C116 Equilibrium Work Sheets 013013

1. Given that the following equilibrium is present in a 0.250L stoppered glass flask: $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) = 2\text{HI}(\text{g})$
   $\text{K}_c = 56.00$. The equilibrium concentrations are $[\text{H}_2(\text{g})] = [\text{I}_2(\text{g})] = 0.42\text{M}$ and $[2\text{HI}(\text{g})] = 3.16\text{M}$.
   Suppose we disturb the equilibrium by adding enough HI to bring its concentration to 5.00 M.

   A. Which direction will the reaction shift to re-establish equilibrium?

   $Q_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{(5.00)^2}{(0.42)^2} = 142 > \text{K}_c$ $\Rightarrow$ reaction shifts left towards reactants

   B. When equilibrium is re-established, what are the equilibrium concentrations of all three components?

   The initial conditions for the ICE table are those instantly after the equilibrium is perturbed.

   $\begin{align*}
   \text{H}_2 + \text{I}_2 & \rightleftharpoons 2\text{HI} \\
   I & \quad 0.42\text{M} \\
   C & \quad 0.42\text{M} \\
   E & \quad 0.42\text{M} \\
   \text{K}_c & = 56.00 = \frac{(5.00-3x)^2}{(0.42+x)^2} \\
   -1.48x & = 5.00-3x \\
   0.196 & = x
   \end{align*}$

   $\begin{align*}
   [\text{H}_2]_e & = [\text{I}_2]_e = 0.196\text{M} \\
   [\text{HI}]_e & = 5.00-3(0.196) = 4.11\text{M}
   \end{align*}$

2. Consider the equilibrium reaction: $\text{I}_2(\text{g}) = 2\text{I}(\text{g})$. $\text{K}_c = 5.6 \times 10^{-12}$
   If the initial concentration of $\text{I}_2(\text{g})$ is 0.45M, what are the equilibrium concentrations of the reactant and the product?

   $\begin{align*}
   \text{I}_2 & = 2\text{I} \\
   I & \quad 0.45\text{M} \\
   C & \quad x \\
   E & \quad 2x \\
   \text{K}_c & = 5.6 \times 10^{-12} = \frac{(2x)^2}{0.45-x} \Rightarrow x = 2.51 \times 10^{-6} \text{M} \\
   2x & = 5.02 \times 10^{-6} \text{M} = [\text{I}] \\
   0.45-x & = 0.444 \text{M} = [\text{I}_2]
   \end{align*}$

3. $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) = 2\text{NH}_3(\text{g})$ summarizes nitrogen fixation, and, $\text{K}_c = 3.5 \times 10^8$.
   Calculate $\text{K}_c$ for the reaction: $\text{NH}_3(\text{g}) = \frac{1}{2}\text{N}_2(\text{g}) + 3/2\text{H}_2(\text{g})$.

   $\begin{align*}
   \text{Rxn 3} & = \frac{1}{2}(\text{Rxn 1}) \Rightarrow \text{K}_c\text{Rxn 3} = \frac{1}{(\text{K}_c\text{Rxn 1})^{1/2}} = (\frac{1}{3.5 \times 10^8})^{1/2} \\
   \text{K}_c\text{Rxn 3} & = 5.35 \times 10^{-5}
   \end{align*}$
4. A. Does the equilibrium concentration of NOCl increase or decrease as the temperature of the reaction mixture is increased?

\[
\text{Heat} + 2\text{NOCl(g)} \rightleftharpoons 2\text{NO(g)} + \text{Cl}_2(g) \quad \Delta H = +77.2 \text{ kJ}
\]

Endothermic

Heating shifts reactants to products
\[\text{[NOCl] decreases}\]

B. Does the equilibrium concentration of SO\textsubscript{3}(g) increase or decrease when the temperature of the reaction mixture is increased?

\[
2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g) \quad \Delta H = -198.7 \text{ kJ}
\]

This was meant to be exothermic

For exothermic reactions, \[\text{[SO}_3\text{]} \text{ increases}\]

5. A. For each reaction in problem 4: Which way does the reaction shift if the volume of the reaction container is increased? Volume increase favors side with greater moles.

4A: Shifts right
4B: Shifts left

B. For each reaction in problem 4: Which way does the reaction shift if the volume of the reaction container is decreased? Volume decrease favors side with fewer moles.

4A: Shifts left
4B: Shifts right