

## Equilibrium Constant Problems

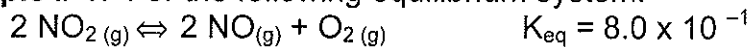
Calculating Equilibrium Concentrations:

Step 1: Write out the equation of the reaction

Step 2: Write out the  $K_{eq}$  expression

Step 3: Use ICE diagrams if needed

**Example # 1:** For the following equilibrium system:

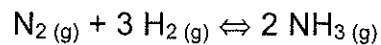


When the system is at equilibrium it is known that  $[\text{NO}_2(g)] = 0.60 \text{ mol/L}$  and  $[\text{NO}(g)] = 0.40$ . Calculate the equilibrium concentration of  $\text{O}_2(g)$ .

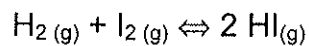
**Example #2:** At the same temperature, the same system as above is at equilibrium with

$[\text{NO}(g)] = 0.25 \text{ mol/L}$  and  $[\text{O}_2(g)] = 0.20 \text{ mol/L}$

**Example #3:** Initially 0.32 moles of  $\text{N}_2(\text{g})$  gas and 0.66 moles  $\text{H}_2(\text{g})$  were placed in a 1.0 L container. At equilibrium the concentration of  $\text{H}_2(\text{g})$  was 0.30 mol/L. Calculate the equilibrium concentrations of all the entities and then calculate the  $K_{\text{eq}}$ .

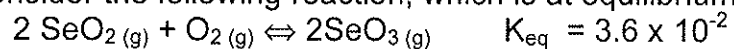


**Example #4:** Using the following information calculate the value of  $K_{\text{eq}}$  for  $\text{HI}(\text{g})$ . Initially the concentrations of  $\text{H}_2(\text{g})$  and  $\text{I}_2(\text{g})$  were 0.200 mol/L. Calculate the concentration of  $\text{HI}(\text{g})$  at equilibrium.



Questions:

1. Consider the following reaction, which is at equilibrium:

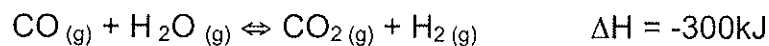


At equilibrium,  $[\text{SeO}_2] = 4.5 \times 10^{-1} \text{ mol/L}$  and  $[\text{O}_2] = 7.0 \times 10^{-1} \text{ mol/L}$ .  
Calculate the equilibrium  $[\text{SeO}_3]$ .

2. For the reaction:  $2 \text{SO}_3(g) \rightleftharpoons 2 \text{SO}_2(g) + \text{O}_2(g)$

A chemist placed 0.100 mol of  $\text{SO}_3(g)$  and 0.200 mol of  $\text{SO}_2(g)$  in a 5.00 L container. At equilibrium, the  $[\text{SO}_3]$  was found to be 0.0160 mol/L. Calculate the value of  $K_{\text{eq}}$ .

3. Consider the equation of the reversible reaction:



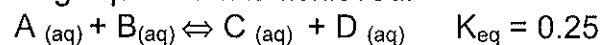
Initially 0.500 mol of  $\text{CO}_{(g)}$  and 0.500  $\text{H}_2\text{O}_{(g)}$  are placed in a 2.00 L container at  $400^\circ\text{C}$ . Calculate the equilibrium concentrations of all species at equilibrium. Calculate  $K_{\text{eq}}$  if the final  $[\text{H}_2\text{O}] = 0.100\text{M}$ .

If the reaction were carried out at  $700^\circ\text{C}$  how would the value of  $K_{\text{eq}}$  compare with that of  $400^\circ\text{C}$ ?

4. For the reaction:  $2 \text{H}_2\text{S}_{(g)} \rightleftharpoons 2 \text{H}_{2(g)} + \text{S}_{2(g)}$

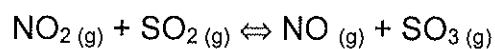
$4.8 \times 10^{-2}$  mol of  $\text{H}_2\text{S}_{(g)}$  was placed in a 2.00 L container. At equilibrium, the  $[\text{S}_2]$  was found to be  $1.1 \times 10^{-2}$  mol/L. Calculate the value of  $K_{\text{eq}}$ .

5. 1.00 mol of each of A<sub>(aq)</sub>, B<sub>(aq)</sub>, C<sub>(aq)</sub>, D<sub>(aq)</sub> are dissolved in 500 mL of water. The following equilibrium is achieved:



Calculate the equilibrium concentrations of all four species. (Which direction will the system move to achieve equilibrium?)

6. At 1000°C,  $K_{eq} = 3.60$  for the reaction:



Initially, 0.200 mol of NO<sub>2(g)</sub> and 0.200 mol of SO<sub>2(g)</sub> are placed in a 5.00 L container. Calculate the equilibrium concentrations of all four species.