14.26 3.00×10^{-2} mol of COCl_2 (phosgene gas) was placed into 1.50 L container. It was heated to 800K and at equilibrium, the pressure of CO was found to be 0.497 atm. Calculate the equilibrium constant K_p for the reaction, \text{CO(g) + Cl}_2 (g) \rightleftharpoons \text{COCl}_2 (g)

a) Since, you will need to solve for K_p, you need partial pressures. Thus, first calculate the initial partial pressure of COCl_2. (Hint: Use PV=nRT)

\[ P_{\text{COCl}_2} = \frac{nRT}{V} = \frac{(3.00 \times 10^{-2} \text{ mol})(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(800 \text{ K})}{(1.50 \text{ L})} = 1.31 \text{ atm} \]

b) You are starting out with pure COCl_2. Given the reaction as written, which way must the reaction shift to come to equilibrium? \underline{Left}

c) Now, plug in all partial pressure values that you know into the ICE chart. Then, fully fill out chart by determining any missing values. \textit{(Make sure to shift rxn in the correct direction.)}

\[
\begin{array}{ccc}
\text{CO(g)} & + & \text{Cl}_2(\text{g}) \rightleftharpoons \text{COCl}_2(\text{g}) \\
\text{Initial (atm):} & 0 & 0 & \textbf{1.31} \\
\text{Change (atm):} & +0.497 & +0.497 & -0.497 \\
\text{Equil (atm):} & \textbf{0.497} & 0.497 & 0.81
\end{array}
\]

d) Calculate K_p.

\[ K_p = \frac{P_{\text{COCl}_2}}{P_{\text{CO}}P_{\text{Cl}_2}} = \frac{0.81 \text{ atm}}{(0.497 \text{ atm})^2} = 3.3 \]

14.40 For the synthesis of ammonia, \text{N}_2 (g) + 3 \text{H}_2 (g) \rightleftharpoons 2 \text{NH}_3 (g), the equilibrium constant K_c is 1.2 at 375°C. If one starts with a reaction mixture consisting of the following concentrations: [H_2]= 0.76 M, [N_2] = 0.60 M, [NH_3] = 0.48 M, which gases will have increased in concentration and which will have decreased in concentration when the mixture comes to equilibrium?

a) This question is basically asking you which way the reaction will shift to reach equilibrium. Thus, solve for “Q” by using the given initial concentrations.

\[ Q_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(0.48M)^2}{(0.60M)(0.76M)^3} = 0.87 \]

b) Now compare Q to the K_c and determine which direction the reaction must shift to reach equilibrium and thus how the concentrations of each gas will change.

\( Q_c \quad < \quad K_c \quad \text{eqil. Thus, rxn shifts to the right} \quad \text{[NH}_3\text{] will increase} \quad \text{[N}_2\text{] and [H}_2\text{] will decrease} \)

\( 0.87 \quad < \quad 1.2 \quad \text{to get more products.} \)
14.42 At 1000 K, a sample of pure NO\textsubscript{2} gas decomposes as shown: \(2 \text{NO}_2 (g) \rightleftharpoons 2 \text{NO} (g) + \text{O}_2 (g)\) 

The equilibrium constant \(K_P\) is 158. Analysis shows that the partial pressure of \(\text{O}_2\) is 0.25 atm at equilibrium. Calculate the pressures of \(\text{NO}\) and \(\text{NO}_2\) in the mixture at equilibrium.

**Overall logic:** This question does not involve any initial pressure values, thus you don’t need to use an ICE chart. Instead, you need to find two equilibrium pressure values. Obviously, you will be plugging values into the \(K_P\) expression, but you have two variables right now. Thus, you need to find another way to figure out either the \(\text{NO}\) or the \(\text{NO}_2\) pressure value.

a) Determine the partial pressure of NO. *(Hint: You started with pure \text{NO}_2 gas. You know the equilibrium partial pressure of \text{O}_2. Look at the balanced equation.)*

At equilibrium, \(P_{\text{NO}} = 2 \text{P}_{\text{O}_2}\) (Balanced eq)

Thus, \(P_{\text{NO}} = (2 \times 0.25 \text{ atm}) = 0.50 \text{ atm.}\

b) Now, solve for the equilibrium partial pressure of \(\text{NO}_2\) using the \(K_P\) expression. *(Please, write the \(K_P\) expression first and then plug in numbers.)*

\[ P_{\text{NO}_2} = \sqrt{\frac{(P_{\text{NO}})^2 \text{P}_{\text{O}_2}}{K_P}} = \sqrt{\frac{(0.50 \text{ atm})^2 \times 0.25 \text{ atm}}{158}} = \boxed{0.020 \text{ atm}}\]

c) You have already answered the question as written, but let’s expand the question and now figure out the initial partial pressure of \(\text{NO}_2\). To do so, you can use an ICE chart. Please fill out the ICE chart below and determine the initial partial pressure of \(\text{NO}_2\).

<table>
<thead>
<tr>
<th>(2 \text{NO}_2 (g))</th>
<th>(\leftarrow)</th>
<th>(2 \text{NO} (g))</th>
<th>+</th>
<th>(\text{O}_2 (g))</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial (atm):</td>
<td>0.52</td>
<td>0</td>
<td>0</td>
<td></td>
</tr>
<tr>
<td>Change (atm):</td>
<td>-0.50</td>
<td>+0.50</td>
<td>+0.25</td>
<td></td>
</tr>
<tr>
<td>Equil. (atm):</td>
<td>0.02</td>
<td>0.50</td>
<td>0.25</td>
<td></td>
</tr>
</tbody>
</table>

14.43 The equilibrium constant \(K_C\) for the reaction, \(\text{H}_2(g) + \text{Br}_2(g) \rightleftharpoons 2 \text{HBr(g)}\) is \(2.18 \times 10^6\) at 730°C. Starting with 3.20 mol HBr in a 12.0-L reaction vessel, calculate the equilibrium concentrations of all components.

a) First calculate the initial concentration of HBr.

\([\text{HBr}]_0 = \frac{3.20 \text{ mol}}{12.0 \text{ L}} = 0.267 \text{ M}\)

b) You are starting with pure HBr. Given the reaction as written, which way must the reaction shift to come to equilibrium? **Left**

c) Completely fill out the ICE chart using x’s where necessary.

<table>
<thead>
<tr>
<th>(\text{H}_2 (g))</th>
<th>+</th>
<th>(\text{Br}_2(g))</th>
<th>(\rightleftharpoons)</th>
<th>(2 \text{HBr (g)})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial (M):</td>
<td>0</td>
<td>0</td>
<td>0.267</td>
<td></td>
</tr>
<tr>
<td>Change (M):</td>
<td>+x</td>
<td>+x</td>
<td>-2x</td>
<td></td>
</tr>
<tr>
<td>Equil (M):</td>
<td>x</td>
<td>x</td>
<td>(0.267 - 2x)</td>
<td></td>
</tr>
</tbody>
</table>

d) Write the \(K_C\) expression. Plug in values & solve for x. *(Note-no quadratic eq. is needed b/c it’s a perfect square.)*

\[ K_C = \frac{[\text{HBr}]^2}{[\text{H}_2][\text{Br}_2]} = \frac{(0.267 - 2x)^2}{x^2} = 2.18 \times 10^6 \]

Taking the square root of both sides we obtain:

\[ \frac{0.267 - 2x}{x} = 1.48 \times 10^3 \quad \text{and} \quad x = 1.80 \times 10^{-4} \]

e) What are the equilibrium concentrations of all components in the reaction?

\([\text{H}_2] = [\text{Br}_2] = 1.80 \times 10^{-4} \text{ M}\)

\([\text{HBr}] = 0.267 - 2(1.80 \times 10^{-4}) = 0.267 \text{ M}\)
14.44 The dissociation of molecular iodine into iodine atoms is represented as $\text{I}_2(g) \rightleftharpoons 2\text{I}(g)$. At 1000 K, the equilibrium constant $K_C$ is $3.80 \times 10^{-5}$. Suppose you start with 0.0456 mole of $\text{I}_2$ in 2.30 L flask at 1000 K. What are the concentrations of gases at equilibrium?

a) What is the initial concentration of $\text{I}_2$? $0.0456 \text{ mol}/2.30 \text{ L} = 0.0198 \text{ M}$.

b) Fill in the ICE chart using “x’s” where necessary.

$$
\begin{array}{c|c|c}
 & \text{I}_2 (g) & \rightleftharpoons & 2 \text{I} (g) \\
\text{Initial (M):} & 0.0198 & & 0 \\
\text{Change (M):} & -x & & +2x \\
\text{Equil (M):} & 0.0198 - x & & 2x \\
\end{array}
$$

c) Write the $K_C$ expression and plug in equilibrium values from the ICE chart. Then, solve for “x” (without using the quadratic equation) by using the approximation method.

$$
K_C = \frac{[\text{I}]^2}{[\text{I}_2]} \ ; \quad 3.80 \times 10^{-5} = \frac{(2x)^2}{(0.0198 - x)}
$$

$$
3.80 \times 10^{-5} \approx \frac{(2x)^2}{(0.0198)} \\
7.52 \times 10^{-7} = 4x^2 \\
1.88 \times 10^{-7} = x^2 \\
4.34 \times 10^{-4} \text{ M} = x
$$

d) Check to see if this is an acceptable approximation (See if less than 5% of initial $\text{I}_2$ was lost.)

$$
\% \text{I}_2 \text{ lost} = \frac{\text{I}_2 \text{ lost}}{\text{initial } \text{I}_2} \times 100 = \frac{4.34 \times 10^{-4}}{0.0198} \times 100 = 2.2\% \text{ lost from initial}. \text{ Yes, acceptable.}
$$

e) What are the concentrations of all gases at equilibrium?

$$
[\text{I}] = 2x = (2)(4.34 \times 10^{-4} \text{ M}) = 8.68 \times 10^{-4} \text{ M} \\
[\text{I}_2] = (0.0198 - x) = [0.0198 - (4.34 \times 10^{-4})] \text{ M} = 0.0194 \text{ M}
$$

14.48 The equilibrium constant $K_c$ for the reaction, $\text{H}_2(g) + \text{CO}_2(g) \rightleftharpoons \text{H}_2\text{O}(g) + \text{CO}(g)$, is 4.2 at 1650°C. Initially, 0.80 mol of $\text{H}_2$ and 0.80 moles of $\text{CO}_2$ are injected into a 5.0 L flask. Calculate the concentration of each species at equilibrium.

a) Calculate initial concentrations of $\text{H}_2$ and $\text{CO}_2$.

$[\text{H}_2] = 0.80 \text{ mol}/5.0 \text{ L} = 0.16 \text{ M} \quad [\text{CO}_2] = 0.80 \text{ mol}/5.0 \text{ L} = 0.16 \text{ M}$.

b) Fill in the ICE chart, using x’s where necessary.

$$
\begin{array}{c|c|c|c}
 & \text{H}_2 (g) & + & \text{CO}_2(g) & \rightleftharpoons & \text{H}_2\text{O}(g) & + & \text{CO}(g) \\
\text{Initial (M):} & 0.16 & & 0.16 & & 0 & & 0 \\
\text{Change (M):} & -x & & -x & & +x & & +x \\
\text{Equil (M):} & 0.16 - x & & 0.16 - x & & x & & x \\
\end{array}
$$

c) Write the $K_C$ expression. Plug in values & solve for x. (Note-no quadratic eq. is needed b/c it’s a perfect square.)

$$
K_c = \frac{[\text{H}_2\text{O}][\text{CO}]}{[\text{H}_2][\text{CO}_2]} \\
4.2 = \frac{x^2}{(0.16 - x)^2}
$$

Taking the square root of both sides, we obtain:

$$
2.0 = \frac{x}{0.16 - x} \\
x = 0.11 \text{ M}
$$

d) What are the equilibrium concentrations of all of the substances?

$$
[\text{H}_2] = [\text{CO}_2] = (0.16 - 0.11) \text{ M} = 0.05 \text{ M} \\
[\text{H}_2\text{O}] = [\text{CO}] = 0.11 \text{ M}
$$
A. For the equilibrium reaction, \( \text{Br}_2(g) + \text{Cl}_2(g) \rightleftharpoons 2\text{BrCl}(g) \) at 400K, the equilibrium constant \( K_c \) is 7.0. If 0.60 mol Br\(_2\) gas, 0.60 mol Cl\(_2\) gas and 2.4 mol BrCl gas are placed into an evacuated 2.0-L reaction flask, what are the equilibrium concentrations of all substances.

a) What are the initial concentrations of each component?

\[
[\text{Br}_2]_0 = [\text{Cl}_2]_0 = \frac{0.60 \text{ mol}}{2.0 \text{ L}} = 0.30 \text{ M}; \quad [\text{BrCl}]_0 = \frac{2.4 \text{ mol}}{2.0 \text{ L}} = 1.2 \text{ M}
\]

b) In this question, initially there are some reactants and some products. Thus, it is not obvious which way the reaction must shift to come to equilibrium. (In contrast, in the previous questions, you were either starting with all reactants or all products.) Thus, solve for “Q” and determine which way the reaction must shift to come to equilibrium.

\[
Q_c = \frac{[\text{BrCl}]^2}{[\text{Br}_2][\text{Cl}_2]} = \frac{(1.2M)^2}{(0.30M)(0.30M)} = 16.
\]

\( Q_c > K_c \) \( \quad \text{Initial ratio of [P]/[R] is larger than at equilibrium.} \)

\( 16 > 7.0 \) \( \quad \text{Thus, rxn shifts to the left to lose products and get more reactants.} \)

c) Fill in ICE chart using x’s whenever necessary. (Remember to shift eq. in the correct direction.)

\[
\begin{array}{c|c|c|c}
 & \text{Br}_2(g) & + & \text{Cl}_2(g) & \rightleftharpoons & 2 \text{BrCl} (g) \\
\hline
\text{Initial (M)} & 0.30 & 0.30 & & & 1.2 \\
\text{Change (M)} & +x & +x & & & -2x \\
\text{Equil (M)} & (0.30+x) & (0.30+x) & & & (1.2 - 2x) \\
\end{array}
\]

d) Now write the \( K_c \) expression, plug in values & solve for “x.” (You should get a perfect square.)

\[
K_c = \frac{[\text{BrCl}]^2}{[\text{Br}_2][\text{Cl}_2]} = \frac{(1.2 - 2x)^2}{(0.30 + x)^2} = 7.0
\]

Taking the square root of both sides we obtain:

\[
\frac{1.2 - 2x}{0.30 + x} = 2.6 \text{ and } x = 0.090 \text{ M}
\]

e) What are the equilibrium concentrations of all substances?

\[
[\text{Br}_2] = [\text{Cl}_2] = (0.30 + 0.090) = 0.39 \text{ M}; \quad [\text{BrCl}] = 1.2 - 2(0.090) = 1.0 \text{ M}
\]