REVIEW QUESTIONS
Chapter 14

1. A mixture of 0.10 mol of NO, 0.050 mol of H₂ and 0.10 mol of H₂O is placed in a 1.0-L flask and allowed to reach equilibrium as shown below:

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

At equilibrium \([\text{NO}] = 0.062 \text{ M}\). Calculate the equilibrium constant, \(K_c\), for this reaction.

\[
\begin{array}{|c|c|c|c|c|}
\hline
& 2 \text{NO} & + & 2 \text{H}_2 & \rightleftharpoons & \text{N}_2 & + & 2 \text{H}_2\text{O} & \text{g} \\
\hline
\text{Initial} & 0.10 \text{ M} & 0.050 \text{ M} & 0 & 0.10 \text{ M} \\
\Delta & -0.038 & -0.038 & +0.019 & +0.038 \\
\text{Equilibrium} & 0.062 & 0.012 & 0.019 & 0.138 \\
\hline
\end{array}
\]

\[
K_c = \frac{[\text{N}_2][\text{H}_2\text{O}]^2}{[\text{NO}]^2[\text{H}_2]^2} = \frac{(0.019)(0.138)^2}{(0.062)^2(0.012)^2} = 650
\]

2. At 700°C, \(K_c = 20.4\) for the reaction shown below:

\[(\text{A})\quad \text{SO}_2 (g) + \frac{1}{2} \text{O}_2 (g) \rightleftharpoons \text{SO}_3 (g)\]

Calculate \(K_c\) and \(K_p\) for the reaction shown below:

\[(\text{B})\quad 2 \text{SO}_2 (g) + \text{O}_2 (g) \rightleftharpoons 2 \text{SO}_3 (g)\]

\[
K_c (\text{B}) = [K_c (\text{A})]^2 = (20.4)^2 = 416
\]

\[
K_p = K_c (RT)^\Delta n = (416) [(0.0821)(973 \text{ K})]^{-1} = 5.21
\]

3. At 100°C, \(K_c = 0.078\) for the following reaction:

\[\text{SO}_2\text{Cl}_2 (g) \rightleftharpoons \text{SO}_2 (g) + \text{Cl}_2 (g)\]

In an equilibrium mixture, \([\text{SO}_2\text{Cl}_2] = 0.136 \text{ M}\) and \([\text{SO}_2] = 0.072 \text{ M}\). What is the concentration of \(\text{Cl}_2\) in the equilibrium mixture?

\[
K_c = \frac{[\text{SO}_2][\text{Cl}_2]}{[\text{SO}_2\text{Cl}_2]} = 0.078
\]

\[
[\text{Cl}_2] = \frac{0.078 [\text{SO}_2\text{Cl}_2]}{[\text{SO}_2]} = \frac{(0.078)(0.136)}{0.072} = 0.15 \text{ M}
\]
4. At 373 K, $K_P = 0.416$ for the equilibrium:

$$2 \text{ NOBr (g)} \rightleftharpoons 2 \text{ NO (g)} + \text{ Br}_2 (g)$$

If the partial pressures of NOBr and NO are equal at equilibrium, what is the partial pressure of Br$_2$?

$$K_P = \frac{P_{\text{NO}}^2}{P_{\text{NOBr}}^2} = 0.416 \quad \text{since } P_{\text{NOBr}} = P_{\text{NO}}$$

$$P_{\text{Br}_2} = 0.416 \text{ atm}$$

5. At 250°C, the reaction

$$\text{PCl}_5 (g) \rightleftharpoons \text{PCl}_3 (g) + \text{Cl}_2 (g)$$

has an equilibrium constant $K_c = 1.80$. If 0.100 mol of PCl$_5$ is added to a 5.00-L flask, what are the concentrations of PCl$_5$, PCl$_3$ and Cl$_2$ at equilibrium at this temperature?

$$[\text{PCl}_5] = \frac{0.100 \text{ mol}}{5.00 \text{ L}} = 0.0200 \text{ M}$$

<table>
<thead>
<tr>
<th></th>
<th>PCl$_5$ (g)</th>
<th>$\rightleftharpoons$</th>
<th>PCl$_3$ (g)</th>
<th>+</th>
<th>Cl$_2$ (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.0200 M</td>
<td>0</td>
<td>0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\Delta$</td>
<td>$-x$</td>
<td>+ $x$</td>
<td>+ $x$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Equilibrium</td>
<td>0.0200 $-x$</td>
<td>$x$</td>
<td>$x$</td>
<td></td>
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</tbody>
</table>

$$K_c = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]} = \frac{x^2}{0.0200 -x} = 1.80$$

$$x^2 + 1.80 \times x - 0.0360 = 0$$

Solving the quadratic equation,

$$x = \frac{-1.80 \pm \sqrt{(1.80)^2 - 4(-0.0360)}}{2} = 0.0198$$

$[\text{PCl}_3] = [\text{Cl}_2] = 0.0198 \text{ M}$

$[\text{PCl}_5] = 0.0200 - 0.0198 = 0.0002 \text{ M}$
6. When 2.00 mol each of hydrogen and iodine are mixed in a 1.00-L flask, 3.50 mol of HI is produced at equilibrium:

\[
\text{H}_2 \ (g) \ + \ \text{I}_2 \ (g) \ \rightleftharpoons \ 2 \ \text{HI} \ (g)
\]

Calculate the equilibrium constant \(K_c\) for this reaction.

| \(\text{H}_2\) (g) \(\rightarrow\) \(\text{I}_2\) (g) \(\rightleftharpoons\) 2 \(\text{HI}\) (g) |
|---|---|---|
| Initial | 2.00 M | 2.00 M | 0 |
| \(\Delta\) | \(-1.75\) | \(-1.75\) | +3.50 |
| Equilibrium | 0.25 | 0.25 | 3.50 |

\[
K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{(3.50)^2}{(0.25)^2} = 196
\]

7. The equilibrium constant for the reaction

\[
\text{SO}_2 \ (g) \ + \ \text{NO}_2 \ (g) \ \rightleftharpoons \ \text{SO}_3 \ (g) \ + \ \text{NO} \ (g)
\]

has a numerical value of 3.00 at a given temperature. 1.50 mol each of \(\text{SO}_2\) and \(\text{NO}_2\) are mixed in a 1.00-L flask and allowed to reach equilibrium. What percent of \(\text{SO}_2\) is converted to product?

| \(\text{SO}_2\) (g) \(\rightarrow\) \(\text{NO}_2\) (g) \(\rightleftharpoons\) \(\text{SO}_3\) (g) \(\rightleftharpoons\) \(\text{NO}\) (g) |
|---|---|---|---|
| Initial | 1.50 M | 1.50 M | 0 | 0 |
| \(\Delta\) | \(-x\) | \(-x\) | +x | +x |
| Equilibrium | 1.50 \(-x\) | 1.50 \(-x\) | x | x |

\[
K_c = \frac{[\text{SO}_3][\text{NO}]}{[\text{SO}_2][\text{NO}_2]} = \frac{x^2}{(1.50 - x)^2} = 3.00
\]

Taking square root of each side,

\[
\frac{x}{1.50 - x} = 1.73
\]

\[
x + 1.73 \times \ x = 2.595
\]

\[
x = 0.951
\]

\[
\% \ = \ \frac{0.951}{1.50} \times 100 = 63.4\%
\]
8. The following equilibrium exists at 1000 °C with $K_C = 2.00$.

$$2 \text{COF}_2 (g) \rightleftharpoons \text{CO}_2 (g) + \text{CF}_4 (g)$$

If a 5.00-L mixture contains 0.145 mol COF$_2$, 0.262 mol of CO$_2$ and 0.074 mol of CF$_4$ at 1000 °C, in which direction will the mixture proceed to reach equilibrium?

$$[\text{COF}_2] = \frac{0.145 \text{ mol}}{5.00 \text{ L}} = 0.0290 \text{ M}$$

$$[\text{CO}_2] = \frac{0.262 \text{ mol}}{5.00 \text{ L}} = 0.0524 \text{ M}$$

$$[\text{CF}_4] = \frac{0.074 \text{ mol}}{5.00 \text{ L}} = 0.0148 \text{ M}$$

$$Q_c = \frac{[\text{CO}_2][\text{CF}_4]}{[\text{COF}_2]^2} = \frac{(0.0524)(0.0148)}{(0.0290)^2} = 0.922$$

Since $Q_c < K_c$, reaction will proceed in the forward direction.

9. Formamide decomposes at high temperatures according to the equation shown below:

$$\text{HCONH}_2 (g) \rightleftharpoons \text{NH}_3 (g) + \text{CO} (g) \quad K_C = 4.84 \text{ at } 400 \text{ K}$$

If 0.186 mol of formamide is placed in a 2.16-L flask and allowed to decompose at 400 K, what will be the total pressure at equilibrium?

$$P_{\text{HCONH}_2} = \frac{nRT}{V} = \frac{(0.186)(0.0821)(400 \text{ K})}{2.16 \text{ L}} = 2.83 \text{ atm}$$

$$K_p = K_c (RT)^\Delta n = 4.84 [(0.0821)(400)]^1 = 159$$

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<thead>
<tr>
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<th>HCONH$_2$ (g) $\rightleftharpoons$ NH$_3$ (g) + CO (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>2.83 atm</td>
</tr>
<tr>
<td>$\Delta$</td>
<td>$-x$ $+x$</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>2.83 $-x$ $x$ $x$</td>
</tr>
</tbody>
</table>

$$K_p = \frac{P_{\text{NH}_3} P_{\text{CO}}}{P_{\text{HCONH}_2}} = \frac{(x)^2}{(2.83 - x)} = 159$$

Solving the quadratic equation, $x = 2.78$

$P_{\text{HCONH}_2} = 2.83 - 2.78 = 0.05 \text{ atm}$

$P_{\text{NH}_3} = P_{\text{CO}} = 2.78 \text{ atm}$

$P_{\text{Total}} = 2.78 + 2.78 + 0.05 = 5.61 \text{ atm}$
10. Predict how each of the following changes affect the amount of H\(_2\) present in an equilibrium mixture in the reaction

\[
3 \text{ Fe (s) } + 4 \text{ H}_2\text{O (g)} \rightleftharpoons \text{ Fe}_3\text{O}_4 \text{ (s) } + 4 \text{ H}_2 \text{ (g)} \quad \Delta H = -150 \text{ kJ}
\]

a) Raising the temperature of the mixture.

Since reaction is exothermic, raising temperature will shift the equilibrium to the left (\(\leftarrow\)) and reduce amount of hydrogen.

b) Adding more H\(_2\)O (g).

Adding more water, will shift the equilibrium to the right (\(\rightarrow\)) and increase the amount of hydrogen.

c) Doubling the volume of the container holding the mixture.

Increasing the volume of the container will reduce the pressure but the equilibrium will not be affected, and the amount of hydrogen will not change.

d) Adding a catalyst.

Adding catalyst does not alter the equilibrium and the amount of hydrogen produced.

11. At 2000 °C the equilibrium constant for the reaction below is \(K_c = 2.4 \times 10^3\). If the initial concentration of NO is 0.500 M, what are the equilibrium concentrations of each substance?

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

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<tr>
<th>\hspace{1cm}</th>
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</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.500 M</td>
</tr>
<tr>
<td>(\Delta)</td>
<td>(-2x)</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>(0.500 - x)</td>
</tr>
</tbody>
</table>

\[
K_c = \frac{[\text{N}_2][\text{O}_2]}{[\text{NO}]^2} = \frac{x^2}{(0.500 - 2x)^3} = 2.4 \times 10^3
\]

Taking square root of each side,

\[
\frac{x}{0.500 - 2x} = 49 \quad \Rightarrow \quad [\text{N}_2] = [\text{O}_2] = x = 0.25 \text{ M}
\]

\[
x + 98x = 24.5 \quad \Rightarrow \quad [\text{NO}] = 0.500 - 2x = 0.500 - 0.494
\]

\[
x = 0.247 \quad \Rightarrow \quad [\text{NO}] = 6.0 \times 10^{-3} \text{ M}
\]
12. The reaction below has an equilibrium constant $K_c = 6.90$. If 0.100 mol of BrCl is placed in a 500-mL flask and allowed to come to equilibrium, what are the equilibrium concentrations of each substance?

$$\text{Br}_2 (g) + \text{Cl}_2 (g) \rightleftharpoons 2 \text{BrCl} (g)$$

$$[\text{BrCl}] = \frac{0.100 \text{ mol}}{0.500 \text{ L}} = 0.200 \text{ M}$$

<table>
<thead>
<tr>
<th></th>
<th>Br$_2$ (g)</th>
<th>Cl$_2$ (g)</th>
<th>2 BrCl (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0</td>
<td>0</td>
<td>0.200 M</td>
</tr>
<tr>
<td>$\Delta$</td>
<td>+x</td>
<td>+x</td>
<td>-2x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>x</td>
<td>x</td>
<td>0.200 - 2x</td>
</tr>
</tbody>
</table>

$$K_c = \frac{[\text{BrCl}]^2}{[\text{Br}_2][\text{Cl}_2]} = \frac{(0.200 - 2x)^2}{x^3} = 6.90$$

Taking square root of each side,

$$\frac{0.200 - 2x}{x} = 2.63$$

$$[B_2] = [C_2] = x = 0.0432 \text{ M}$$

$$2.63x + 2x = 0.200$$

$$[\text{BrCl}] = 0.200 - 2x = 0.200 - 0.0864$$

$$x = 0.0432$$

$$[\text{BrCl}] = 0.114 \text{ M}$$