CHAPTER 14 (MOORE)
CHEMICAL EQUILIBRIUM

This chapter deals with chemical equilibrium, or how far chemical reactions proceed. Some reactions convert reactants to products with near 100% efficiency but others do not. Chemical equilibrium is a dynamic process, and occurs when the rates of two opposing processes (reactions) become the same. Unless conditions are somehow changed, the concentrations of both reactants and products remain constant after equilibrium is achieved.

Consider the decomposition reaction in which \( \text{N}_2\text{O}_4 \) \( \rightleftharpoons \) 2 NO\(_2\) \( (g) \)

What happens to the concentration of each, based upon the rate equations, as the reaction proceeds? Assuming reversible elementary processes,

\[
\text{Rate (forward)} = k \text{ (forward)} \times [\text{N}_2\text{O}_4] \quad \text{and} \quad \text{rate (reverse)} = k \text{ (reverse)} \times [\text{NO}_2]^2
\]

As [\( \text{N}_2\text{O}_4 \)] decreases, rate forward decreases, and, as [\( \text{NO}_2 \)] increases, rate reverse increases, but more quickly … why? Let’s look at the graph of concentration vs. time …the concentration lines converge at equilibrium.

We can also graph changes in concentrations with time for

2 HI\( (g) \rightleftharpoons \) H\(_2\)\( (g) \) + I\(_2\)\( (g) \)

(see also Fig. 14.2, Moore, p. 676)

Is there a mathematical relationship between [HI], [H\(_2\)] and [I\(_2\)]?

**Expressing Equilibrium Relationships**

2 HI\( (g) \rightleftharpoons \) H\(_2\)\( (g) \) + I\(_2\)\( (g) \), where  
rate\( \text{(forward)} = k_f[H\text{I}]^2 \) \quad \text{and} \quad \text{rate\( \text{(reverse)} = k_r[H_2][I_2] \)}

or,  
rate\( \text{(forward)} = \text{rate\( \text{(reverse)} \) \Leftrightarrow k_f[H\text{I}]^2 = k_r[H_2][I_2] \)}

and,  
\( (k_f/k_r) = [H_2][I_2]/[HI]^2 \quad \text{constant} = K_C \text{ (the equilibrium constant)} \)

**\( K_C \) is called the “concentration-based” equilibrium constant!** The product concentrations are on top and the reactant concentrations are on the bottom.

Regardless of the starting concentrations, the value of the expression, with product concentrations in the numerator, and reactant concentrations in the denominator, where each concentration is raised to the power of its coefficient in the balanced chemical equation, is equal to a constant.
The Equilibrium Constant Expression (sometimes called the “mass action” expression)

For the general reaction: \( aA + bB \rightarrow gG + hH \), the equilibrium expression is:

\[
K_c = \frac{[G]^g[H]^h}{[A]^a[B]^b}
\]

Products Reactants

The Equilibrium Constant
The equilibrium constant is constant regardless of the initial concentrations of reactants and products, as long as the temperature remains constant.

The concentration equilibrium constant is \( K_C \)

Concentrations of the products appear in the numerator and concentrations of the reactants appear in the denominator. Recall that \([A]\) means molar concentration of \( A \).

Important: the exponents of the concentrations are identical to the stoichiometric coefficients in the chemical equation!!

Example 14.1
If the equilibrium concentrations of \( \text{COCl}_2 \) and \( \text{Cl}_2 \) are the same at 395 \( ^\circ \text{C} \), find the equilibrium concentration of \( \text{CO} \) in the reaction:

\( \text{CO}(g) + \text{Cl}_2(g) \rightleftharpoons \text{COCl}_2(g) \quad K_C = 1.2 \times 10^3 \text{ at 395 } ^\circ \text{C} \)

First, write the \( K_C \) expression, where \( K_C = 1.2 \times 10^3 = [\text{COCl}_2] / [\text{CO}][\text{Cl}_2] \)

Then, \([\text{COCl}_2] = [\text{Cl}_2]\), so substitute in the equilibrium expression, and

\[
K_C = 1.2 \times 10^3 = [\text{Cl}_2] / [\text{CO}][\text{Cl}_2] = 1/[\text{CO}] \quad \Rightarrow \quad [\text{CO}] = 8.3 \times 10^{-4} \text{ M}
\]

Reversing the Direction of the Chemical Equation
How does this affect the form of \( K_C \)? Consider the reaction: \( 2 \text{NO}(g) + \text{O}_2(g) \rightleftharpoons 2 \text{NO}_2(g) \)

Then, \( K_C = [\text{NO}_2]^2 / [\text{NO}]^2 [\text{O}_2] \)

Now consider \( 2 \text{NO}_2(g) \rightleftharpoons 2 \text{NO}(g) + \text{O}_2(g) \), where \( K_C' = [\text{NO}]^2[\text{O}_2] / [\text{NO}_2]^2 \)

You should see that \( K_C' = 1/K_C \)

When the reaction direction is reversed, \( K_C' = 1/K_C \)

Modifying the Chemical Equation: Changing Coefficients
Changing the coefficients in a balanced equation affects the form of the equilibrium expression!
Example 14.2. The equilibrium constant at 718 K for the reaction below is 7.07.

\[ \frac{1}{2} \text{H}_2(g) + \frac{1}{2} \text{I}_2(g) \rightleftharpoons \text{HI}(g) \]

(a) What is the value of \( K_C \) at 718 K for the reaction, \( \text{HI}(g) \rightleftharpoons \frac{1}{2} \text{H}_2(g) + \frac{1}{2} \text{I}_2(g) \)?

**Solution.** Reversing the reaction means taking the reciprocal of \( K_C \): \( K'_C = \frac{1}{7.07} = 0.141 \)

(b) What is the value of \( K_C \) at 718 K for the reaction, \( \text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2 \text{HI}(g) \)?

**Solution.** Doubling the coefficients in the original balanced equation means squaring \( K_C \):

\[ K'_C = (7.07)^2 = 50.0 \]

**Modifying the Chemical Equation: Changing Coefficients**

Another example …

If \( K_C = 1.97 \times 10^{-20} \) for \( \frac{1}{2} \text{N}_2(g) + 3/2 \text{H}_2\text{O}(g) \rightleftharpoons \text{NH}_3(g) + \frac{3}{4} \text{O}_2(g) \) at 900 K, what is \( K_C \) at the *same temperature* for the reaction \( 4 \text{NH}_3(g) + 3 \text{O}_2(g) \rightleftharpoons 2 \text{N}_2(g) + 6 \text{H}_2\text{O}(g) \)?

**Solution.** Reversing the reaction and quadrupling the coefficients means raising all exponents by a factor of 4 and taking the reciprocal of \( K_C \):

\[ K'_C = 1/(1.97 \times 10^{-20})^4 = 6.64 \times 10^{78} \]

**Equilibrium Constants for Overall Reactions**

When reactions occur in two or more steps, we can use the \( K_C \) values for each step to calculate \( K_C \) for the overall process.

Suppose we need: \( \text{N}_2\text{O}(g) + 3/2 \text{O}_2(g) \rightleftharpoons 2 \text{NO}_2(g) \) \( K_C(1) = ?? \) and we’re given:

\[
\begin{align*}
\text{N}_2\text{O}(g) + \frac{1}{2} \text{O}_2(g) & \rightleftharpoons 2 \text{NO}(g) & K_C(2) = 1.7 \times 10^{-13} \\
2 \text{NO}(g) + \text{O}_2(g) & \rightleftharpoons 2 \text{NO}_2(g) & K_C(3) = 4.67 \times 10^{13} 
\end{align*}
\]

*Adding* the given equations gives the desired equation. *Multiplying* the given values of \( K \) gives the equilibrium constant for the overall reaction.

Adding equations 2 and 3 gives: \( \text{N}_2\text{O}(g) + 3/2 \text{O}_2(g) \rightleftharpoons 2 \text{NO}_2(g) \)

Multiplying \( K_C(2) \) by \( K_C(3) \), we get \( K_C(1) = (1.7 \times 10^{-13})(4.67 \times 10^{13}) = 8.0 \)

**Equilibria Involving Gases** - In reactions involving gases, we often use partial pressures in place of molarities. This means that we have to use a *partial pressure equilibrium constant*.

For the reaction, \( a\text{A}(g) + b\text{B}(g) \rightleftharpoons g\text{G}(g) + h\text{H}(g) \), 

\[ K_P = \frac{\{P_G\}^g \{P_H\}^h}{\{P_A\}^a \{P_B\}^b} \]

where the \( P \) values represent partial pressures, typically in atm or mmHg

\( K_C \) and \( K_P \) are related by: 

\[ K_P = K_C \left(\frac{RT}{\Delta n_{\text{gas}}}\right) \]
where $\Delta n(\text{gas})$ is the change in the total number of moles of gas as the reaction occurs in the forward direction. More specifically, $\Delta n(\text{gas}) = \text{mol gaseous products} - \text{mol gaseous reactants}$

**Example 14.3**
Consider the equilibrium between dinitrogen tetroxide and nitrogen dioxide:

$$\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2 \text{NO}_2(\text{g}) \quad K_p = 0.660 \text{ at } 319 \text{ K}$$

(a) What is the value of $K_c$ for this reaction?

(b) What is the value of $K_p$ for the reaction $2 \text{NO}_2(\text{g}) \rightleftharpoons \text{N}_2\text{O}_4(\text{g})$?

(c) If the equilibrium partial pressure of NO$_2$(g) is 0.332 atm, what is the equilibrium partial pressure of N$_2$O$_4$(g)?

**Solution.**  
$$\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2 \text{NO}_2(\text{g}) \quad K_p = 0.660 \text{ atm at } 319 \text{ K}$$

(a) What is the value of $K_c$ for this reaction?

\[K_p = K_c \cdot \frac{(RT)^{\Delta n(\text{gas})}}{K_c} \text{ where } R = 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1} \]
\[
\text{Here, } \Delta n(\text{gas}) = 2 - 1 = +1
\]
\[
K_c = 0.660 \text{ atm}/\left[(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(319 \text{ K})\right]^{+1} = 0.0252 \text{ mol L}^{-1}
\]

(b) What is the value of $K_p$ for the reaction $2 \text{NO}_2(\text{g}) \rightleftharpoons \text{N}_2\text{O}_4(\text{g})$?

Now, the reaction direction is reversed ... recall that $K_p' = 1 / K_p$  
So, $K_p' = 1/(0.660 \text{ atm}) = 1.52 \text{ atm}^{-1}$

(c) If the equilibrium partial pressure of NO$_2$(g) is 0.332 atm, what is the equilibrium partial pressure of N$_2$O$_4$(g)?

Begin by writing the $K_p$ equilibrium expression ...  
\[K_p = \frac{(p\text{NO}_2)^2}{(p\text{N}_2\text{O}_4)} \quad K_p = 0.660 \text{ atm} \]

and, then, substitute for $p(\text{NO}_2) = 0.332 \text{ atm}$

\[K_p = \frac{(0.332 \text{ atm})^2}{(0.660 \text{ atm})} = 0.660 \text{ atm} \]

\[p(\text{N}_2\text{O}_4) = \frac{(0.332 \text{ atm})^2}{(0.660 \text{ atm})} = 0.167 \text{ atm} \]

**Equilibria Involving Pure Solids and Liquids**
The equilibrium constant expression does not include terms for pure solid and liquid phases because their concentrations do not change in a reaction. Although the amounts of pure solid and liquid phases change during a reaction, these phases remain pure and their concentrations do not change.

Example:  
$$\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g}) \quad K_c = ?? \quad K_c = [\text{CO}_2]^1$$
Example 14.4 (see also Moore, Problem-Solving Example 14.1 (b), p. 678)
The reaction of steam and coke (a form of carbon) produces a mixture of carbon monoxide and hydrogen, called water-gas. This reaction has long been used to make combustible gases from coal: C(s) + H₂O(g) ⇌ CO(g) + H₂(g)

Write the equilibrium constant expression for \( K_C \) for this reaction.
\[
K_C = \frac{[CO][H_2]}{[H_2O]}
\]
Why no [C]? Because C is a pure solid, C(s)!

Equilibrium Constants: When Do We Need Them?
A very large numerical value of \( K_c \) or \( K_p \) signifies that a reaction goes (essentially) to completion. A very small numerical value of \( K_c \) or \( K_p \) signifies that the forward reaction, as written, occurs only to a slight extent.

An equilibrium constant expression applies only to a reversible reaction at equilibrium. Although a reaction may be thermodynamically favored, it may be kinetically controlled ... Thermodynamics tells us “it’s possible (or not)”. Kinetics tells us if “it’s practical (or not)”

Example 14.5
Is the reaction CaO(s) + CO₂(g) ⇌ CaCO₃(s) likely to occur to any appreciable extent at 298 K? We know that for CaCO₃(s) ⇌ CaO(s) + CO₂(g) \( K_P = 1.9 \times 10^{-23} \)

What does this tell us? That this reaction is “reactant favored” or that the equilibrium “lies to the left.” Lots of CaCO₃ with very little CO₂ .... where \( K_P = p(CO_2)^1 \). Therefore, this reaction favors the formation of CaCO₃, so the answer is YES!

The Reaction Quotient, \( Q \)
For nonequilibrium conditions, the expression having the same form as \( K_c \) or \( K_p \) is called the reaction quotient, \( Q_c \) or \( Q_p \). The reaction quotient is not constant for a reaction, but is useful for predicting the direction in which a net change must occur to establish equilibrium. To determine the direction of net change, we compare the magnitude of \( Q_c \) to that of \( K_c \). If \( Q_c > K_c \), then numerator is too large relative to the denominator and some product must be converted to reactant (i.e. equilibrium shifts to the left). When \( Q < K \), the denominator of \( Q \) is too big; we have “too much reactants.” When \( Q = K \), equilibrium has been reached. When \( Q > K \), the numerator of \( Q \) is too big; we have “too much products.”

Example 14.6
Predict the direction of net change for Experiment 3 in Table 14.1.\[
[HI] = 1.000 \text{ M; } [H_2] = 0.000 \text{ M; } [I_2] = 0.000 \text{ M}
\]
\[
Q_C = \frac{(0.000)(0.000)(1.000)^2}{1.000} = 0.000 \quad K_C = 1.84 \times 10^{-2}
\]
\( Q_C < K_C \), so equilibrium must shift to the right to form more product(s).

See also Examples and Exercises, Moore, pp. 690-691

Le Châtelier’s Principle
When any change in concentration, temperature, pressure, or volume is imposed on a system at equilibrium, the system responds by attaining a new equilibrium condition that minimizes the impact of the imposed change.
When a system at equilibrium is “stressed,” it responds in such a way that it relieves the stress and attains a new state of equilibrium.

**Changing the Amounts of Reacting Species**
At equilibrium, \[ Q_C = K_c. \] If the concentration of one of the reactants is increased, the denominator of the reaction quotient increases. \( Q_C \) is now less than \( K_c \). This condition is only temporary because the concentrations of all species must change in such a way so as to make \( Q_C = K_c \) again. In order to do this, the concentrations of the products increase; the “equilibrium is shifted to the right.”

**Heterogeneous Equilibria and Le Chatelier’s Principle**
Addition or removal of pure solids or pure liquids from a system at equilibrium does not affect the position of the equilibrium.

**Changing External Pressure or Volume in Gaseous Equilibria**
When the external pressure is increased (or system volume is reduced), an equilibrium shifts in the direction producing the smaller number of moles of gas. When the external pressure is decreased (or the system volume is increased), an equilibrium shifts in the direction producing the larger number of moles of gas.

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

Increasing pressure shifts to the side with fewer molecules of gas, so 2 \( \text{NO}_2 \) combine to form 1 \( \text{N}_2\text{O}_4 \) (equilibrium shifts “to the left”).

If there is *no change in the number of moles of gas* in a reaction, changes in external pressure (or system volume) have *no effect* on an equilibrium.

**Example:** \( \text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2 \text{HI} \quad \Delta n(\text{gas}) = 0 \)
This equilibrium is unaffected by pressure changes.

**Example 14.8**
An equilibrium mixture of \( \text{O}_2(g) \), \( \text{SO}_2(g) \), and \( \text{SO}_3(g) \) is transferred from a 1.00-L flask to a 2.00-L flask. The balanced equation for the reaction is \( 2 \text{SO}_3(g) \rightleftharpoons 2 \text{SO}_2(g) + \text{O}_2(g) \)
In which direction does a net reaction proceed in order to restore equilibrium?

Volume increases (from 1.00 L to 2.00 L), so pressure decreases, favoring more molecules or the side with more moles of gas … 2 mols on left, 3 mols on right, so equilibrium shift is left-to-right.

**Temperature Changes**
Raising the temperature of an equilibrium mixture shifts equilibrium in the direction of the endothermic reaction; lowering the temperature shifts equilibrium in the direction of the exothermic reaction. We can treat heat as a product of an exothermic reaction or as a reactant of an endothermic reaction, and then apply Le Châtelier’s principle.

Heating an endothermic rxn shifts the equilibrium toward more products; heating an exothermic rxn shifts the equilibrium toward more reactants.
Example 14.9
Is the amount of NO(g) formed from given amounts of N\(_2\)(g) and O\(_2\)(g),
N\(_2\)(g) + O\(_2\)(g) \rightleftharpoons 2\text{NO}(g) \quad \Delta H^\circ = +180.5 \text{kJ}
greater at high or at low temperatures?

From (+) sign of \(\Delta H^\circ\) we see that the reaction is endothermic as written. This means we can essentially write heat as a reactant, or, N\(_2\)(g) + O\(_2\)(g) + Heat \rightleftharpoons 2\text{NO}(g)

Adding heat (higher temp.) means equilibrium should shift left to right and more product should form at higher temperatures.

Effect of Catalysts and Catalysis
A catalyst lowers the activation energy of both the forward and the reverse reactions.

\[ \text{\Rightarrow Adding a catalyst does not affect an equilibrium state. A catalyst merely causes equilibrium to be achieved more rapidly.} \]

Equilibrium Calculations
1. Finding \(K_C\) from concentration data
2. Finding one equilibrium concentrations from \(K_C\) and the remaining concentrations
3. Finding equilibrium concentrations using square roots
4. Finding equilibrium concentrations using the quadratic equation

Determining Values of Equilibrium Constants from Experimental Data
When initial amounts and equilibrium amounts of one or more species are given, the amounts of the remaining species in the equilibrium state and the equilibrium concentrations can be calculated.

A useful general approach is to tabulate under the chemical equation:
1. concentrations of substances initially present (I)
2. changes in concentrations that occur (C)
3. the equilibrium concentrations (E)

We use this procedure to make a table called an “ICE” table: Initial / Change /Equilibrium.

Example 14.11
In a 10.0-L vessel at 1000 K, 0.250 mol SO\(_2\) and 0.200 mol O\(_2\) react to form 0.162 mol SO\(_3\) at equilibrium. Calculate \(K_C\), at 1000 K, for the reaction 2 SO\(_2\)(g) + O\(_2\)(g) \rightleftharpoons 2\text{SO}_3(g)

\[
[\text{SO}_2] = (0.250 \text{ mol}/10.0 \text{ L}) = 0.0250 \text{ M} \\
[\text{O}_2] = (0.200 \text{ mol}/10.0 \text{ L}) = 0.0200 \text{ M} \\
[\text{SO}_3] = (0.162 \text{ mol}/10.0 \text{ L}) = 0.0162 \text{ M}
\]

Now set up the ICE table:
Example 14.12

Consider the reaction, \( \text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2 \text{HI}(g) \) \( K_c = 54.3 \) at 698 K

If we start with 0.500 mol \( \text{I}_2(g) \) and 0.500 mol \( \text{H}_2(g) \) in a 5.25-L vessel at 698 K, how many moles of each gas will be present at equilibrium?

Initial \( [\text{H}_2] = (0.500 \text{ mol/5.25 L}) = 0.0952 \text{ M} = [\text{I}_2] \)

<table>
<thead>
<tr>
<th>( [\text{SO}_2] )</th>
<th>( [\text{O}_2] )</th>
<th>( [\text{SO}_3] )</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>0.0250</td>
<td>0.0200</td>
</tr>
<tr>
<td>C</td>
<td>-2x</td>
<td>-x</td>
</tr>
<tr>
<td>E</td>
<td>0.0250 - 2x</td>
<td>0.0200 - x</td>
</tr>
</tbody>
</table>

\[ 0 + 2x = 0.0162, \text{ so } x = 0.0162/2 = 0.00810 \]

\[ [\text{SO}_2] = 0.0250 - 2 (0.00810) = 0.088 \text{ M} \]
\[ [\text{O}_2] = 0.0200 - 0.0081 = 0.0119 \text{ M} \]

\[ K_c = [\text{SO}_3]^2 / ([\text{SO}_2]^2[\text{O}_2]) = (0.0162)^2 / ((0.0088)(0.0119)) = 2.8 \times 10^2 \]

Now, solve for “x” ...

\[ \frac{(2x)^2}{(0.0952 - x)(0.0952 - x)} = 54.3 \]
\[ (2x)^2/(0.0952 - x)^2 = 54.3 \]

Taking the square root of each side, \( (2x)/(0.0952 - x) = (54.3)^{1/2} = 7.368 \)
\[ 2x = (7.368)(0.0952 - x) = 0.702 - 7.368x \]
\[ x = 0.0749 \]

\[ [\text{H}_2] = [\text{I}_2] = 0.0952 - x = 0.0952 - 0.0749 = 0.0203 \text{ M} \]
\[ [\text{HI}] = 2x = 2(0.0749) = 0.150 \text{ M} \]

Now “check” by plugging back into the \( K_c \) expression:
\[ K_c = [\text{HI}]^2 / ([\text{H}_2][\text{I}_2]) = (0.150)^2 / ((0.0203)(0.0203)) \]
\[ K_c = 54.6 \text{ (w/ round off error)} \]
Calculating Equilibrium Quantities from $K_C$ and $K_P$ Values

We can calculate the amount of substances present at equilibrium, starting with only initial reactants, no products and the known value of the equilibrium constant. 
1. Set up an ICE table, using “$x$” to identify the changes in concentration that occur in reaching equilibrium.
2. All concentration changes are related to $x$, and the appropriate terms are substituted into the equilibrium constant expression.
3. Solve the equation for $x$.

Example 14.13

Suppose that in the reaction of Example 14.12, the initial amounts are 0.800 mol H$_2$ and 0.500 mol I$_2$. What will be the amounts of reactants and products when equilibrium is attained?

$$\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2 \text{HI}(\text{g}) \quad K_C = 54.3 \text{ at } 698 \text{ K}$$

If we start with 0.500 mol I$_2$(g) and 0.500 mol H$_2$(g) in a 5.25-L vessel at 698 K, how many moles of each gas will be present at equilibrium? $K_C = 54.3 = [\text{HI}]^2 / ([\text{H}_2][\text{I}_2])$

Example 14.14

Carbon monoxide and chlorine react to form phosgene, COCl$_2$, which is used in the manufacture of pesticides, herbicides, and plastics:

$$\text{CO}(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons \text{COCl}_2(\text{g}) \quad K_C = 1.2 \times 10^3 \text{ at } 668 \text{ K}$$

How much of each substance, in moles, will there be at equilibrium in a reaction mixture that initially has 0.0100 mol CO, 0.0100 mol Cl$_2$, and 0.100 mol COCl$_2$ in a 10.0-L flask?

Initial concentrations:

$[\text{CO}] = 0.0100 \text{ mol/10.0 L} = 0.00100 \text{ M} = [\text{Cl}_2]$  
$[\text{COCl}_2] = 0.100 \text{ mol/10.0 L} = 0.0100 \text{ M}$

Calculate $Q_C$ and compare to $K_C$ to see the direction of the reaction.

$$Q_C = C(\text{COCl}_2) / C(\text{CO})\cdot C(\text{Cl}_2) = (0.01000) / (0.00100)^2 = 1 \times 10^4$$

Since $Q_C >> K_C$, reaction goes to the left (more reactants)

<table>
<thead>
<tr>
<th></th>
<th>[CO]</th>
<th>[Cl$_2$]</th>
<th>[COCl$_2$]</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>0.00100</td>
<td>0.00100</td>
<td>0.0100</td>
</tr>
<tr>
<td>C</td>
<td>+ x</td>
<td>+ x</td>
<td>- x</td>
</tr>
<tr>
<td>E</td>
<td>0.00100 + x</td>
<td>0.00100 + x</td>
<td>0.0100 - x</td>
</tr>
</tbody>
</table>

Then, $K_C = [\text{COCl}_2] / ([\text{CO}] [\text{Cl}_2]) = (0.0100 - x) / ([0.00100 + x])(0.00100 + x) = 1.2 \times 10^3$
\[(0.0100 - x) = (1.2 \times 10^3)(0.00100 + x)^2 = 1.2 + 2.4x + (1.2 \times 10^3)x^2\]

Collecting terms, \(1.2 \times 10^3 x^2 + 3.4x - 0.0088 \equiv 0\)

Solve as a quadratic: \(a = 1.2 \times 10^3; b = 3.4; c = -0.0088\)

Positive root: \(x = 0.0016\)

So, \([CO] = [Cl_2] = 0.00100 + x = 0.00100 + 0.0016 = 0.0026\) M

and \([COCl_2] = 0.0100 - x = 0.0100 - 0.0016 = 0.0084\) M

Then, \(\text{mol CO} = \text{mol Cl}_2 = (0.0026\ \text{mol/L})(10.0\ \text{L}) = 0.026\ \text{mol}\)

and \(\text{mol COCl}_2 = (0.0084\ \text{mol/L})(10.0\ \text{L}) = 0.084\ \text{mol}\)