equilibrium constant for the decomposition of \(\ce{CaCO3(s)}\) to \(\ce{CaO(s)}\) and \(\ce{CO2(g)}\) is \(K = \ce{[CO2]}\). At 800°C, the concentration of \(\ce{CO2}\) in equilibrium with solid \(\ce{CaCO3}\) and \(\ce{CaO}\) is \(2.5 \times 10^{-3}\; M\). Thus \(K\) at 800°C is \(2.5 \times 10^{-3}\). (Remember that equilibrium constants are unitless.)

A more complex example of this type of problem is the conversion of n-butane, an additive used to increase the volatility of gasoline, into isobutane (2-methylpropane).

\[\text{n-butane} \rightleftharpoons \text{isobutane}\]

This reaction can be written as follows:

\[\ce{n-butane_{(g)} \rightleftharpoons isobutane_{(g)}} \label{Eq1}\]

and the equilibrium constant

\[K = \frac{[\text{isobutane}]}{[\text{n-butane}]} \label{Eq2}\]

At equilibrium, a mixture of n-butane and isobutane at room temperature was found to contain 0.041 M isobutane and 0.016 M n-butane. Substituting these concentrations into the equilibrium constant expression,

\[K = \frac{0.041}{0.016} = 2.6 \label{Eq2}\]
Thus the equilibrium constant for the reaction as written is 2.6.

Example (PageIndex(1))

The reaction between gaseous sulfur dioxide and oxygen is a key step in the industrial synthesis of sulfuric acid:

\[
\text{\ce{2SO2(g) + O2(g) <=> 2SO3(g)}}
\]

A mixture of \(\text{\ce{SO_2}}\) and \(\text{\ce{O_2}}\) was maintained at 800 K until the system reached equilibrium. The equilibrium mixture contained

- \(5.0 \times 10^{-2}\; \text{M}\; \text{\ce{SO_3}}\),
- \(3.5 \times 10^{-3}\; \text{M}\; \text{\ce{O_2}}\), and
- \(3.0 \times 10^{-3}\; \text{M}\; \text{\ce{SO_2}}\).

Calculate \(K\) and \(K_p\) at this temperature.

**Given:** balanced equilibrium equation and composition of equilibrium mixture

**Asked for:** equilibrium constant

**Strategy**

Write the equilibrium constant expression for the reaction. Then substitute the appropriate equilibrium concentrations into this equation to obtain \(K\).

**SOLUTION**

Substituting the appropriate equilibrium concentrations into the equilibrium constant expression,

\[
K = \dfrac{[\text{\ce{SO_3}}]^2}{[\text{\ce{SO_2}}]^2[\text{\ce{O_2}}]} = \dfrac{(5.0 \times 10^{-2})^2}{(3.0 \times 10^{-3})^2(3.5 \times 10^{-3})} = 7.9 \times 10^4
\]

To solve for \(K_p\), we use the relationship between \(K\) and \(K_p\), where \(\Delta n = 2 - 3 = -1\):

\[
K_p = K(RT)^\Delta n = 7.9 \times 10^4 [(0.08206; \text{L-atm/mol-K})(800 \text{ K})]^{-1} = 1.2 \times 10^3
\]

Exercise (PageIndex(1))

Hydrogen gas and iodine react to form hydrogen iodide via the reaction

\[
\text{\ce{H2(g) + I2(g) <=> 2HI(g)}}
\]

A mixture of \(\text{\ce{H_2}}\) and \(\text{\ce{I_2}}\) was maintained at 740 K until the system reached equilibrium. The equilibrium mixture contained
• \(1.37 \times 10^{-2}\; M\; HI\),
• \(6.47 \times 10^{-3}\; M\; H_2\), and
• \(5.94 \times 10^{-4}\; M\; I_2\).

Calculate \(K\) and \(K_p\) for this reaction.

**Answer:**

• \(K = 48.8\)
• \(K_p = 48.8\)

Chemists are not often given the concentrations of all the substances, and they are not likely to measure the equilibrium concentrations of all the relevant substances for a particular system. In such cases, we can obtain the equilibrium concentrations from the initial concentrations of the reactants and the balanced chemical equation for the reaction, as long as the equilibrium concentration of one of the substances is known. Example \(\PageIndex{2}\) shows one way to do this.

**Example \(\PageIndex{2}\)**

A 1.00 mol sample of \(\text{NOCl}\) was placed in a 2.00 L reactor and heated to 227°C until the system reached equilibrium. The contents of the reactor were then analyzed and found to contain 0.056 mol of \(\text{Cl}_2\). Calculate \(K\) at this temperature. The equation for the decomposition of \(\text{NOCl}\) to \(\text{NO}\) and \(\text{Cl}_2\) is as follows:

\[
\ce{2 NOCl(g) <=> 2NO(g) + Cl2(g)}
\]

**Given:** balanced equilibrium equation, amount of reactant, volume, and amount of one product at equilibrium

**Asked for:** \(K\)

**Strategy:**

A. Write the equilibrium constant expression for the reaction. Construct a table showing the initial concentrations, the changes in concentrations, and the final concentrations (as initial concentrations plus changes in concentrations).

B. Calculate all possible initial concentrations from the data given and insert them in the table.

C. Use the coefficients in the balanced chemical equation to obtain the changes in concentration of all other substances in the reaction. Insert those concentration changes in the table.

D. Obtain the final concentrations by summing the columns. Calculate the equilibrium constant for the reaction.

**SOLUTION**

A The first step in any such problem is to balance the chemical equation for the reaction (if it is not already balanced) and use it to derive the equilibrium constant expression. In this case, the equation is already balanced, and the equilibrium constant expression is as follows:

\[
K=\dfrac{[\text{NO}]^2[\text{Cl}_2]}{[\text{NOCl}]^2}
\]

To obtain the concentrations of \(\text{NOCl}\), \(\text{NO}\), and \(\text{Cl}_2\) at equilibrium, we construct a table showing
what is known and what needs to be calculated. We begin by writing the balanced chemical equation at the top of the table, followed by three lines corresponding to the initial concentrations, the changes in concentrations required to get from the initial to the final state, and the final concentrations.

\[
\text{ICE} \quad \text{ \([\text{NOCl}]\)} \quad \text{\([\text{NO}]\)} \quad \text{\([\text{Cl}_2]\)}
\]

<table>
<thead>
<tr>
<th>Initial</th>
<th>Change</th>
<th>Final</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.500</td>
<td>0</td>
<td>0.028</td>
</tr>
</tbody>
</table>

B Initially, the system contains 1.00 mol of \([\text{NOCl}]\) in a 2.00 L container. Thus \([\text{NOCl}]_i = 1.00 \text{ mol}/2.00 \text{ L} = 0.500 \text{ M}\). The initial concentrations of \([\text{NO}]\) and \([\text{Cl}_2]\) are 0 (M) because initially no products are present. Moreover, we are told that at equilibrium the system contains 0.056 mol of \([\text{Cl}_2]\) in a 2.00 L container, so \([\text{Cl}_2]_f = 0.056 \text{ mol}/ 2.00 \text{ L} = 0.028 \text{ M}\). We insert these values into the following table:

\[
\text{ICE} \quad \text{\([\text{NOCl}]\)} \quad \text{\([\text{NO}]\)} \quad \text{\([\text{Cl}_2]\)}
\]

<table>
<thead>
<tr>
<th>Initial</th>
<th>Change</th>
<th>Final</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.500</td>
<td>0</td>
<td>0.028</td>
</tr>
</tbody>
</table>

C We use the stoichiometric relationships given in the balanced chemical equation to find the change in the concentration of \([\text{Cl}_2]\), the substance for which initial and final concentrations are known:

\[
\Delta[\text{Cl}_2] = 0.028 \text{ M}_{(\text{final})} - 0.00 \text{ M}_{(\text{initial})} = +0.028 \text{ M}
\]

According to the coefficients in the balanced chemical equation, 2 mol of \([\text{NO}]\) are produced for every 1 mol of \([\text{Cl}_2]\), so the change in the \([\text{NO}]\) concentration is as follows:

\[
\Delta[\text{NO}] = \left(\text{frac}{0.028}{\text{L}}\right) \left(\text{frac}{2 \text{ mol}}{1 \text{ mol}}\right) = 0.056 \text{ M}
\]

Similarly, 2 mol of \([\text{NOCl}]\) are consumed for every 1 mol of \([\text{Cl}_2]\) produced, so the change in the \([\text{NOCl}]\) concentration is as follows:

\[
\Delta[\text{NOCl}] = \left(\text{frac}{0.028}{\text{L}}\right) \left(\text{frac}{-1 \text{ mol}}{1 \text{ mol}}\right) = -0.056 \text{ M}
\]

We insert these values into our table:

\[
\text{2 NOCl} \Leftrightarrow 2\text{NO} + \text{Cl}_2
\]
\[ \begin{array}{|c|c|c|c|} 
\hline
\text{ICE} & \text{\([\text{NOCl}]\)} & \text{\([\text{NO}]\)} & \text{\([\text{Cl}_2]\)} \\
\hline
\text{Initial} & 0.500 & 0 & 0 \\
\text{Change} & -0.056 & +0.056 & +0.028 \\
\text{Final} & 0.444 & 0.056 & 0.028 \\
\hline
\end{array} \]

We sum the numbers in the \([\text{NOCl}]\) and \([\text{NO}]\) columns to obtain the final concentrations of \([\text{NO}]\) and \([\text{NOCl}]\):

\[ \text{\([\text{NO}] \_f = 0.000 \text{ M} + 0.056 \text{ M} = 0.056 \text{ M} \)} \]

\[ \text{\([\text{NOCl}] \_f = 0.500 \text{ M} + (-0.056 \text{ M}) = 0.444 \text{ M} \)} \]

We can now complete the table:

\[ \begin{array}{|c|c|c|c|} 
\hline
\text{ICE} & \text{\([\text{NOCl}]\)} & \text{\([\text{NO}]\)} & \text{\([\text{Cl}_2]\)} \\
\hline
\text{initial} & 0.500 & 0 & 0 \\
\text{change} & -0.056 & +0.056 & +0.028 \\
\text{final} & 0.444 & 0.056 & 0.028 \\
\hline
\end{array} \]

We can now calculate the equilibrium constant for the reaction:

\[ K = \frac{[\text{NO}]^2[\text{Cl}_2]}{[\text{NOCl}]^2} = \frac{(0.056)^2(0.028)}{(0.444)^2} = 4.5 \times 10^{-4} \]

Exercise

The German chemist Fritz Haber (1868–1934; Nobel Prize in Chemistry 1918) was able to synthesize ammonia (\([\text{NH}_3]\)) by reacting \([0.1248 \text{ M \text{H}_2}]\) and \([0.0416 \text{ M \text{N}_2}]\) at about 500°C. At equilibrium, the mixture contained 0.00272 M \([\text{NH}_3]\). What is \(K\) for the reaction

\[ \text{\([\text{N}_2 + 3\text{H}_2 \leftrightarrow 2\text{NH}_3]\)} \]

at this temperature? What is \(K_p\)?

Answer:

- \(K = 0.105\)
- \(K_p = 2.61 \times 10^{-5}\)
The original laboratory apparatus designed by Fritz Haber and Robert Le Rossignol in 1908 for synthesizing ammonia from its elements. A metal catalyst bed, where ammonia was produced, is in the large cylinder at the left. The Haber-Bosch process used for the industrial production of ammonia uses essentially the same process and components but on a much larger scale. Unfortunately, Haber’s process enabled Germany to prolong World War I when German supplies of nitrogen compounds, which were used for explosives, had been exhausted in 1914.

Calculating Equilibrium Concentrations from the Equilibrium Constant

To describe how to calculate equilibrium concentrations from an equilibrium constant, we first consider a system that contains only a single product and a single reactant, the conversion of n-butane to isobutane (Equation \ref{Eq1}), for which $K = 2.6$ at 25°C. If we begin with a 1.00 M sample of n-butane, we can determine the concentration of n-butane and isobutane at equilibrium by constructing a table showing what is known and what needs to be calculated, just as we did in Example \PageIndex{2}).

\[
\text{n-butane}(g) \rightleftharpoons \text{isobutane}(g)
\]

<table>
<thead>
<tr>
<th>\text{ICE}</th>
<th>\text{n-butane}</th>
<th>\text{isobutane}</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>\text{n-butane} _ _ _ (g)</td>
<td>\text{isobutane} _ _ _ (g)</td>
</tr>
<tr>
<td>Change</td>
<td>Δ\text{n-butane} = −x</td>
<td>Δ\text{isobutane} = +x</td>
</tr>
<tr>
<td>Final</td>
<td>\text{n-butane} _ _ _ (g)</td>
<td>\text{isobutane} _ _ _ (g)</td>
</tr>
</tbody>
</table>

The initial concentrations of the reactant and product are both known: \([\text{n-butane}]_i = 1.00 \text{ M}\) and \([\text{isobutane}]_i = 0 \text{ M}\). We need to calculate the equilibrium concentrations of both n-butane and isobutane. Because it is generally difficult to calculate final concentrations directly, we focus on the change in the concentrations of the substances between the initial and the final (equilibrium) conditions. If, for example, we define the change in the concentration of isobutane (Δ\text{isobutane}) as +x, then the change in the concentration of n-butane is Δ\text{n-butane} = −x. This is because the balanced chemical equation for the reaction tells us that 1 mol of n-butane is consumed for every 1 mol of isobutane produced. We can then express the final concentrations in terms of the initial concentrations and the changes they have...
undergone.

\[ \text{n-butane}(g) \rightleftharpoons \text{isobutane}(g) \]

**ICE**

<table>
<thead>
<tr>
<th></th>
<th>([\text{n-butane}]_{(g)})</th>
<th>([\text{isobutane}]_{(g)})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>1.00</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>(-x)</td>
<td>(+x)</td>
</tr>
<tr>
<td>Final</td>
<td>((1.00 - x))</td>
<td>((0 + x) = x)</td>
</tr>
</tbody>
</table>

Substituting the expressions for the final concentrations of n-butane and isobutane from the table into the equilibrium equation,

\[ K = \frac{[\text{isobutane}]}{[\text{n-butane}]} = \frac{x}{1.00 - x} = 2.6 \]

Rearranging and solving for \(x\),

\[ \begin{align*}
  x &= 2.6 - 2.6x \\
  2.6x + x &= 2.6 \\
  x &= 0.72
\end{align*} \]

We obtain the final concentrations by substituting this \(x\) value into the expressions for the final concentrations of n-butane and isobutane listed in the table:

\[ \begin{align*}
  [\text{n-butane}]_f &= (1.00 - x) M \\
  &= (1.00 - 0.72) M \\
  &= 0.28 \; M \\
  [\text{isobutane}]_f &= (0.00 + x) M \\
  &= (0.00 + 0.72) M \\
  &= 0.72 \; M
\end{align*} \]

We can check the results by substituting them back into the equilibrium constant expression to see whether they give the same \(K\) that we used in the calculation:

\[ K = \frac{0.72 \; \cancel{M}}{0.28 \; \cancel{M}} = 2.6 \]

This is the same \(K\) we were given, so we can be confident of our results.

Example \(\PageIndex{3}\) illustrates a common type of equilibrium problem that you are likely to encounter.

Example \(\PageIndex{3}\): water–gas shift

The water–gas shift reaction is important in several chemical processes, such as the production of \(\text{H}_2\) for fuel cells. This reaction can be written as follows:

\[ \text{H}_2(g) + \text{CO}_2(g) \rightleftharpoons \text{H}_2\text{O}(g) + \text{CO}(g) \]

\(K = 0.106\) at 700 K. If a mixture of gases that initially contains 0.0150 M \(\text{H}_2\) and 0.0150 M \(\text{CO}_2\) is allowed to equilibrate at 700 K, what are the final concentrations of all substances present?
**Given:** balanced equilibrium equation, \( (K) \), and initial concentrations

**Asked for:** final concentrations

**Strategy:**

A. Construct a table showing what is known and what needs to be calculated. Define \( \Delta x \) as the change in the concentration of one substance. Then use the reaction stoichiometry to express the changes in the concentrations of the other substances in terms of \( x \). From the values in the table, calculate the final concentrations.

B. Write the equilibrium equation for the reaction. Substitute appropriate values from the ICE table to obtain \( x \).

C. Calculate the final concentrations of all species present. Check your answers by substituting these values into the equilibrium constant expression to obtain \( K \).

**Solution**

A. The initial concentrations of the reactants are \( [\ce{H_2}]_i = [\ce{CO_2}]_i = 0.0150 \; \text{M} \). Just as before, we will focus on the change in the concentrations of the various substances between the initial and final states. If we define the change in the concentration of \( [\ce{H_2O}]_i \) as \( \Delta x \), then \( \Delta [\ce{H_2O}] = +x \). We can use the stoichiometry of the reaction to express the changes in the concentrations of the other substances in terms of \( x \). For example, 1 mol of \( [\ce{CO}]_i \) is produced for every 1 mol of \( [\ce{H_2O}]_i \), so the change in the \( [\ce{CO}] \) concentration can be expressed as \( \Delta [\ce{CO}] = +x \). Similarly, for every 1 mol of \( [\ce{H_2O}]_i \) produced, 1 mol each of \( [\ce{H_2}]_i \) and \( [\ce{CO_2}]_i \) are consumed, so the change in the concentration of the reactants is \( \Delta [\ce{H_2}] = \Delta [\ce{CO_2}] = -x \).

We enter the values in the following table and calculate the final concentrations.

\[
\begin{array}{ccccc}
\text{ICE} & [\ce{H_2}] & [\ce{CO_2}] & [\ce{H_2O}] & [\ce{CO}] \\
\text{Initial} & 0.0150 & 0.0150 & 0 & 0 \\
\text{Change} & -x & -x & +x & +x \\
\text{Final} & (0.0150 - x) & (0.0150 - x) & x & x
\end{array}
\]

B. We can now use the equilibrium equation and the given \( (K) \) to solve for \( x \):

\[
K = \frac{[\ce{H_2O}] [\ce{CO}]}{[\ce{H_2}] [\ce{CO_2}]} = \frac{x^2}{(0.0150 - x)^2} = 0.106
\]

We could solve this equation with the quadratic formula, but it is far easier to solve for \( x \) by recognizing that the left side of the equation is a perfect square; that is,

\[
\frac{x^2}{(0.0150 - x)^2} = \left( \frac{x}{0.0150 - x} \right)^2 = 0.106
\]

Taking the square root of the middle and right terms,

\[
\begin{align*}
\frac{x}{0.0150 - x} &= \pm \sqrt{0.106} \\
\therefore x &= \frac{0.0150 \pm \sqrt{0.106}}{2}
\end{align*}
\]

We could simplify this expression to find the value of \( x \).
The final concentrations of all species in the reaction mixture are as follows:

- \([H_2]_f = [H_2]_i + \Delta [H_2] = (0.0150 - 0.00369) \; M = 0.0113 \; M\)
- \([CO_2]_f = [CO_2]_i + \Delta [CO_2] = (0.0150 - 0.00369) \; M = 0.0113 \; M\)
- \([H_2O]_f = [H_2O]_i + \Delta [H_2O] = (0 + 0.00369) \; M = 0.00369 \; M\)
- \([CO]_f = [CO]_i + \Delta [CO] = (0 + 0.00369) \; M = 0.00369 \; M\)

We can check our work by inserting the calculated values back into the equilibrium constant expression:

\[
K = \dfrac{[H_2O][CO]}{[H_2][CO_2]} = \dfrac{(0.00369)^2}{(0.0113)^2} = 0.107
\]

To two significant figures, this \(K\) is the same as the value given in the problem, so our answer is confirmed.

Exercise

Hydrogen gas reacts with iodine vapor to give hydrogen iodide according to the following chemical equation:

\[
\ce{H2(g) + I2(g) <=> 2HI(g)}
\]

\(K = 54\) at 425°C. If 0.172 M \(H_2\) and \(I_2\) are injected into a reactor and maintained at 425°C until the system equilibrates, what is the final concentration of each substance in the reaction mixture?

Answer:

- \([HI]_f = 0.270 \; M\)
- \([H_2]_f = [I_2]_f = 0.037 \; M\)

In Example 3, the initial concentrations of the reactants were the same, which gave us an equation that was a perfect square and simplified our calculations. Often, however, the initial concentrations of the reactants are not the same, and/or one or more of the products may be present when the reaction starts. Under these conditions, there is usually no way to simplify the problem, and we must determine the equilibrium concentrations with other means. Such a case is described in Example 4.

Example

In the water–gas shift reaction shown in Example 3, a sample containing 0.632 M \(\ce{CO2}\) and 0.570 M \(\ce{H_2}\) is allowed to equilibrate at 700 K. At this temperature, \(K = 0.106\). What is the composition of the reaction mixture at equilibrium?

Given: balanced equilibrium equation, concentrations of reactants, and \(K\)

Asked for: composition of reaction mixture at equilibrium

Strategy:

A. Write the equilibrium equation. Construct a table showing the initial concentrations of all substances in the mixture.
Complete the table showing the changes in the concentrations (x) and the final concentrations.

B. Write the equilibrium constant expression for the reaction. Substitute the known K value and the final concentrations to solve for x.

C. Calculate the final concentration of each substance in the reaction mixture. Check your answers by substituting these values into the equilibrium constant expression to obtain K.

Solution

A \(|\text{CO}_2|_i = 0.632\; M|\) and \(|\text{H}_2|_i = 0.570\; M|\). Again, x is defined as the change in the concentration of \(|\text{H}_2O|\): \(\Delta|\text{H}_2O| = +x|\). Because 1 mol of \(|\text{CO}|\) is produced for every 1 mol of \(|\text{H}_2O|\), the change in the concentration of \(|\text{CO}|\) is the same as the change in the concentration of \(|\text{H}_2O|\), so \(\Delta|\text{CO}| = +x|\). Similarly, because 1 mol each of \(|\text{H}_2|\) and \(|\text{CO}_2|\) are consumed for every 1 mol of \(|\text{H}_2O|\) produced, \(\Delta|\text{H}_2| = \Delta|\text{CO}_2| = -x|\). The final concentrations are the sums of the initial concentrations and the changes in concentrations at equilibrium.

\[
\text{H}_2(g) + \text{CO}_2(g) \rightleftharpoons \text{H}_2O(g) + \text{CO}(g)
\]

| \(\text{ICE}\) | \(|\text{H}_2(g)|\) | \(|\text{CO}_2(g)|\) | \(|\text{H}_2O(g)|\) | \(|\text{CO} (g)|\) |
|---|---|---|---|---|
| Initial | 0.570 | 0.632 | 0 | 0 |
| Change | -x | -x | +x | +x |
| Final | (0.570 - x) | (0.632 - x) | x | x |

B We can now use the equilibrium equation and the known \(|K|\) value to solve for \(|x|\):

\[
K = \frac{|\text{H}_2O| |\text{CO}|}{|\text{H}_2| |\text{CO}_2|} = \frac{x^2}{(0.570-x)(0.632-x)} = 0.106
\]

In contrast to Example \(|\text{PageIndex}(3)|\), however, there is no obvious way to simplify this expression. Thus we must expand the expression and multiply both sides by the denominator:

\[
x^2 = 0.106(0.360 - 1.20x + x^2)
\]

Collecting terms on one side of the equation,

\[
0.894x^2 + 0.127x - 0.0382 = 0
\]

This equation can be solved using the quadratic formula:

\[
x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-0.127 \pm \sqrt{(0.127)^2 - 4(0.894)(-0.0382)}}{2(0.894)}
\]

so

\[
x = 0.148 \text{ and } -0.290
\]
Only the answer with the positive value has any physical significance, so \(\Delta[H_2O] = \Delta[CO] = +0.148 \text{ M}\), and \(\Delta[H_2] = \Delta[CO_2] = -0.148 \text{ M}\).

C The final concentrations of all species in the reaction mixture are as follows:

- \([H_2]_f = [H_2]_i + \Delta[H_2] = 0.570 \text{ M} - 0.148 \text{ M} = 0.422 \text{ M}\)
- \([CO_2]_f = [CO_2]_i + \Delta[CO_2] = 0.632 \text{ M} - 0.148 \text{ M} = 0.484 \text{ M}\)
- \([H_2O]_f = [H_2O]_i + \Delta[H_2O] = 0 \text{ M} + 0.148 \text{ M} = 0.148 \text{ M}\)
- \([CO]_f = [CO]_i + \Delta[CO] = 0 \text{ M} + 0.148 \text{ M} = 0.148 \text{ M}\)

We can check our work by substituting these values into the equilibrium constant expression:

\[
\begin{align*}
K &= \dfrac{[H_2O][CO]}{[H_2][CO_2]} \\
&= \dfrac{(0.148)^2}{(0.422)(0.484)} \\
&= 0.107
\end{align*}
\]

Because \(K\) is essentially the same as the value given in the problem, our calculations are confirmed.

Exercise \(\PageIndex{4}\)

The exercise in Example \(\PageIndex{1}\) showed the reaction of hydrogen and iodine vapor to form hydrogen iodide, for which \(K = 54\) at 425°C. If a sample containing 0.200 M \([H_2]\) and 0.0450 M \([I_2]\) is allowed to equilibrate at 425°C, what is the final concentration of each substance in the reaction mixture?

Answer:

- \([HI]_f = 0.0882 \text{ M}\)
- \([H_2]_f = 0.156 \text{ M}\)
- \([I_2]_f = 9.2 \times 10^{-4} \text{ M}\)

In many situations it is not necessary to solve a quadratic (or higher-order) equation. Most of these cases involve reactions for which the equilibrium constant is either very small (\(K < 10^{-3}\)) or very large (\(K > 10^3\)), which means that the change in the concentration (defined as \(x\)) is essentially negligible compared with the initial concentration of a substance. Knowing this simplifies the calculations dramatically, as illustrated in Example \(\PageIndex{5}\).

Example \(\PageIndex{5}\)

Atmospheric nitrogen and oxygen react to form nitric oxide:

\[
N_2(g) + O_2(g) \rightleftharpoons 2NO(g)
\]

with \(K_p = 2.0 \times 10^{-31}\) at 25°C.

What is the partial pressure of \(\ce{NO}\) in equilibrium with \(\ce{N_2}\) and \(\ce{O_2}\) in the atmosphere (at 1 atm, \(P_{\ce{N_2}} = 0.78\); atm) and \(P_{\ce{O_2}} = 0.21\); atm)?

**Given**: balanced equilibrium equation and values of \(K_p\), \(P_{\ce{N_2}}\), and \(P_{\ce{O_2}}\)
Asked for: partial pressure of NO

Strategy:

A. Construct a table and enter the initial partial pressures, the changes in the partial pressures that occur during the course of the reaction, and the final partial pressures of all substances.

B. Write the equilibrium equation for the reaction. Then substitute values from the table to solve for the change in concentration (x).

C. Calculate the partial pressure of $\ce{NO}$. Check your answer by substituting values into the equilibrium equation and solving for $\langle \text{K}_p \rangle$.

Solution

A Because we are given $\langle \text{K}_p \rangle$ and partial pressures are reported in atmospheres, we will use partial pressures. The initial partial pressure of $\ce{O_2}$ is 0.21 atm and that of $\ce{N_2}$ is 0.78 atm. If we define the change in the partial pressure of $\ce{NO}$ as $2x$, then the change in the partial pressure of $\ce{O_2}$ and of $\ce{N_2}$ is $-x$ because 1 mol each of $\ce{N_2}$ and of $\ce{O_2}$ is consumed for every 2 mol of $\ce{NO}$ produced. Each substance has a final partial pressure equal to the sum of the initial pressure and the change in that pressure at equilibrium.

$$\ce{N2(g) + O2(g) <=> 2NO(g)}$$

<table>
<thead>
<tr>
<th>$\langle P_{\ce{N_2}} \rangle$</th>
<th>$\langle P_{\ce{O_2}} \rangle$</th>
<th>$\langle P_{\ce{NO}} \rangle$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.78</td>
<td>0.21</td>
</tr>
<tr>
<td>Change</td>
<td>$-x$</td>
<td>$-x$</td>
</tr>
<tr>
<td>Final</td>
<td>$(0.78-x)$</td>
<td>$(0.21-x)$</td>
</tr>
</tbody>
</table>

B Substituting these values into the equation for the equilibrium constant,

$$\langle \text{K}_p \rangle = \frac{(2x)^2}{(0.78)(0.21)} = 2.0 \times 10^{-31}$$

In principle, we could multiply out the terms in the denominator, rearrange, and solve the resulting quadratic equation. In practice, it is far easier to recognize that an equilibrium constant of this magnitude means that the extent of the reaction will be very small; therefore, the $\langle x \rangle$ value will be negligible compared with the initial concentrations. If this assumption is correct, then to two significant figures, $(0.78-x) = 0.78$ and $(0.21-x) = 0.21$. Substituting these expressions into our original equation,

$$\langle \text{K}_p \rangle = \frac{4x^2}{0.16} = 2.0 \times 10^{-31}$$

$$x^2 = \frac{0.33 \times 10^{-31}}{4}$$

$$x = 9.1 \times 10^{-17}$$

C Substituting this value of x into our expressions for the final partial pressures of the substances,

- $\langle P_{\ce{NO}} \rangle = 2x \text{ atm} = 1.8 \times 10^{-16} \text{ atm}$
- $\langle P_{\ce{N_2}} \rangle = (0.78-x) \text{ atm} = 0.78 \times 0.21 \text{ atm}$
From these calculations, we see that our initial assumption regarding \( x \) was correct: given two significant figures, \( 2.0 \times 10^{-16} \) is certainly negligible compared with 0.78 and 0.21. When can we make such an assumption? As a general rule, if \( x \) is less than about 5% of the total, or \( 10^{-3} > K > 10^{3} \), then the assumption is justified. Otherwise, we must use the quadratic formula or some other approach. The results we have obtained agree with the general observation that toxic \( \text{NO} \), an ingredient of smog, does not form from atmospheric concentrations of \( \text{N}_2 \) and \( \text{O}_2 \) to a substantial degree at 25°C. We can verify our results by substituting them into the original equilibrium equation:

\[
K_p = \dfrac{(P_{\text{NO}})^2}{(P_{\text{N}_2})(P_{\text{O}_2})} = \dfrac{(1.8 \times 10^{-16})^2}{(0.78)(0.21)} = 2.0 \times 10^{-31}
\]

The final \( K_p \) agrees with the value given at the beginning of this example.

Exercise \( \PageIndex{5} \)

Under certain conditions, oxygen will react to form ozone, as shown in the following equation:

\[
\text{O}_2(g) \rightleftharpoons 2\text{O}_3(g)
\]

with \( K_p = 2.5 \times 10^{-59} \) at 25°C. What ozone partial pressure is in equilibrium with oxygen in the atmosphere \( (P_{\text{O}_2} = 0.21 \text{ atm}) \)?

**Answer**

\( 4.8 \times 10^{-31} \text{ atm} \)

Another type of problem that can be simplified by assuming that changes in concentration are negligible is one in which the equilibrium constant is very large \( (K \geq 10^3) \). A large equilibrium constant implies that the reactants are converted almost entirely to products, so we can assume that the reaction proceeds 100% to completion. When we solve this type of problem, we view the system as equilibrating from the products side of the reaction rather than the reactants side. This approach is illustrated in Example \( \PageIndex{6} \).

Example \( \PageIndex{6} \)

The chemical equation for the reaction of hydrogen with ethylene \( (\text{C}_2\text{H}_4) \) to give ethane \( (\text{C}_2\text{H}_6) \) is as follows:

\[
\text{H}_2(g) + \text{C}_2\text{H}_4(g) \rightleftharpoons \text{C}_2\text{H}_6(g)
\]

with \( K = 9.6 \times 10^{18} \) at 25°C. If a mixture of 0.200 M \( \text{H}_2 \) and 0.155 M \( \text{C}_2\text{H}_4 \) is maintained at 25°C in the presence of a powdered nickel catalyst, what is the equilibrium concentration of each substance in the mixture?

**Given**: balanced chemical equation, \( K \), and initial concentrations of reactants

**Asked for**: equilibrium concentrations

**Strategy**:

A. Construct a table showing initial concentrations, concentrations that would be present if the reaction were to go to
A. From the magnitude of the equilibrium constant, we see that the reaction goes essentially to completion. Because the initial concentration of ethylene (0.155 M) is less than the concentration of hydrogen (0.200 M), ethylene is the limiting reactant; that is, no more than 0.155 M ethane can be formed from 0.155 M ethylene. If the reaction were to go to completion, the concentration of ethane would be 0.155 M and the concentration of ethylene would be 0 M. Because the concentration of hydrogen is greater than what is needed for complete reaction, the concentration of unreacted hydrogen in the reaction mixture would be 0.200 M − 0.155 M = 0.045 M. The equilibrium constant for the forward reaction is very large, so the equilibrium constant for the reverse reaction must be very small. The problem then is identical to that in Example \(\PageIndex{5}\). If we define −x as the change in the ethane concentration for the reverse reaction, then the change in the ethylene and hydrogen concentrations is +x. The final equilibrium concentrations are the sums of the concentrations for the forward and reverse reactions.

\[
\begin{array}{ccc}
\text{\[H_2(g)\]+C}_2\text{H}_4(g) & \overset{\text{Ni}}{\rightleftharpoons} & \text{C}_2\text{H}_6(g) \\
\text{\textbf{Initial}} & 0.200 & 0.155 & 0 \\
\text{Assuming 100\% reaction} & 0.045 & 0 & 0.155 \\
\text{Change} & +x & +x & -x \\
\text{Final} & (0.045 + x) & (0 + x) & (0.155 - x)
\end{array}
\]

B. Substituting values into the equilibrium constant expression,

\[
\text{\[K=\dfrac{[\text{C}_2\text{H}_6]}{[\text{H}_2][\text{C}_2\text{H}_4]}=\dfrac{0.155-x}{(0.045+x)x}=9.6 \times 10^{18}\]}
\]

Once again, the magnitude of the equilibrium constant tells us that the equilibrium will lie far to the right as written, so the reverse reaction is negligible. Thus x is likely to be very small compared with either 0.155 M or 0.045 M, and the equation can be simplified \((0.045 + x) = 0.045 \text{ and } (0.155 - x) = 0.155\) as follows:

\[
\text{\[K=\dfrac{0.155}{0.045x} = 9.6 \times 10^{18}\]}
\]

\[
\text{x=3.6 \times 10^{-19}}
\]

C. The small x value indicates that our assumption concerning the reverse reaction is correct, and we can therefore calculate the final concentrations by evaluating the expressions from the last line of the table:

- \([\text{C}_2\text{H}_6]_f = (0.155 - x); \ M = 0.155 \ \text{M}\)
We can verify our calculations by substituting the final concentrations into the equilibrium constant expression:

\[ K = \frac{[C_2H_6]}{[H_2][C_2H_4]} = \frac{0.155}{(0.045)(3.6 \times 10^{-19})} = 9.6 \times 10^{18} \]

This \(K\) value agrees with our initial value at the beginning of the example.

Exercise \(\PageIndex{6}\)

Hydrogen reacts with chlorine gas to form hydrogen chloride:

\[
\ce{ H2(g) + Cl 2(g) <=> 2HCl(g)}
\]

with \(K_p = 4.0 \times 10^{31}\) at 47°C. If a mixture of 0.257 M \(H_2\) and 0.392 M \(Cl_2\) is allowed to equilibrate at 47°C, what is the equilibrium composition of the mixture?

Answer:

- \([H_2]_f = 4.8 \times 10^{-32}\ M\)
- \([Cl_2]_f = 0.135\ M\)
- \([HCl]_f = 0.514\ M\)

Summary

When an equilibrium constant is calculated from equilibrium concentrations, molar concentrations or partial pressures are substituted into the equilibrium constant expression for the reaction. Equilibrium constants can be used to calculate the equilibrium concentrations of reactants and products by using the quantities or concentrations of the reactants, the stoichiometry of the balanced chemical equation for the reaction, and a tabular format to obtain the final concentrations of all species at equilibrium.

Key Takeaway

Various methods can be used to solve the two fundamental types of equilibrium problems: (1) those in which we calculate the concentrations of reactants and products at equilibrium and (2) those in which we use the equilibrium constant and the initial concentrations of reactants to determine the composition of the equilibrium mixture.

Conceptual Problems

1. Describe how to determine the magnitude of the equilibrium constant for a reaction when not all concentrations of the substances are known.
2. Calculations involving systems with very small or very large equilibrium constants can be dramatically simplified by making certain assumptions about the concentrations of products and reactants. What are these assumptions?
when \(K\) is (a) very large and (b) very small? Illustrate this technique using the system \(A+2B \rightleftharpoons C\) for which you are to calculate the concentration of the product at equilibrium starting with only A and B. Under what circumstances should simplifying assumptions not be used?

**Numerical Problems**

1. In the equilibrium reaction \(A+B \rightleftharpoons C\), what happens to \(K\) if the concentrations of the reactants are doubled? tripled? Can the same be said about the equilibrium reaction \(2A \rightleftharpoons B+C\)?

2. The following table shows the reported values of the equilibrium \(P\) for the reaction \(Ag_2O\text{,}(s) \rightleftharpoons 2Ag\text{,}(s)+12O\text{,}(g)\), for which \(\Delta H^\circ = 31 kJ/mol\). Are these data consistent with what you would expect to occur? Why or why not?

<table>
<thead>
<tr>
<th>T (°C)</th>
<th>(P\text{,(O}_2\text{)}): (mmHg)</th>
</tr>
</thead>
<tbody>
<tr>
<td>150</td>
<td>182</td>
</tr>
<tr>
<td>184</td>
<td>143</td>
</tr>
<tr>
<td>191</td>
<td>126</td>
</tr>
</tbody>
</table>

3. Given the equilibrium system \(N_2O_4\text{,(g)} \rightleftharpoons 2NO\text{,(g)}\), what happens to \(K_p\) if the initial pressure of \(N_2O_4\text{(4g)\text{,)}}\) is doubled? If \(K_p\) is \(1.7 \times 10^4\) at 2300°C, and the system initially contains 100% \(N_2O_4\text{(4g)\text{,)}}\) at a pressure of \(2.6 \times 10^2\) atm, what is the equilibrium pressure of each component?

4. At 430°C, 4.20 mol of HI in a 9.60 L reaction vessel reaches equilibrium according to the following equation: \(H\text{,(g)}+I\text{,(g)} \rightleftharpoons 2HI\text{(g)}\). At equilibrium, \([H_2] = 0.047; \text{M}\) and \([HI] = 0.345; \text{M}\). What are \(K\) and \(K_p\) for this reaction?

5. Methanol, a liquid used as an automobile fuel additive, is commercially produced from carbon monoxide and hydrogen at 300°C according to the following reaction: \(CO\text{,(g)}+2H\text{,(g)} \rightleftharpoons CH_3OH\text{(g)}\) with \(K_p = 1.3 \times 10^4\text{, at 1 atm}\). If 56.0 g of \(CO\text{)}\) is mixed with excess hydrogen in a 250 mL flask at this temperature, and the hydrogen pressure is continuously maintained at 100 atm, what would be the maximum percent yield of methanol? What pressure of hydrogen would be required to obtain a minimum yield of methanol of 95% under these conditions?

6. Starting with pure A, if the total equilibrium pressure is 0.969 atm for the reaction \(A(s) \rightleftharpoons 2B_{(g)}+C_{(g)}\), what is \(K_p\)?

7. The decomposition of ammonium carbamate to \(NH_3\text{)}\) and \(CO_2\text{)}\) at 40°C is written as \(NH_4CO_2NH_2\text{(s)} \rightleftharpoons 2NH\text{,(g)}+CO_2\text{(g)}\). If the initial concentration of \(NH_4CO_2NH_2\text{(s)}\) is 0.242 atm, what is the equilibrium partial pressure of \(CO_2\text{(g)}\)? What is the total gas pressure of the system? What is \(K_p\)?

8. At 375 \(K\), \(K_p\) for the reaction \(SO_2Cl_2\text{(g)} \rightleftharpoons SO_2\text{(g)}+Cl_2\text{(g)}\) is 2.4, with pressures expressed in atmospheres. At 303 \(K\), \(K_p\) is \(2.9 \times 10^2\text{, atm}.\)
   a. What is \(K\) for the reaction at each temperature?
   b. If a sample at 375 K has 0.100 M \(Cl_2\text{(l)}\) and 0.200 M \(SO_2\text{(l)}\) at equilibrium, what is the concentration of \(SO_2Cl_2\text{(g)}\)?
   c. If the sample given in part b is cooled to 303 K, what is the pressure inside the bulb?

9. For the gas-phase reaction \(aA \rightleftharpoons bB\text{(l)}\), show that \(K_p = K\text{(RT)}^a\text{(Δn)}\) assuming ideal gas behavior.
10. For the gas-phase reaction \( \text{I}_2 \rightleftharpoons 2\text{I} \), show that the total pressure is related to the equilibrium pressure by the following equation: \( P_T = \sqrt{K_p P_{\text{I}_2}} + P_{\text{I}_2} \)

11. Experimental data on the system \( \text{Br}_2(l) \rightleftharpoons \text{Br}_2(aq) \) are given in the following table. Graph \( [\text{Br}_2] \) versus moles of \( [\text{Br}_2(l)] \) present, then write the equilibrium constant expression and determine \( K \).

<table>
<thead>
<tr>
<th>Grams ( [\text{Br}_2(l)] ) in 100 mL Water</th>
<th>( [\text{Br}_2] ) (M)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.0</td>
<td>0.0626</td>
</tr>
<tr>
<td>2.5</td>
<td>0.156</td>
</tr>
<tr>
<td>3.0</td>
<td>0.188</td>
</tr>
<tr>
<td>4.0</td>
<td>0.219</td>
</tr>
<tr>
<td>4.5</td>
<td>0.219</td>
</tr>
</tbody>
</table>

12. Data accumulated for the reaction \( \text{n-butane(g)} \rightleftharpoons \text{isobutane(g)} \) at equilibrium are shown in the following table. What is the equilibrium constant for this conversion? If 1 mol of \( \text{n-butane} \) is allowed to equilibrate under the same reaction conditions, what is the final number of moles of \( \text{n-butane} \) and \( \text{isobutane} \)?

<table>
<thead>
<tr>
<th>Moles ( \text{n-butane} )</th>
<th>Moles ( \text{Isobutane} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.5</td>
<td>1.25</td>
</tr>
<tr>
<td>1.0</td>
<td>2.5</td>
</tr>
<tr>
<td>1.50</td>
<td>3.75</td>
</tr>
</tbody>
</table>

13. Solid ammonium carbamate \( (\text{NH}_4\text{CO}_2\text{NH}_2(\text{s})) \) dissociates completely to ammonia and carbon dioxide when it vaporizes: \( \text{NH}_4\text{CO}_2\text{NH}_2(\text{s}) \rightleftharpoons 2\text{NH}_3(\text{g})+\text{CO}_2(\text{g}) \) At 25°C, the total pressure of the gases in equilibrium with the solid is 0.116 atm. What is the equilibrium partial pressure of each gas? What is \( (K_p) \)? If the concentration of \( (\text{CO}_2) \) is doubled and then equilibrates to its initial equilibrium partial pressure +x atm, what change in the \( (\text{NH}_3) \) concentration is necessary for the system to restore equilibrium?

14. The equilibrium constant for the reaction \( \text{COCl}_2(\text{g}) \rightleftharpoons \text{CO}(\text{g})+\text{Cl}_2(\text{g}) \) is \( K_p = 2.2 \times 10^{-10} \) at 100°C. If the initial concentration of \( (\text{COCl}_2) \) is \( 3.05 \times 10^{-3} \) M, what is the partial pressure of each gas at equilibrium at 100°C? What assumption can be made to simplify your calculations?

15. Aqueous dilution of \( (\text{IO}_4^-) \) results in the following reaction: \( \text{IO}_4^-+(\text{aq})+2\text{H}_2\text{O}-(\text{l}) \rightleftharpoons \text{H}_4\text{IO}_6^-+(\text{aq}) \) with \( K = 3.5 \times 10^2 \). If you begin with 50 mL of a 0.896 M Solution of \( (\text{IO}_4^-) \) that is diluted to 250 mL with water, how many moles of \( (\text{H}_4\text{IO}_6^-) \) are formed at equilibrium?

16. Iodine and bromine react to form \( (\text{IBr}) \), which then sublimes. At 184.4°C, the overall reaction proceeds according to the following equation: \( \text{I}_2(\text{g})+\text{Br}_2(\text{g}) \rightleftharpoons 2\text{IBr}(\text{g}) \) with \( K_p = 1.2 \times 10^2 \). If you begin the reaction with 7.4 g of \( (\text{I}_2) \) vapor and 6.3 g of \( (\text{Br}_2) \) vapor in a 1.00 L container, what is the concentration of \( (\text{IBr}) \) at equilibrium? What is the partial pressure of each gas at equilibrium? What is the total pressure of the system?

17. For the reaction \( \text{C}_-(\text{s})+12\text{N}_2(\text{g})+\text{I}_2(\text{g}) \rightleftharpoons \text{CH}_3\text{NH}_2(\text{g}) \) with \( K = 1.8 \times 10^{-6} \). If you begin the reaction with 1.0 mol of \( (\text{N}_2) \), 2.0 mol of \( (\text{H}_2) \), and sufficient \( (\text{C}_-(\text{s})) \) in a 2.00 L container, what are the concentrations of \( (\text{N}_2) \) and \( (\text{CH}_3\text{NH}_2) \) at equilibrium? What happens to \( (K) \) if the
concentration of $H_2$ is doubled?