TLC
Teaching and Learning Center
UNIVERSITY OF WASHINGTON TACOMA

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 to solve for K by directly substituting given equilibrium quantities into the reaction quotient:

For example, let's use the following reaction:

$N_2(g)\ +\ 3H_2(g) \longleftrightarrow 2NH_3(g)$

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:

$$Q_c = \frac{[NH_3]^2}{[N_2][H_2]^3} \quad \text{therefore,} \quad K_c = \frac{(3.50)^2}{(2.25)(5.00)^3} = \boxed{0.218}$$

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:

$2NO\ (g) \longleftrightarrow N_2\ (g) + O_2\ (g)$

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_c 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.

Concentration (M)	**2NO(g)**	$\longleftrightarrow$	**N_2(g)**	**+**	**O_2(g)**
Initial	0.215M		0		0
Change	-2x		+x		+x
Equilibrium	0.215-2x		x		x

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

x= 0.066M

Now that we have determined x, we can substitute the concentration into Q_c:

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

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:

$2HI(g) \longleftrightarrow I_2(g) + H_2(g)$ $\quad K_c = 0.67$

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:

Concentration (M)	**$2HI(g)$**	$\longleftrightarrow$	**$I_2(g)$**	**+**	**$H_2(g)$**
Initial	3.0		0		0
Change	-2x		x		+x
Equilibrium	3.0-2x		x		x

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

$$Q_c = \frac{[I_2][H_2]}{[HI]^2} \quad \text{therefore} \quad K_c = 0.67 = \frac{(x)^2}{(3.0-2x)^2}$$

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.

$[I_2]$ =$[H_2]$=x=0.93 M

[HI]= 3.0 – 2(0.93) = 1.14M