CHEM 1412
Chapter 17 Homework

1. Explain the significance of
   (a) a very large value of $K$
   A large $K_c$ means the products are favored in the reaction.

   (b) a very small value of $K$
   A small value of $K_c$ means the reactants are favored in the reaction.

   (c) a value of $K$ of about 1.
   This means the reactants are products are more or less equally favored.

2. What does the value of the equilibrium constant tell us about the time required for the reaction to reach equilibrium?
   Nothing. Reaction rates are not related to the position of the equilibrium.

3. Write the expression for $K_c$ for each of the following reactions.
   (a) $2 \text{NO}_2(g) \rightleftharpoons 2 \text{NO}(g) + \text{O}_2(g)$
   $$K_c = \frac{[\text{NO}]^2[\text{O}_2]}{[\text{NO}_2]^2}$$

   (b) $2 \text{CHCl}_3(g) + 3 \text{H}_2(g) \rightleftharpoons 2 \text{CH}_4(g) + 3 \text{Cl}_2(g)$
   $$K_c = \frac{[\text{CH}_4]^2[\text{Cl}_2]^3}{[\text{CHCl}_3]^2[\text{H}_2]^3}$$

   (c) $\text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2 \text{HI}(g)$
   $$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

4. Which is NOT necessarily true about a reaction at chemical equilibrium?
   (a) The equilibrium is dynamic rather than static.
   (b) There is no net change for either reactants or products.
   (c) The number of moles of reactants is equal to the number of moles of products.
   (d) Individual reactant and product molecules continue to undergo reaction.

5. Equilibrium in a chemical reaction can be attained by
   (a) starting with reactants.
   (b) starting with products.
   (c) starting with a combination of reactants and products.
   (d) all of the above
6. Consider the following compounds in the states indicated, as possible reactants or products in a chemical equilibrium. Which of these compounds would be omitted from the equilibrium expression?

\[
\text{H}_2\text{O}_\text{(s)} \quad \text{H}_2\text{O}_\text{(l)} \quad \text{H}_2\text{O}_\text{(g)} \quad \text{HCl}_\text{(g)} \quad \text{HCl}_\text{(aq)} \quad \text{Fe}_2\text{O}_3\text{(s)} \quad \text{NaHCO}_3\text{(aq)} \quad \text{NH}_3\text{(g)} \quad \text{NH}_3\text{(aq)}
\]

7. The reaction between \( \text{N}_2 \) and \( \text{O}_2 \) to form \( \text{NO} \) is represented by the chemical equation

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

Equilibrium concentrations of the gases at 1500 K are \( 1.7 \times 10^{-5} \) M \( \text{O}_2 \), \( 6.4 \times 10^{-3} \) M \( \text{N}_2 \), and \( 1.1 \times 10^{-3} \) M \( \text{NO} \). Calculate \( K_c \) at 1500 K.

\[
K_c = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]} = \frac{(1.1 \times 10^{-5})^2}{(6.4 \times 10^{-3})(1.7 \times 10^{-5})} = 1.1 \times 10^3
\]

8. At elevated temperatures, \( \text{BrF}_5 \) establishes the following equilibrium

\[
2 \text{BrF}_3\text{(g)} \rightleftharpoons \text{Br}_2\text{(g)} + 5 \text{F}_2\text{(g)}
\]

The equilibrium concentration of the gases are \( 0.0064 \) M \( \text{BrF}_5 \), \( 0.0018 \) M \( \text{Br}_2 \), and \( 0.0090 \) M \( \text{F}_2 \). What is the value of \( K_c \)?

\[
K_c = \frac{[\text{Br}_2][\text{F}_2]^5}{[\text{BrF}]^2} = \frac{(0.0018)(0.0090)^5}{(0.0064)^2} = 2.6 \times 10^{-9}
\]

9. The value of \( K_c \) for a chemical reaction is 1.50. What does this mean?
   (a) Products are favored, but just barely.
   (b) Most of the initial concentrations of reactants are present under equilibrium conditions.
   (c) It takes a relatively long time for this system to come to equilibrium.
   (d) This value has no meaning until we have the balanced chemical equation.

10. Which statement about \( K_c \) is NOT true?
   (a) \( K_c \) is independent of temperature.
   (b) \( K_c \) is independent of starting concentrations.
   (c) \( K_c \) measures the extent of the reaction.
   (d) \( 1/K_c \) measures the extent of the reverse reaction.

11. For the following equation, \( K_c = 7.9 \times 10^{11} \) at 500 K.

\[
\text{H}_2\text{(g)} + \text{Br}_2\text{(g)} \rightleftharpoons 2 \text{HBr}\text{(g)}
\]

Find \( K_c \) for the following equilibria.

For the given reaction, \( K_c = \frac{[\text{HBr}]^2}{[\text{H}_2][\text{Br}_2]} = 7.9 \times 10^{11} \)
(a) \[ \frac{1}{2} \text{H}_2(\text{g}) \, + \, \frac{1}{2} \text{Br}_2(\text{g}) \, \rightleftharpoons \, \text{HBr}(\text{g}) \]

\[
K_c = \frac{[\text{HBr}]}{[\text{H}_2]^{1/2}[\text{Br}_2]^{1/2}} \quad \text{this expression is the square root of the original expression}
\]

\[
K_c = \sqrt{7.9 \times 10^{11}} = 8.9 \times 10^5
\]

(b) \[2 \text{HBr}(\text{g}) \, \rightleftharpoons \, \text{H}_2(\text{g}) \, + \, \text{Br}_2(\text{g})\]

\[
K_c = \frac{[\text{H}_2][\text{Br}_2]}{[\text{HBr}]^2} \quad \text{this expression is the inverse of the original expression}
\]

\[
K_c = \frac{1}{7.9 \times 10^{11}} = 1.3 \times 10^{-12}
\]

(c) \[2 \text{H}_2(\text{g}) \, + \, 2 \text{Br}_2(\text{g}) \, \rightleftharpoons \, 4 \text{HBr}(\text{g})\]

\[
K_c = \frac{[\text{HBr}]^4}{[\text{H}_2]^2[\text{Br}_2]^2} \quad \text{this expression is the square of the original expression}
\]

\[
K_c = (7.9 \times 10^{11})^2 = 6.2 \times 10^{23}
\]

12. For the reaction

\[ \text{H}_2\text{CO} \, \rightleftharpoons \, \text{H}_2 \, + \, \text{CO} \]

\[K_c = 0.50 \text{ at a given temperature. After some period of time, a reaction mixture was tested and found to contain 0.50 M \text{H}_2\text{CO}, 1.50 M \text{H}_2 \text{ and 0.25 M CO. Is the reaction mixture at equilibrium? If not, in which direction must the reaction proceed to reach equilibrium?}}\]

\[
Q_c = \frac{[\text{H}_2][\text{CO}]}{[\text{H}_2\text{CO}]} = \frac{(1.50)(0.25)}{(0.50)} = 0.75
\]

\[Q_c > K_c, Q_c \text{ must decrease in value for equilibrium to be attained. Therefore, the reaction must proceed to the left (decrease products) to reach equilibrium.}\]

13. If \(Q\), the reaction quotient, is less than \(K\), the equilibrium constant, then

(a) the reaction will proceed in the forward direction.

(b) the reaction will proceed in the reverse direction.

(c) the number is too large.

(d) the denominator is too small.
14. A 1.00-L sealed tube initially contains $9.84 \times 10^{-4}$ mol $H_2$ and $1.38 \times 10^{-3}$ mol $I_2$. It is kept at $350 \, ^\circ C$ until the reaction

$$H_2(g) + I_2(g) \rightleftharpoons 2 \, HI(g)$$

comes to equilibrium. At equilibrium $4.73 \times 10^{-4}$ mol $I_2$ is present. Calculate (a) the number of moles of $H_2$ and HI present at equilibrium, and (b) the equilibrium constant, $K_c$.

$$
\begin{array}{ccc}
\text{initial} & 9.84 \times 10^{-4} & 1.38 \times 10^{-3} & 0 \\
\Delta & -9.1 \times 10^{-4} & -9.1 \times 10^{-4} & +2(9.1 \times 10^{-4}) \\
equil. & 7 \times 10^{-5} & 4.73 \times 10^{-4} & 1.8 \times 10^{-3}
\end{array}
$$

$$K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{(1.8 \times 10^{-3})^2}{(7 \times 10^{-5})(4.73 \times 10^{-4})} = 100 \text{ or } 1 \times 10^2$$

15. NO and $O_2$ are mixed in a container of fixed volume and kept at 1000 K. Their initial concentrations are 0.0200 M and 0.0300 M, respectively. When the reaction

$$2 \, NO(g) + O_2(g) \rightleftharpoons 2 \, NO_2(g)$$

has come to equilibrium, the concentration of $NO_2$ is $2.2 \times 10^{-3}$ M. Calculate the equilibrium constant, $K_c$.

$$
\begin{array}{ccc}
\text{initial} & 0.0200 & 0.0300 & 0 \\
\Delta & -2.2 \times 10^{-3} & -1.1 \times 10^{-3} & +2.2 \times 10^{-3} \\
equil. & 0.0178 & 0.0289 & 2.2 \times 10^{-3}
\end{array}
$$

$$K_c = \frac{[NO_2]^2}{[NO]^2[O_2]} = \frac{(2.2 \times 10^{-3})^2}{(0.0178)^2(0.0289)}$$

$$K_c = 0.53$$
16. For the reaction described by the equation

\[ \text{N}_2(\text{g}) + \text{C}_2\text{H}_2(\text{g}) \rightleftharpoons 2 \text{HCN(}\text{g}) \]

\( K_c = 2.3 \times 10^{-4} \) at 300 °C. What is the equilibrium concentration of HCN if the initial concentrations of \( \text{N}_2 \) and \( \text{C}_2\text{H}_2 \) are 2.2 M and 1.0 M, respectively?

\[ \begin{align*}
\text{initial} & \quad 2.2 & \quad 1.0 & \quad 0 \\
\Delta & \quad -x & \quad -x & \quad +2x \\
\text{equil} & \quad 2.2-x & \quad 1.0-x & \quad 2x
\end{align*} \]

\[ K_c = \frac{[\text{HCN}]^2}{[\text{N}_2][\text{C}_2\text{H}_2]} \]

\[ 2.3 \times 10^{-4} = \frac{(2x)^2}{(2.2 - x)(1.0 - x)} \]

\[ 2.3 \times 10^{-4} = \frac{4x^2}{(2.2 - 3.2x + x^2)} \]

\[ 2.3 \times 10^{-4}(2.2 - 3.2x + x^2) = 4x^2 \]

\[ 5.1 \times 10^{-4} - 7.4 \times 10^{-4}x + 2.3 \times 10^{-4}x^2 = 4x^2 \]

\[ 4x^2 + 7.4 \times 10^{-4}x - 5.1 \times 10^{-4} = 0 \]

\[ x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} \]

\[ x = \frac{-7.4 \times 10^{-4} \pm \sqrt{(7.4 \times 10^{-4})^2 - 4(4)(-5.1 \times 10^{-4})}}{2(4)} \]

\[ x = \frac{-7.4 \times 10^{-4} \pm \sqrt{8.2 \times 10^{-3}}}{8} \]

\[ x = -7.4 \times 10^{-4} \pm 0.090 \]

\[ x = 0.011 \text{ and } -0.011 \]

[HCN] = 2x = 2(0.011) = 0.022 M
17. \( K_c \) for the reaction
\[ \text{Br}_2(\text{g}) + \text{F}_2(\text{g}) \rightleftharpoons 2 \text{BrF}(\text{g}) \]
is 55.3. What are the equilibrium concentrations of all these gases if the initial concentration of \( \text{Br}_2 \) and \( \text{F}_2 \) are both 0.250 M?

\[
\begin{array}{ccc}
\text{Br}_2(\text{g}) + \text{F}_2(\text{g}) & \rightleftharpoons & 2 \text{BrF}(\text{g}) \\
\text{initial} & 0.250 & 0.250 \\
\Delta & -x & -x & 2x \\
\text{equil.} & 0.250-x & 0.250-x & 2x \\
\end{array}
\]

\[
K_c = \frac{[\text{BrF}]^2}{[\text{Br}_2][\text{F}_2]}
\]

\[
\sqrt{55.3} = \sqrt{\frac{(2x)^2}{(0.250 - x)^2}}
\]

\[
7.44 = \frac{2x}{0.250 - x}
\]

\[
7.44(0.250 - x) = 2x
\]

\[
1.86 - 7.44x = 2x
\]

\[
9.44x = 1.86
\]

\[
x = 0.197
\]

\[
[\text{Br}_2] = [\text{F}_2] = 0.250 - x = 0.250 - 0.197 = 0.053 \text{ M}
\]

\[
[\text{BrF}] = 2x = 2(0.197) = 0.394 \text{ M}
\]

18. Carbon dioxide reacts with graphite to form an equilibrium with carbon monoxide. At 850 °C, \( K_c = 10.0 \) for the reaction
\[ \text{CO}_2(\text{g}) + \text{C}_{(\text{graph})} \rightleftharpoons 2 \text{CO}(\text{g}) \]

If 220. g of CO is placed in a 2.50 L flask and heated to 850 °C, what is the mass of \( \text{CO}_2 \) at equilibrium?

\[
\frac{220 \text{ g CO}}{2.50 \text{ L}} = \frac{1 \text{ mol CO}}{28.0 \text{ g CO}} = 3.14 \text{ M CO}
\]

\[
\begin{array}{ccc}
\text{CO}_2(\text{g}) + \text{C}_{(\text{graph})} & \rightleftharpoons & 2 \text{CO}(\text{g}) \\
\text{initial} & 0 & 3.14 \\
\Delta & +x & -2x \\
\text{equil.} & x & 3.14 - 2x \\
\end{array}
\]
\[ K_c = \frac{[CO]^2}{[CO_2]} \]

\[ 10.0 = \frac{(3.14 - 2x)^2}{x} \]

\[ 10.0x = (9.86 - 12.6x + 4x^2) \]

\[ 4x^2 - 22.6x + 9.86 = 0 \]

\[ x = \frac{-(22.6) \pm \sqrt{(-22.6)^2 - 4(4)(9.86)}}{2(4)} \]

\[ x = \frac{22.6 \pm \sqrt{353}}{8} \]

\[ x = 22.6 \pm 18.8 = 0.475 \text{ or } 5.18 \]

\[ x = 0.475 \text{ M CO}_2 \]

\[ \frac{2.50 \text{ L}}{1 \text{ L}} \left( \frac{0.475 \text{ mol}}{1 \text{ mol CO}_2} \right) = 52.2 \text{ g CO}_2 \]

19. Predict whether the equilibrium for the photosynthesis reaction described by the equation

\[ 6 \text{ CO}_2 (g) + 6 \text{ H}_2\text{O} (l) \rightleftharpoons C_6\text{H}_{12}\text{O}_6 (s) + 6 \text{ O}_2 (g) \quad \Delta H^0 = 2801.69 \text{ kJ/mol rxn} \]

would (i) shift right, (ii) shift left, or (iii) remain unchanged if

(a) [CO\textsubscript{2}] is increased
(b) \( P_\text{oxygen} \) is increased
(c) one half of \( C_6H_{12}O_6 \) is removed
(d) total pressure is decreased
(e) temperature is increased
(f) a catalyst is added

(i) shifts right
(ii) shifts left
(iii) remains unchanged
(iii) remains unchanged
(i) shifts right
(iii) remains unchanged

20. Predict whether the equilibrium for the reaction

\[ 2 \text{ SO}_2 (g) + \text{ O}_2 (g) \rightleftharpoons 2 \text{ SO}_3 (g) \quad \Delta H^0 = -198 \text{ kJ/mol} \]

would (i) shift right, (ii) shift left, or (iii) remain unchanged if

(a) [O\textsubscript{2}] is decreased
(b) total pressure is increased
(c) temperature is increased

(ii) shifts left
(i) shifts right
(ii) shifts left
21. Predict whether the equilibrium for the reaction
\[ C_{(g)} + H_2O_{(g)} \rightleftharpoons CO_{(g)} + H_2_{(g)} \quad \Delta H^\circ = 131 \text{ kJ/mol} \]
would (i) shift right, (ii) shift left, or (iii) remain unchanged if

(a) pressure of H$_2$O is decreased (ii) shifts left
(b) total pressure is increased (ii) shifts left
(c) temperature is increased (i) shifts right
(d) C$_{(g)}$ is added (iii) remains unchanged

22. Consider the hypothetical reaction:
\[ A_{(g)} \rightleftharpoons 2 B_{(g)} \]
A flask is charged with 0.75 atm of pure A, after which it is allowed to reach equilibrium. At equilibrium, the partial pressure of A is 0.50 atm. What is the value of K$_p$?

\[
\begin{align*}
A_{(g)} & \rightleftharpoons 2 B_{(g)} \\
\text{init.} & \quad 0.75 \quad 0 \\
\Delta & \quad -0.25 \quad +0.50 \\
\text{equil.} & \quad 0.50 \quad 0.50 \\
K_p & = \frac{(P_B)^2}{P_A} = \frac{(0.50)^2}{0.50} = 0.50
\end{align*}
\]

23. For the reaction:
\[ H_2_{(g)} + I_2_{(g)} \rightleftharpoons 2 HI_{(g)} \]
K$_p$ = 55.3 at 700 K. If 4.00 atm of H$_2$ and 4.00 atm of I$_2$ are mixed initially, what are the equilibrium pressures of all three components at 700 K?

\[
\begin{align*}
\text{init.} & \quad H_2_{(g)} \quad I_2_{(g)} \quad 2 HI_{(g)} \\
& \quad 4.00 \quad 4.00 \quad 0 \\
\Delta & \quad -x \quad -x \quad +2x \\
\text{equil.} & \quad 4.00 - x \quad 4.00 - x \quad 2x \\
K_p & = \frac{(P_{HI})^2}{(P_{H2})(P_{I2})} \\
\sqrt{55.3} & = \sqrt{\frac{(2x)^2}{(4.00 - x)^2}} \\
7.44 & = \frac{2x}{4.00 - x} \\
7.44(4.00 - x) & = 2x \\
9.44x & = 29.8 \\
x & = 3.16 \\
P_{H2} = P_{I2} & = 4.00 - x = 4.00 - 3.16 = 0.84 \text{ atm} \\
P_{HI} & = 2x = 2(3.16) = 6.32 \text{ atm}
\end{align*}
\]
24. For the equilibrium

\[ N_2(g) + 3 H_2(g) \rightleftharpoons 2 NH_3(g) \]

\[ K_p = 4.52 \times 10^{-5} \text{ at } 450^\circ C. \] For each of the mixtures listed here, indicate whether the mixture is at equilibrium at 450 \( ^\circ C \); if it is not at equilibrium, indicate the direction (toward product or toward reactants) in which the mixture must shift to achieve equilibrium.

(a) 105 atm \( NH_3 \), 35 atm \( N_2 \), and 495 atm \( H_2 \)

\[ Q_p = \frac{(P_{NH_3})^2}{(P_{N2})(P_{H2})^3} \]

\[ Q_p = \frac{(105)^2}{(35)(495)^3} = 2.6 \times 10^{-6} \quad Q_p < K_p, \text{ reaction must shift right, toward products} \]

(b) 35 atm \( NH_3 \), no \( N_2 \), and 595 atm \( H_2 \)

\[ Q_p = \frac{(35)^2}{(0)(595)^3} = \text{undefined (infinitely large)} \quad Q_p > K_p, \text{ reaction must shift left, toward reactants} \]

(c) 26 atm \( NH_3 \), 202 atm \( N_2 \), and 42 atm \( H_2 \)

\[ Q_p = \frac{(26)^2}{(202)(42)^3} = 4.52 \times 10^{-5} \quad Q_p = K_p, \text{ reaction is at equilibrium} \]

(d) 105 atm \( NH_3 \), 5.0 atm \( N_2 \), and 55 atm \( H_2 \)

\[ Q_p = \frac{(105)^2}{(5.0)(55)^3} = 0.013 \quad Q_p > K_p, \text{ reaction must shift left, toward reactants} \]

25. For the reaction

\[ NH_4Cl(g) \rightleftharpoons NH_3(g) + HCl(g) \]

\( K_c \) is 0.011 at 330 \( ^\circ C \). Calculate the value of \( K_p \) at 330 \( ^\circ C \).

\[ K_p = K_c(RT)^\Delta n \quad \Delta n = 2 \text{ mol} - 1 \text{ mol} = 1 \text{ mol} \]

\[ K_p = (0.011)[(0.0821)(603)]^1 \]

\[ K_p = 0.54 \]
26. For the reaction

\[ \text{H}_2 \text{(g)} + \text{Cl}_2 \text{(g)} \rightleftharpoons 2 \text{HCl} \text{(g)} \]

\( K_c \) is 193 at 2500. \( K \). What is the value of \( K_p \)?

\[ \Delta n = 2 \text{ mol} - 2 \text{ mol} = 0 \]

\[ K_p = K_c (RT) + \Delta n \]

\[ K_p = 193 \]

27. For the reaction

\[ 2 \text{ BrF}_5 \text{(g)} \rightleftharpoons \text{Br}_2 \text{(g)} + 5 \text{F}_2 \text{(g)} \]

the value of \( K_p \) at 500. °C is 0.042. Calculate \( K_c \) at this same temperature.

\[ K_c = K_p (RT) + \Delta n = 6 \text{ mol} - 2 \text{ mol} = 4 \text{ mol} \]

\[ K_c = (0.042)[(0.0821)(773)]^{-4} \]

\[ K_c = \frac{(0.042)}{[(0.0821)(773)]^{-4}} \approx 0.042 = \frac{0.042}{(63.5)^4} = \frac{0.042}{1.62 \times 10^7} \]

\[ K_c = 2.6 \times 10^{-9} \]

28. For the reaction

\[ \text{CO} \text{(g)} + \text{H}_2\text{O} \text{(g)} \rightleftharpoons \text{CO}_2 \text{(g)} + \text{H}_2 \text{(g)} \]

(a) Calculate \( K_p \) at 25 °C, given \( \Delta G^o = -28.7 \text{ kJ/mol} \).

\[ \Delta G^o = -RT \ln K \]

\[ -28.7 \text{ kJ/mol} = -(8.314 \times 10^{-3} \text{ kJ/mol·K})(298 \text{ K})(\ln K_p) \]

\[ \ln K_p = \frac{28.7 \text{ kJ/mol}}{(8.314 \times 10^{-3} \text{ kJ/mol·K})(298 \text{ K})} \]

\[ \ln K_p = 11.6 \]

\[ K_p = e^{11.6} = 1.09 \times 10^5 \]

(b) Calculate \( K_p \) at 200. °C, given \( \Delta G^o = -21.3 \text{ kJ/mol} \).

\[ -21.3 \text{ kJ/mol} = -(8.314 \times 10^{-3} \text{ kJ/mol·K})(473 \text{ K})(\ln K_p) \]

\[ \ln K_p = 5.42 \]

\[ K_p = e^{5.42} = 225 \]
(c) Given the values of $K_p$ at 25 and 200 °C calculated above, calculate $\Delta H$ for the reaction.

$$\ln \left( \frac{K_2}{K_1} \right) = \frac{\Delta H}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right)$$

$$\ln \left( \frac{225}{1.09 \times 10^5} \right) = \frac{\Delta H}{8.314 \text{ J/mol} \cdot \text{K}} \left( \frac{1}{298 \text{ K}} - \frac{1}{473 \text{ K}} \right)$$

$$-6.16 = \frac{\Delta H}{8.314 \text{ J/mol} \cdot \text{K}} \left( 1.24 \times 10^{-3} \text{ K}^{-1} \right)$$

$$-6.16 = \Delta H \left( 1.49 \times 10^{-4} \text{ mol/J} \right)$$

$$\Delta H = -41,300 \text{ J/mol} = -41.3 \text{ kJ/mol}$$

29. For the reaction

$$2 \text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2 \text{SO}_3(g) \quad \Delta H^\circ = -197.6 \text{ kJ/mol}$$

At 25 °C, the value of $K_p$ is $7.1 \times 10^{24}$. (a) Predict qualitatively whether the value of $K_p$ for this reaction at 500 °C would be greater than, the same as, or less than the value at 25 °C. (b) What is the value of $K_p$ at 500 °C?

(a) The reaction is exothermic. Increasing temperature causes the equilibrium to shift left. As a result, the equilibrium value of $K_p$ at a higher temperature would be less than $K_p$ at 25 °C.

(b)

$$\ln \left( \frac{K_2}{K_1} \right) = \frac{\Delta H}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right)$$

$$\ln \left( \frac{K_2}{7.1 \times 10^{24}} \right) = \frac{-197.6 \text{ kJ/mol}}{8.314 \times 10^{-3} \text{ kJ/mol} \cdot \text{K}} \left( \frac{1}{298 \text{ K}} - \frac{1}{773 \text{ K}} \right)$$

$$\ln \left( \frac{K_2}{7.1 \times 10^{24}} \right) = (-23770 \text{ K})(0.00206 \text{ K} - 1)$$

$$\ln \left( \frac{K_2}{7.1 \times 10^{24}} \right) = -49.0$$

$$\frac{K_2}{7.1 \times 10^{24}} = e^{-49.0} = 5.24 \times 10^{-22}$$

$$K_2 = (5.24 \times 10^{-22})(7.1 \times 10^{24}) = 3700$$