1. **True/False statements.** [5 pts] **Homework Problem**

   a) At equilibrium, the concentrations of the reactants and products are no longer changing with time. **T**

   b) When the value of $K_c$ for a reaction is large, it indicates that the equilibrium concentrations of the reactants are large compared to the equilibrium concentrations of products. **F**

   c) At equilibrium, the rates of the forward and reverse reactions are equal. **T**

   d) At equilibrium, the reaction has stopped. No reactants are converted to products and no products are converted to reactants. **F**

   e) At equilibrium, the concentrations of products and reactants are equal. **F**

2. Write the equilibrium constant expressions in terms of $K_c$ for the following two reactions. [4 pts]

   a) $2 \text{O}_3(g) \rightleftharpoons 3 \text{O}_2(g)$

      $$K_c = \frac{[\text{O}_2]^3}{[\text{O}_3]^2}$$

   b) $\text{SnO}_2(s) + 2 \text{CO}(g) \rightleftharpoons \text{Sn}(s) + 2 \text{CO}_2(g)$

      $$K_c = \frac{[\text{CO}_2]^2}{[\text{CO}]^2}$$

3. The value of the equilibrium constant, $K_c$, for a reaction is 0.36. If the equation for the reaction is multiplied through by 3, that is, all the coefficients are multiplied by 3, and the direction of the reaction is reversed, what is the value of the equilibrium constant, $K_c$, for the reaction as it is now written? [3 pts] **Homework Problem**

   a) 0.36     b) 1.1     c) 0.93     d) 0.047     e) 21

   $$K_c' = \frac{1}{(K_c)^3} = \frac{1}{(0.36)^3} = 21$$

4. Consider the following equilibrium at 1000K:

   $$2 \text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2 \text{SO}_3(g)$$

   A study of this system reveals that there are $3.6 \times 10^{-3}$ moles of SO$_2$ gas, and $4.1 \times 10^{-3}$ moles of O$_2$ present in a 12.0 L flask at equilibrium. The equilibrium constant for this reaction is $3.4 \times 10^{-3}$. Calculate the number of moles of SO$_3$(g) in the flask at equilibrium. [4 pts] **Homework Problem**

   $$[\text{SO}_2] = 3.6 \times 10^{-3} \text{ mol/12.0 L} = 3.00 \times 10^{-4} M$$

   $$[\text{O}_2] = 4.1 \times 10^{-3} \text{ mol/12.0 L} = 3.42 \times 10^{-4} M$$
\[ K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]} \]

\[
3.4 \times 10^{-3} = \frac{[\text{SO}_3]^2}{(3.00 \times 10^{-4})^2(3.42 \times 10^{-4})}
\]

\[
[\text{SO}_3]^2 = 1.05 \times 10^{-13}
\]

\[
[\text{SO}_3] = 3.24 \times 10^{-7} \text{ M}
\]

\[
\text{mol SO}_3 = \frac{3.24 \times 10^{-7} \text{ mol}}{1 \text{ L}} \times 12.0 \text{ L} = 3.9 \times 10^{-6} \text{ mol}
\]

5. Which statement about the equilibrium reaction below is false? [3 pts]

\[
\text{CO}_2 (g) + \text{H}_2 (g) \rightleftharpoons \text{CO} (g) + \text{H}_2\text{O} (g) \quad K = 1 \times 10^{-2} \text{ at } 327^\circ \text{C}
\]

[a] \([\text{CO}_2][\text{H}_2]\) equals \([\text{CO}][\text{H}_2\text{O}]\) at equilibrium at 327°C.
[b] The equilibrium constant for the reverse reaction is \(1 \times 10^2\) at 327°C.
[c] \(\text{H}_2\text{O}\) appears in the equilibrium constant expression for the reaction.
[d] At equilibrium, the rate of formation of \(\text{CO}_2\) and \(\text{H}_2\) is equal to the rate of formation of \(\text{CO}\) and \(\text{H}_2\text{O}\).
[e] For any equilibrium position, \(K\) will be \(1 \times 10^{-2}\) as long as the temperature remains at 327°C.

6. At a particular temperature, \(K_p = 1.0 \times 10^{-3}\) for the reaction

\[
3 \text{ H}_2 (g) + \text{N}_2 (g) \rightleftharpoons 2 \text{NH}_3 (g)
\]

The initial pressure of \(\text{NH}_3\) is 0.105 atm. Which of the following represents the correct setup to solve for the equilibrium pressures? [3 pts]

[a] \(1.0 \times 10^{-3} = \frac{(0.105 - 2x)^2}{27x^4}\)
[b] \(1.0 \times 10^{-3} = \frac{(0.105 - 2x)^2}{9x^4}\)
[c] \(1.0 \times 10^{-3} = \frac{(0.105 - 2x)^2}{3x^4}\)
[d] \(1.0 \times 10^{-3} = \frac{(0.105 - x)^2}{x^4}\)
[e] \(1.0 \times 10^{-3} = \frac{0.105 - 2x}{3x^2}\)

Initial (atm): 0 0 0.105
Change (atm): +3x +x −2x
Equilibrium (atm): 3x x 0.105 −2x

\[
K_p = \frac{(P_{\text{NH}_3})^2}{(P_{\text{H}_2})^3(P_{\text{N}_2})}
\]

\[
1.0 \times 10^{-3} = \frac{(0.105 - 2x)^2}{(3x)^3(x)} = \frac{(0.105 - 2x)^2}{27x^4}\]
7. Nitrogen and oxygen gas react to form nitrogen monoxide, NO.

$$\text{N}_2 \ (g) + \text{O}_2 \ (g) \rightleftharpoons 2 \text{NO} \ (g)$$

At room temperature, the equilibrium constant, $K_c$, equals $7.1 \times 10^{-31}$. If the reaction is started with $[\text{N}_2] = 0.500 \ M$ and $[\text{O}_2] = 0.800 \ M$, calculate the concentration of NO at equilibrium. [Hint: Note the magnitude of $K_c$] [6 pts]

\[
\begin{array}{c|ccc}
\text{Initial (M):} & \text{N}_2 & \text{O}_2 & 2 \text{NO} \\
0.500 & 0.800 & 0 \\
\text{Change (M):} & -x & -x & +2x \\
\text{Equilibrium (M):} & 0.500 - x & 0.800 - x & 2x \\
\end{array}
\]

\[
K_c = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]}
\]

\[
7.1 \times 10^{-31} = \frac{(2x)^2}{(0.500 - x)(0.800 - x)} \approx \frac{4x^2}{(0.500)(0.800)}
\]

\[
x = 2.665 \times 10^{-16} \ M
\]

\[
[\text{NO}] = 2x = 5.33 \times 10^{-16} \ M
\]