Solving Equilibrium Problems

We are able to group equilibrium problems into two types:

1) We have been given equilibrium concentrations (or partial pressures) and must solve for \( K \) (equilibrium constant).
2) We have been given \( K \) and the initial concentrations and must solve for the equilibrium concentrations.

For the first type of equilibrium problem, we can solve for \( K \) by directly substituting given equilibrium quantities into the reaction quotient:

For example, let’s use the following reaction:

\[
N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)
\]

At equilibrium, the above reaction contains 2.25 M of \( N_2 \), 5.00 M of \( H_2 \), and 3.5 M of \( NH_3 \). Calculate the equilibrium constant.

Because we have been given the equilibrium concentrations of each reactant AND product, we can simply substitute these quantities into the reaction quotient:

\[
Q_c = \frac{[NH_3]^2}{[N_2][H_2]^3}
\]

therefore,

\[
K_c = \frac{(3.50)^2}{(2.25)(5.00)^3} = 0.218
\]

We can also solve for \( K \) if we have only been given some quantities, but not all. For instance, if you were given the initial concentrations and equilibrium concentrations you can set up a reaction table or ‘ICE’ table to help you calculate the equilibrium constant, \( K \).

For example, let’s use the following reaction and quantities to solve for \( K \).

The decomposition of nitrogen oxide is shown by the reaction below:

\[
2NO \rightleftharpoons N_2 + O_2
\]

This reaction was studied at 298 K with initial amount of 0.215 M of NO gas. At equilibrium, the concentration of NO was 0.083M. Calculate \( K \) for this reaction.

It is important to recognize that we have initial and equilibrium amounts for NO, but we don’t know the equilibrium amounts of our products, \( N_2 \) and \( O_2 \). When we don’t know some of our equilibrium amounts, we must set up a reaction or ICE table.

<table>
<thead>
<tr>
<th>Concentration (M)</th>
<th>2NO(g)</th>
<th></th>
<th>N_2(g)</th>
<th>+</th>
<th>O_2(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.215M</td>
<td></td>
<td>0</td>
<td></td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-2x</td>
<td>+x</td>
<td></td>
<td></td>
<td>+x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.215-2x</td>
<td>x</td>
<td></td>
<td></td>
<td>x</td>
</tr>
</tbody>
</table>

In order to solve for \( K_c \), we need equilibrium concentrations for all reactants and products. Based on the balanced equation, we know that when 2x moles of NO reacts, x moles of \( N_2 \) and \( O_2 \) will form. We also were given the equilibrium concentration for NO (0.083M), so we can solve for \( x \):

\[
0.215 - 2x = 0.083M
\]

Solve for \( x \):

\[
x = (0.083M - 0.215M) / -2
\]
Now that we have determined \( x \), we can substitute the concentration into \( Q_c \):

\[
Q_c = \frac{[N_2][O_2]}{[NO]^2} \quad \text{therefore,} \quad K_c = \frac{(0.066)(0.066)}{(0.083)^2} = 0.632
\]

The second type of equilibrium problem you may encounter will give you both initial concentrations and \( K \) and then ask you to solve for the equilibrium concentrations.

Let's examine the reaction involving the decomposition of HI:

\[
2\text{HI}(g) \rightleftharpoons \text{I}_2(g) + \text{H}_2(g) \quad K_c = 0.67
\]

If 3.0 M of HI is placed in a flask, what is the equilibrium concentration of each product and reactant?

Because we were given initial amounts, we must complete an ICE table:

<table>
<thead>
<tr>
<th>Concentration (M)</th>
<th>2HI(g)</th>
<th>( \text{I}_2 ) (g)</th>
<th>+</th>
<th>( \text{H}_2 ) (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>3.0</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-2x</td>
<td>x</td>
<td>+x</td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>3.0-2x</td>
<td>x</td>
<td>x</td>
<td></td>
</tr>
</tbody>
</table>

We were given the equilibrium constant for this reaction (\( K_c = 0.67 \)), so we can set up our reaction quotient:

\[
Q_c = \frac{[\text{I}_2][\text{H}_2]}{[\text{HI}]^2} \quad \text{therefore} \quad K_c = 0.67 = \frac{(x)^2}{(3.0-2x)^2}
\]

We can take the square root of each side of the equation:

\[
\sqrt{0.67} = \frac{x}{3.0 - 2x}
\]

Next, we can multiply each side by (3.0-2x):

\[
2.46 - 1.64x = x
\]

So, \( 2.46 = 2.64x \) therefore \( x = 0.93 \)

Now that we have solved for \( x \), we can calculate the equilibrium concentrations.

\[
[\text{I}_2] = [\text{H}_2] = x = 0.93 \text{ M}
\]

\[
[\text{HI}] = 3.0 - 2(0.93) = 1.14 \text{ M}
\]