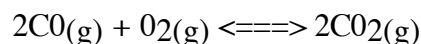


EQUILIBRIUM PROBLEMS

If you can determine the equilibrium concentrations of the chemicals involved in a reaction, then you can write an equilibrium expression and determine the value for the K_{eq} for that reaction.

For example let's use the following reaction:



Suppose that the equilibrium concentrations of the chemicals was found to be...

$$[CO(g)] = 0.2 \text{ moles/liter}$$

$$[O_2(g)] = 0.2 \text{ moles/liter}$$

$$[CO_2(g)] = .00075 \text{ moles/liter}$$

The brackets [] around the formulas of the chemicals actually mean their concentration in moles/liter.

We can write the equilibrium expression for this reaction --- substitute the concentrations and solve for the equilibrium constant.

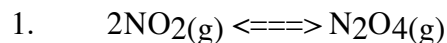
$$K_{eq} = \frac{[CO_2]^2}{[CO]^2[O_2]}$$

$$K_{eq} = \frac{[0.00075]^2}{[0.2]^2[0.2]}$$

$$K_{eq} = 7.03 \times 10^{-5}$$

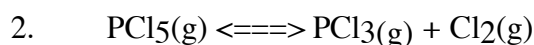
PRACTICE PROBLEMS

For each of the following reactions calculate the equilibrium constant. The equilibrium concentrations are given.



$$[NO_2(g)] = 0.0045 \text{ moles/liter}$$

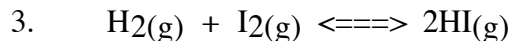
$$[N_2O_4(g)] = .000005 \text{ moles/liter}$$



$$[PCl_5(g)] = 0.5 \text{ moles/liter}$$

$$[PCl_3(g)] = 0.0075 \text{ moles/liter}$$

$$[Cl_2(g)] = 0.0075 \text{ moles/liter}$$



$$[H_2(g)] = 4.5647 \times 10^{-3} \text{ moles/liter}$$

$$[I_2(g)] = 0.7378 \times 10^{-3} \text{ moles/liter}$$

$$[HI(g)] = 13.544 \times 10^{-3} \text{ moles/liter}$$