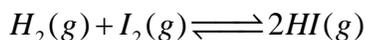


## What Does the Equilibrium Constant Tell Us?

First of all, you remember that the equilibrium constant is a constant for a particular reaction at a particular temperature. If the temperature changes, the value of the equilibrium constant also changes. In addition, the equilibrium constant expression is nothing but the quotient of concentrations of products and concentrations of reactants.

If we start the reaction with certain amount of reactants, how do we know when the equilibrium is established? If the system is not in equilibrium, what does it mean? What needs to be done to make it to attain equilibrium? We would like to know the answers to these questions. It is best illustrated using the following example.

Consider the following equilibrium reaction



with equilibrium constant  $K_c$  equal to 54.3 at  $430^\circ\text{C}$ . In a particular experiment, we placed 0.5 mole of  $H_2$ , 0.2 mole of  $I_2$ , and 0.8 mole of  $HI$  in a 1.0-L container at  $430^\circ\text{C}$ . Is the system in equilibrium? If not, what needs to be done to make it to establish an equilibrium?

We substitute the given concentrations into the equilibrium constant expression, but do not call it as an equilibrium constant because these concentrations are not the equilibrium concentrations rather they are initial concentrations. We call this ratio as **reaction quotient ( $Q_c$ )**.  $Q_c$  is calculated exactly the same way as the equilibrium constant except using the given initial concentrations. Further, to distinguish the equilibrium concentration (which is indicated by the pair of square brackets ( $[ ]$ )) from the initial concentration, the same pair of square brackets are using with 0 as subscript to mean initial concentration. So, we write the equation for  $Q_c$  as

$$Q_c = \frac{[HI]_0^2}{[H_2]_0[I_2]_0} = \frac{(0.8)^2}{(0.5)(0.2)} = 6.4$$

Comparing  $Q_c$  with the given  $K_c$ , is clear that  $Q_c$  is much less than  $K_c$ . Hence, the system is not in equilibrium. In order to establish equilibrium,  $Q_c$  must be equal to  $K_c$ . To do that the ratio of concentrations must be increased by forming more  $HI$ . **Thus the net reaction proceeds from left to right to reach equilibrium.**

In general, there are three possibilities:

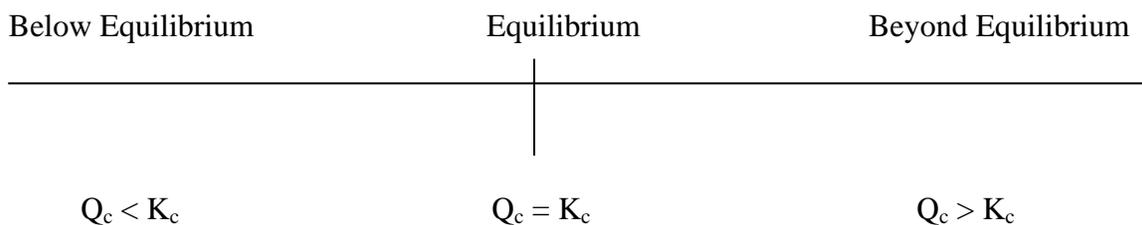
If  $Q_c = K_c$ , The system is in equilibrium. It means that the initial concentrations are equilibrium concentrations.

If  $Q_c < K_c$ , The system is below the equilibrium. The ratio of concentrations is small.

To reach equilibrium, reactants must be converted to products. It means that **the system must proceed from left to right to reach equilibrium.**

If  $Q_c > K_c$ , The system has gone beyond the equilibrium. The ratio of concentrations is high. To reach equilibrium, products must be converted back into reactants. It means that **the system must proceed from right to left to reach equilibrium.**

The following diagram gives better understanding.



Thus, knowing the initial concentrations, one can predict whether the system is in equilibrium. If not, in what direction ( left to right or right to left) it has proceed to reach equilibrium. This gives us important information about the status of the system.

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