Chapter 7 – Chemical Systems and Equilibrium

Dynamic Equilibrium in Chemical Systems

- Many closed systems that do not appear to be changing are in fact undergoing reaction
  - The macroscopic properties are constant (what we can see), but changes may be going on that balance each other out
  - Such a system is at a dynamic equilibrium
- Dynamic equilibrium – a balance between forward and reverse processes occurring at the same rate
  - This is shown using a double arrow (combining a forward and reverse arrow), indicating that the reaction is going in both the forward and reverse directions at the same time
  
  Ex. \[ \text{CO}_2(g) \rightleftharpoons \text{CO}_2(aq) \]

- Forward reaction – in an equilibrium reaction, the left-to-right reaction
- Reverse reaction – in an equilibrium reaction, the right-to-left reaction

Solubility Equilibrium

- A dynamic equilibrium between a solute and a solvent in a saturated solution in a closed system
- In a saturated solution, the concentration of the dissolved solute is constant, but according to the dynamic equilibrium concept, the rate of the dissolving process is equal to the rate of the crystallizing process
- Most substances can dissolve in a solvent to a certain extent
- The initial dissolution (dissolving) of a substance in a solvent is thought to be the result of collisions among solvent particles and solute particles
- Once this solution is saturated, particles continue to collide:
  - When particles of the solvent collide with the solute crystal, solute particles break off and enter the dissolved state
  - When particles of the dissolved solute collide with the crystal, they may form ionic bonds and crystallize out of solution
  - Therefore, solute particles continue to dissolve into the solvent, while other solute particles crystallize out of the solvent
  - A dynamic equilibrium is when these two processes occur at the same rate that there is no observable change in either the concentration of the solute particles in solution or in the quantity of solid present
- Early in the dissolving process, the number of particles entering the dissolved state far exceeds the number that crystallize
- Nearing equilibrium, the rates of dissolution and crystallization approach one another
- At equilibrium, the rate of dissolution equals the rate of crystallization

- Ex. NaCl dissolves in water

Figure 2

(a) When the solute is first added, many more ions dissociate from the crystal than crystallize onto it.
(b) As more ions come into solution, more ions also crystallize.
(c) At solubility equilibrium, solute ions dissolve and crystallize at the same rate.
Ex. radioactive iodine is added to a saturated solution of normal iodine

A mixture exhibiting solubility equilibrium must contain both dissolved and undissolved solute at the same time
- This state can be established by starting with a solute and adding it to a solvent
- Ex. Add calcium sulfate to water in a large enough quantity that not all of it will dissolve
  \[ \text{CaSO}_4(s) \rightleftharpoons \text{Ca}^{2+} + \text{SO}_4^{2-} \]
- Ex. Combine two solutions, one with a very high concentration of calcium ions and the other with a very high concentration of sulfate ions, initially causing solid crystals to form at a higher rate than those crystals would dissolve
  - This causes \( \text{CaSO}_4 \) to precipitate, which will continue until the rate become equal and a dynamic equilibrium is established
- What is important is that, at equilibrium, the rates of the forward and reverse reactions are equal – how the equilibrium is established is irrelevant

Phase Equilibrium
- A dynamic equilibrium between different physical states of a pure substance in a closed system
- Ex. evaporation/condensation equilibrium
  \[ \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{O}(g) \]
  - When a liquid is placed in a closed container, initially only evaporation will occur (some molecules in the liquid gain enough energy in collisions to leave the surface of the liquid phase and move into the space above the liquid to start a second phase, the gas phase)
  - Molecules in the gas phase increase
  - Molecules in the gas phase can still collide with liquid molecules, lose energy, and join the condensed phase
  - In time, equilibrium will be established when the rate at which molecules are evaporating and condensing become equal
  - The amount (and pressure) of the substance in the gas phase remains constant
The tendency of a liquid to evaporate is greater at higher temperatures, so concentration (and pressure) of vapour is greater if equilibrium is established at higher temperatures.

- Note – equilibrium cannot be established in an open container because molecules that leave the surface of the liquid escape from the system and don’t return to the liquid phase.

- **Ex. solid/liquid equilibrium**
  \[ H_2O_{(l)} \rightleftharpoons H_2O_{(g)} \quad t = 0^\circ C \]
  - Established at melting/freezing point.
  - Can be seen in a frozen lake (at 0°C) - equilibrium occurs and keeps the surface temperature at exactly 0°C as long as there is some ice and water present.

### Chemical Reaction Equilibrium

- A dynamic equilibrium between reactants and products of a chemical reaction in a closed system.
- So far, we have dealt mainly with **quantitative reactions** (a reaction in which virtually all the limiting reagent is consumed and essentially all the reactants change into products if enough time passes).
- However, many reactions are not quantitative and when occurring in a closed container, both reactants and products are usually present after the reaction appears to have stopped.
- The final state of a chemical system at equilibrium can be explained as a competition between collisions of reactants to form products and collisions of products to form reactants.
  - This required the system to be closed so reactants and products cannot escape.
- Allowing reactions to reach equilibrium limits the amount of product produced.
- **Continuous processing** – carried out in an open container, products formed are continuously removed while fresh reactants are fed into the system.
  - Ex. producing calcium oxide (lime) - lime and carbon dioxide are continuously removed and calcium carbonate is added.
- **Batch processing** – done in a closed container, fixed amounts of each reactant are put in a closed vessel and the products are removed when the reaction is complete.
  - Ex. making popcorn in a bag in the microwave.

**Ex. Decomposition of dinitrogen tetroxide into nitrogen dioxide**

\[ N_2O_4(g) \rightleftharpoons 2NO_2(g) \]

- \( N_2O_4(g) \) is placed in a closed container and decomposes quickly at first.
  - \([N_2O_4(g)]\) (a colourless gas) decreases.
  - \([NO_2(g)]\) (an orange-brown gas) increases.
  - Eventually, concentrations change more slowly.
  - Equilibrium is reached when concentrations remain constant.
Although the equilibrium established above was created by placing $N_2O_4(g)$ into a closed container, the same equilibrium can occur if $NO_2(g)$ was the starting material.

For a given overall system composition, the same equilibrium concentrations are reached whether equilibrium is approached in the forward or reverse direction.

Percent Reaction at Chemical Equilibrium
- Chemists have studied the reaction of hydrogen gas and iodine gas extensively, because the molecules are simple and the reaction takes place entirely in gas phase
- The reaction is rapid at first, then reaches equilibrium (the dark purple colour of the iodine vapour fades then becomes constant).

The reaction is:

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

Three experiments involving this system are performed and the results recorded in the table below (one in which hydrogen and iodine are mixed, one in which hydrogen, iodine and hydrogen iodide are mixed, and one in which only hydrogen iodide is present initially).

<table>
<thead>
<tr>
<th>System</th>
<th>Initial system concentrations (mmol/L)</th>
<th>Equilibrium system concentrations (mmol/L)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>$1.00$</td>
<td>$1.00$</td>
</tr>
<tr>
<td>2</td>
<td>$0.50$</td>
<td>$0.50$</td>
</tr>
<tr>
<td>3</td>
<td>$0$</td>
<td>$0$</td>
</tr>
</tbody>
</table>
Chemists can use this table to describe the state of equilibrium in two ways
- In terms of percent reaction
- In terms of an equilibrium constant

**Percent Reaction** – the yield of *product* measured at equilibrium compared with the maximum possible yield of *product*
- Can be found for all three systems using the formula:

\[
\text{percent reaction} = \frac{\text{actual product yield}}{\text{theoretical product yield}} \times 100\%
\]
- Actual product yield is found by experiment
- Theoretical product yield can be calculated stoichiometrically, as if the reaction were quantitative and forward

Ex. System 1 – actual product yield is 1.56 mmol/L, while we would expect 2.00 mmol/L of HI\(_{(g)}\) product based on the mole ratios of the chemical equation and the concentrations we started with

\[
\text{percent reaction} = \frac{\text{actual product yield}}{\text{theoretical product yield}} \times 100\%
= \frac{1.56 \text{ mmol/L}}{2.00 \text{ mmol/L}} \times 100\%
= 78.0\%
\]

Percent reaction provides an easily understood way to communicate relative amounts of chemicals present in equilibrium systems.

To communicate the extent of a reaction, the percent reaction is usually written above the equilibrium arrow

\[
\text{H}_2\text{(g)} + \text{I}_2\text{(g)} \rightleftharpoons \text{HI}_\text{(g)}
\]

### Table 4 Percent Reaction of the Hydrogen–Iodine System at 448°C

<table>
<thead>
<tr>
<th>System</th>
<th>Equilibrium [HI]* (mmol/L)</th>
<th>Maximum possible [HI]* (mmol/L)</th>
<th>Percent reaction (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1.56</td>
<td>2.00</td>
<td>78.0</td>
</tr>
<tr>
<td>2</td>
<td>2.10</td>
<td>2.70</td>
<td>77.8</td>
</tr>
<tr>
<td>3</td>
<td>2.50</td>
<td>3.20</td>
<td>78.1</td>
</tr>
</tbody>
</table>

*Square brackets [ ] indicate molar concentration.
Extent of Chemical Reactions

- All reactions are now considered to be reversible – so the question is now to what extent will they go one way or the other.
- Reactions fall loosely into three categories:
  - Reactions that favour reactants very strongly – the percent reaction is much less than 1%. In these reactions, mixing reactants has no observable result.
  - Reactions that favour products very strongly, where the percent reaction is more than 99%. These reactions are observed to be complete (quantitative). These reactions are generally written with a single arrow to indicate that the effect of the reverse reaction is negligible.
  - Reactions that achieve noticeable equilibrium conditions – the percent reaction lies somewhere between 1% and 99%. In these cases, significant amounts of both reactants and products are always present in a mixture in a closed system. If the percent reaction is less than 50%, reactants are favoured; if greater than 50%, products are favoured.

<table>
<thead>
<tr>
<th>Description of Chemical Reactions at Equilibrium</th>
<th>Position of equilibrium</th>
</tr>
</thead>
<tbody>
<tr>
<td>no reaction (NR)</td>
<td>&lt;1%</td>
</tr>
<tr>
<td>reactants favoured</td>
<td>&lt;50%</td>
</tr>
<tr>
<td>products favoured</td>
<td>&gt;50%</td>
</tr>
<tr>
<td>quantitative</td>
<td>&gt;99% or 100%</td>
</tr>
</tbody>
</table>

Stoichiometric Calculations with Equilibrium Reactions

- When reversible reactions achieve equilibrium before all of the reactants become products, the stoichiometry requires more thought than simple mole ratios.
- An ICE table provides a convenient way to organize the information needed to solve a stoichiometric problem involving an equilibrium system.
- I – *initial* concentration of reactants and products (before reaction)
- C – *change* in the concentrations of reactants and products between the start and the point at which equilibrium is achieved.
- E – concentrations of reactants and products at *equilibrium*.
Sample Problem
Ex. 1) Consider the following equation for the formation of hydrogen fluoride from its elements at SATP:

\[ \text{H}_2(g) + \text{F}_2(g) \rightleftharpoons 2\text{HF}(g) \]

If the reaction begins with 1.00 mol/L concentrations of \( \text{H}_2(g) \) and \( \text{F}_2(g) \) and no \( \text{HF}(g) \), calculate the concentrations of \( \text{H}_2(g) \) and \( \text{HF}(g) \) at equilibrium if the equilibrium concentration of \( \text{F}_2(g) \) is measured to be 0.24 mol/L.

Solution

- First, list the information given in the question:

- Now, we can set up an ICE table to organize the information. Make sure to incorporate the balanced chemical reaction in the first row as mole ratios will be important to consider.

- When equilibrium is reached, a certain amount of \( \text{H}_2(g) \) and \( \text{F}_2(g) \) will have changed into \( \text{HF}(g) \). The balanced equation indicates that this change occurs in a 1:1:2 molar ratio (i.e., for every mole of \( \text{H}_2(g) \) and \( \text{F}_2(g) \) that react, two moles of \( \text{HF}(g) \) are created). However, we don’t know what amount of reactants is converted into product.

- So, use the variable \( x \) to represent the changes in concentration of reactants and products, with the coefficients of \( x \) corresponding to the coefficients in the balanced equation.

- The change in \( \text{H}_2(g) \) and \( \text{F}_2(g) \) will be represented by \( -x \) mol/L. The balanced equation indicates that 2 mol of \( \text{HF}(g) \) is produced for every mole of \( \text{H}_2(g) \) and \( \text{F}_2(g) \) that reacts. So the change in concentration of \( \text{HF}(g) \) is represented by \( +2x \) mol/L.

- The final concentrations of \( \text{H}_2(g) \) and \( \text{F}_2(g) \) will be their initial concentrations, 1.00 mol/L, minus \( x \) mol/L. Therefore, we place \( (1.00 - x) \) in the equilibrium concentration row under both \( \text{H}_2(g) \) and \( \text{F}_2(g) \). As there was no \( \text{HF}(g) \) at the beginning of the reaction, the final (equilibrium) concentration of this gas is \( 0 + 2x \), or simply \( 2x \). Therefore, place \( 2x \) in the equilibrium row under \( \text{HF}(g) \).
• Knowing that the equilibrium concentration of $F_2(g)$ is 0.24 mol/L, you can determine the value of $x$.

• Now, use the value of $x$ to calculate the equilibrium concentrations of the other two entities.

• The equilibrium concentrations of $H_2(g)$ and $HF(g)$ are 0.24 mol/L and 1.52 mol/L respectively.
Ex. 2) When ammonia is heated, it decomposes into nitrogen gas and hydrogen gas according to the following equation:

$$2\text{NH}_3(g) \rightleftharpoons \text{N}_2(g) + 3\text{H}_2(g)$$

When 4.0 mol of $\text{NH}_3(g)$ is introduced into a 2.0 L rigid container and heated to a particular temperature, the amount of ammonia changes as shown in the figure shown. Determine the equilibrium concentrations of the other two entities.

**Solution**

- Determine concentrations given by the graph

- Set up an ICE table to incorporate these given values, and indicate the concentration changes

- Determine the value of $x$ using the initial, change, and equilibrium values of $\text{NH}_3(g)$ using the calculated equilibrium concentration for $\text{NH}_3(g)$

- Use the value of $x$ to calculate the equilibrium concentrations of the other two entities
Equilibrium Law in Chemical Reactions

- When considering chemical reactions that reach an equilibrium point, both reactants and products will be in the mixture.
- Repetitive testing of various chemical reactions that reach equilibrium while beginning with different initial concentrations of reactants show that the concentrations of the reactants and products at equilibrium can be related through the equilibrium law:

Consider the reaction:  \( aA + bB \rightleftharpoons cC + dD \)

The Equilibrium Law can be expressed as:

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

where:  
- A, B, C and D are chemical entities in gas or aqueous phases
- \( a, b, c \) and \( d \) are the coefficients of the balanced equation
- \( K \) is the equilibrium constant

- The law of mass action is a relation that states that the value of the equilibrium constant is constant for a given temperature.
- When the concentrations of the reactants and products at equilibrium are arranged according to the equilibrium law expression, they will be equal to the equilibrium constant.
- The equilibrium law describes the behavior of almost all gaseous and aqueous chemical equilibrium.
- The concentrations of the chemical entities inputted into the equilibrium law equation must be expressed in moles per liter (mol/L).

Sample Problems

1. Write equilibrium expressions for the following reversible reactions:
   a. \( 2 \text{NO}_2(g) \rightleftharpoons \text{N}_2\text{O}_4(g) \)
   b. \( \text{N}_2(g) + 3 \text{H}_2(g) \rightleftharpoons 2 \text{NH}_3(g) \)

2. For the reaction,

\( 2 \text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2 \text{SO}_3(g) \)

which of the following is the correct equilibrium expression:

a. \( \frac{[\text{SO}_3]}{[\text{SO}_2][\text{O}_2]} \)

b. \( \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]} \)

c. \( \frac{[\text{SO}_3]^2}{[\text{O}_2]} \)
Kinetics Argument for the Equilibrium Law

- Consider the following example to demonstrate how analyzing rates of reaction can yield the equilibrium law:

\[ \text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g) \]

- Experimental work has shown the rate law equation for the forward reaction is:

\[ r_f = k_f\left[\text{N}_2\text{O}_4(g)\right] \]

while the rate law for the reverse reaction is:

\[ r_r = k_r\left[\text{NO}_2(g)\right]^2 \]

- At equilibrium, the rate of reaction in the forward direction is equal to the rate of reaction for the reverse direction, so:

\[ r_f = r_r \]

\[ k_f\left[\text{N}_2\text{O}_4(g)\right] = k_r\left[\text{NO}_2(g)\right]^2 \]

- Manipulation of the above equation yields:

\[ \frac{k_f}{k_r} = \left[\frac{\text{NO}_2(g)}{\text{N}_2\text{O}_4(g)}\right]^2 \]

- The equilibrium expression can be written as:

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

- Therefore, we can set equation (1) equal to equation (2) so that:

\[ K = \frac{k_f}{k_r} \]

The Magnitude of the Equilibrium Constant

- The magnitude of the rate constant will determine the degree to which a reaction will proceed in the direction specified.

- Consider that the equilibrium constant is a ratio of the rate constant for the forward reaction (making products) to the rate constant for the reverse reaction (making reactants)

<table>
<thead>
<tr>
<th>K</th>
<th>Description</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>&gt;&gt; 1</td>
<td>Products &gt;&gt; Reactants</td>
<td>(2\text{O}_3(g) \leftrightarrow 3\text{O}_2(g)) at 2000 K ( K = 2.0 \times 10^{57})</td>
</tr>
<tr>
<td></td>
<td>The ratio of product concentration to reactant concentration is very large. The equilibrium lies very much to the right, favouring product formation.</td>
<td></td>
</tr>
<tr>
<td>= 1</td>
<td>Products = Reactants</td>
<td>(\text{CO}(g) + \text{H}_2\text{O}(g) \leftrightarrow \text{CO}_2(g) + \text{H}_2(g)) at 700º C ( K = 5.09)</td>
</tr>
<tr>
<td></td>
<td>The ratio of product concentration to reactant concentration is very small. The equilibrium lies very much to the left, favouring reactants.</td>
<td></td>
</tr>
<tr>
<td>&lt;&lt; 1</td>
<td>Products &lt;&lt; Reactants</td>
<td>(\text{N}_2(g) + \text{O}_2(g) \leftrightarrow 2\text{NO}_2(g)) ( K = 1.0 \times 10^{-25})</td>
</tr>
</tbody>
</table>
Sample Problems
1. For each of the following reactions, state whether the value of the equilibrium constant favours the formation of reactants or products.

   a. \( \text{I}_2(g) + \text{Cl}_2(g) \rightleftharpoons 2 \text{ICl}(g) \) \( K_{eq} = 2.0 \times 10^6 \)

   b. \( \text{H}_2(g) + \text{Cl}_2 \rightleftharpoons 2 \text{HCl}(g) \) \( K_{eq} = 1.08 \)

   c. \( \text{I}_2(g) \rightleftharpoons \text{I}(g) + \text{I}(g) \) \( K_{eq} = 3.8 \times 10^{-7} \)

2. The equilibrium constant for the decomposition of molecular chlorine at 298K is \( 1.4 \times 10^{-38} \). Would many chlorine atoms result from the dissociation of the chlorine molecules at this temperature?

- In general, the equilibrium constant of a forward reaction (\( K \)) and the equilibrium constant of the reverse reaction (\( K' \)) are reciprocal quantities
  \( K = \frac{1}{K'} \)

Sample Problems
1. Write the equilibrium expression for the following reaction:

   \( 2 \text{O}_3(g) \rightleftharpoons 3 \text{O}_2(g) \)

   a. If the equilibrium constant for the above reaction is \( 1.0 \times 10^{12} \) at 2300°C, what will be the value of the equilibrium constant at the same temperature, for the following reaction?

   \( 3 \text{O}_2(g) \rightleftharpoons 2 \text{O}_3(g) \)
Depending on the reaction, the equilibrium constant may or may not vary significantly with temperature.

Ex. Production of ammonia from nitrogen and hydrogen gas

- At $T = 25^\circ C$, $K = 4.26 \times 10^8$
- At $T = 300^\circ C$, $K = 1.02 \times 10^{-5}$
- At $T = 400^\circ C$, $K = 8.00 \times 10^{-7}$

**Homogeneous and Heterogeneous Equilibria**

- **Homogeneous** equilibrium involved reactants and products in a single phase
  - Ex. All aqueous or all gas

- **Heterogeneous** equilibrium involved reactants and products in more than one phase
  - Ex. Liquid and gas

Consider the following reaction with solid and gas phases:

$$3\text{Fe(s)} + 4\text{H}_2\text{O(g)} \rightleftharpoons \text{Fe}_3\text{O}_4(s) + 4\text{H}_2(g)$$

The equilibrium law expression is:

$$K = \frac{[\text{Fe}_3\text{O}_4(s)][\text{H}_2(g)]^4}{[\text{Fe(s)}]^3[\text{H}_2\text{O(g)}]^4}$$

But, the concentration of the pure solids (Fe(s) and Fe$_3$O$_4$(s)) do **not** change during the reaction as their densities are fixed - therefore, they do not affect the equilibrium.

As such, the quantities of [Fe(s)] and [Fe$_3$O$_4$(s)] are constant and could be incorporated into the equilibrium constant leaving:

$$K = \frac{[\text{H}_2(g)]^4}{[\text{H}_2\text{O(g)}]^4}$$

Similarly, in heterogeneous reactions that involve pure liquids, the concentration of the pure liquid is constant (fixed density) and can also be incorporated into the equilibrium constant.

The equilibrium constant reported in reference tables for heterogeneous equilibria will already incorporate the densities of pure solids or liquids.

In the case where a heterogeneous reaction involves ions and a pure solid, pure liquid, or gas, then a net ionic equation is developed to eliminate spectator ions and show only those ions altered or involved in the chemical reaction.

**Sample Problem**

1. Write the equilibrium law equation for the reaction of zinc in copper(II) chloride solution.

   - First, write the balanced chemical equation

   - Write a complete ionic equation

   - The spectator ions simply cancel out, resulting in the following net ionic equation:
• Write the equilibrium law equation

• Rearrange the equation to separate the variables from the constants

• The constants are incorporated into the value of $K$, leaving:

Note – the constant concentrations of the solids, as well as the concentration of spectator ions (the chloride ions in this example), don’t appear in the equilibrium expression.

2. Write the equilibrium expressions for each of the following reactions:

   a. $\text{Fe}_3\text{O}_4(s) + \text{H}_2(g) \rightleftharpoons 3 \text{FeO}(s) + \text{H}_2\text{O}(g)$

   b. $\text{Ag}_2\text{S}(s) \rightleftharpoons 2 \text{Ag}^+(aq) + \text{S}^{2-}(aq)$

3. Classify the following equilibria as heterogeneous or homogeneous, and write an equilibrium expression for each.

   a. $\text{NH}_4\text{NO}_2(s) \rightleftharpoons \text{N}_2(g) + 2 \text{H}_2\text{O}(g)$

   b. $\text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{O}(g)$

   c. $\text{SO}_2(g) + \frac{1}{2} \text{O}_2(g) \rightleftharpoons \text{SO}_3(g)$

   d. $\text{S}_8(s) + 8 \text{O}_2(g) \rightleftharpoons 8 \text{SO}_2(g)$
Determining Equilibrium Concentration

- Simple calculations involving the equilibrium law equation requires knowledge of a combination of the concentrations of the reactants and products and the equilibrium constant.
- Ex. Consider the following reaction:
  \[ \text{H}_2(g) + \text{CO}_2(g) \rightleftharpoons \text{H}_2\text{O}(g) + \text{CO}(g) \]

Where the concentrations are given as:
- \([\text{H}_2(g)] = 0.24 \text{ mol/L}\]
- \([\text{CO}_2(g)] = 1.80 \text{ mol/L}\]
- \([\text{H}_2\text{O}(g)] = 0.88 \text{ mol/L}\]
- \([\text{CO}(g)] = 0.88 \text{ mol/L}\]

The equilibrium law expression is:

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

\[
K = \frac{(0.88 \text{ mol/L})(0.88 \text{ mol/L})}{(0.24 \text{ mol/L})(1.80 \text{ mol/L})} = 1.8
\]

- Variations of this question may involve:
  - Determining the concentration of one of the reactants or products (K given along with other reactant and product concentrations)
  - Determining K where balanced chemical equation and other information makes it possible to calculate concentrations