

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:



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.

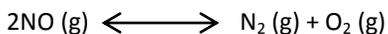
Because we have been given the equilibrium concentrations of each reactant AND product, we can simply substitute these quantities into the reaction quotient:

$$Q_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} \quad \text{therefore,} \quad K_c = \frac{(3.50)^2}{(2.25)(5.00)^3} = \boxed{0.218}$$

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:



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)	\rightleftharpoons	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 :

$$0.215 - 2x = 0.083\text{M}$$

Solve for x :

$$x = (0.083\text{M} - 0.215\text{M}) / -2$$

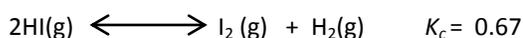
$$x = 0.066\text{M}$$

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:



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)	\rightleftharpoons	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):

$$2.46 - 1.64x = x$$

$$\text{So, } 2.46 = 2.64x \quad \text{therefore } x = 0.93$$

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

$$[I_2] = [H_2] = x = 0.93 \text{ M}$$

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