**Using the Equilibrium Constant**

 **Knowing the value of the equilibrium constant for a chemical reaction can allow you to tell the extent of the reaction, predict the direction and calculate equilibrium concentrations from initial concentrations.**

 **The numerical value of the equilibrium constant for a reaction can be used to measure how far the reaction goes before equilibrium is reached.**

**Example:**

**2H2(g) + O2(g) 2H2O (g) Kc = 2.4 x 1047@ 500K**

**Kc = [H2O]2 / [H2]2[O2] = 2.4 x 1047**

 **Because you divide the products by the reactants, a very large value of kc means that the equilibrium ration of products to reactants is very large.**

 **Therefore, when Kc is very large the reaction proceeds nearly to completion.**

**For example:**

 **If you have a 2 to 1 (from coefficients) ration amounts of H2 gas to O2 and [H2O] = .1 M at equilibrium, the concentrations of H2 and O2 are quite small relative to [H2O] (.1M).**

**[H2] = 4.4 x 10-17 M [O2]= = 2.2 x 10-17 M**

 **By contrast, if a reaction has a very small value of Kc, the equilibrium ratio of products to reactants is very small and the reaction proceeds hardly at all before equilibrium is reached.**

**Example: revers H2 + O2.**

**2H2O (g) 2H2(g) + O2 (g)**

**Kc = ?**

 **The Kc value for a reaction that is reversed can be found by getting the reciprocal of the original Kc from the forward reaction.**

**Kc = 1 / 2.4 x 1047 = 4.2 x 10-48.**

 **The reverse reaction gives the same equilibrium mixture as obtained by the forward reaction ([H2] = 4.4 x10-17, [O2] = 2.2 x 10-17).**

 **However, the reverse reaction does not occur to any appreciable extent because of how small the Kc value is: 4.2 x 10-48.**