FOR ALL EQUILIBRIUM PROBLEMS, YOU MUST:
1) Write all equilibrium equations
2) Write all equilibrium concentrations
3) Write all equilibrium expressions

SET A:
1) a) What is the equilibrium constant expression for the reaction:

$$3 \text{ Fe(s)} + 4 \text{ H}_2\text{O(g)} \rightleftharpoons \text{ Fe}_3\text{O}_4\text{(s)} + 4 \text{ H}_2\text{(g)}$$

Ans: $[\text{H}_2]^4/[\text{H}_2\text{O}]^4$

b) The equilibrium constant, $K_c$, for the reaction:

$$2 \text{ NOCl (g)} \rightleftharpoons 2 \text{ NO (g)} + \text{ Cl}_2\text{(g)}$$

What is the equilibrium constant, $K_c$, for the reaction:

$$1/3 \text{ Cl}_2\text{(g)} + 2/3 \text{ NO (g)} \rightleftharpoons 2/3 \text{ NOCl (g)}$$

$$K = \frac{3}{2.4 \times 10^{-7}}$$

Ans: $1.6 \times 10^2$

c) Given the following equilibrium equations and their corresponding equilibrium constants:

\begin{align*}
\text{i)} & \quad 2 \text{ CO}_2\text{(g)} + \text{ H}_2\text{O (g)} \rightleftharpoons 2 \text{ O}_2\text{(g)} + \text{ CH}_2\text{CO (g)} \quad K_c = 6.1 \times 10^8 \\
\text{ii)} & \quad \text{CH}_4\text{(g)} + 2 \text{ O}_2\text{(g)} \rightleftharpoons \text{CO}_2\text{(g)} + 2 \text{ H}_2\text{O (g)} \quad K_c = 1.2 \times 10^{14}
\end{align*}

Find $K_c$ for the reaction: $\text{CH}_4\text{(g)} + \text{CO}_2\text{(g)} \rightleftharpoons \text{CH}_2\text{CO (g)} + \text{H}_2\text{O (g)}$

Setup: \[ Add \text{ equations (i) and (ii). The sum is equation (iii) } \]

\[ K_{\text{overall}} = K_{(i)} \cdot K_{(ii)} = (6.1 \times 10^8)(1.2 \times 10^{14}) \]

Ans: $7.3 \times 10^{22}$

2) A mixture of 9.22 moles of A, 10.11 moles of B, and 27.83 moles of C is placed in a one-liter container at a certain temperature. The reaction is allowed to reach equilibrium. At equilibrium the number of moles of B is 18.32. Calculate the equilibrium constant for the reaction:

$$\text{A (g)} + 2 \text{B (g)} \rightleftharpoons 3 \text{C (g)}$$

\begin{tabular}{|c|c|c|c|}
\hline
\text{Initial conc} & 9.22 & 10.11 & 27.83 \\
\hline
\text{change in conc.} & +4.10 & +8.20 & -12.3 \\
\hline
\text{Equilibrium conc.} & 9.22+4.10 & 18.32 & 27.83-12.31 \\
\hline
\end{tabular}

\[ 8.21 \text{moles "B" reacted} \]

\[ K_c = \frac{[\text{C}]^3}{[\text{A}]^1[\text{B}]^2} = \frac{(15.51)^3}{(13.22)(18.32)^2} \]

Ans: 0.832
3) a) At a certain temperature, $K_c$ is $4.13 \times 10^{-2}$ for the equilibrium:

$$2 \text{IBr} (g) \rightleftharpoons \text{I}_2 (g) + \text{Br}_2 (g)$$

Assume that equilibrium is established at the above temperature by adding only IBr (g) to the reaction flask. What are the concentrations of I$_2$ (g) and Br$_2$ (g) in equilibrium with 0.0124 moles/liter of IBr(g)?

Setup:

\[
\begin{array}{c|c|c|c}
\text{Initial conc.} & \text{Initial} & \text{I}_2 (g) & \text{Br}_2 (g) \\
\hline
\text{Initial} & -2x & +x & +x \\
\text{Final} & -2x = 0.124 & x & x \\
\end{array}
\]

$$K_c = \frac{[\text{I}_2][\text{Br}_2]}{[\text{IBr}]^2}$$

\[
\sqrt{4.13 \times 10^{-2}} = \sqrt{\frac{(x)(x)}{(0.124)^2}}
\]

\[
x = 2.52 \times 10^{-3}
\]

[Ans: $2.52 \times 10^{-3}$ M]

b) What was the initial concentration of IBr before equilibrium was established?

Setup:

\[
\begin{align*}
\text{M}_{\text{Initial}} - 2x &= 0.124 \\
0.124 - 2x &= 0.124 + 2(2.52 \times 10^{-3}) \\
0.124 - 2x &= 0.174
\end{align*}
\]

[Ans: 0.0174]

4) 0.924 mole of A (g) is placed in a 1.00 liter container at 700 °C, where it is 38.8 % dissociated when equilibrium was established.

\[
3 \text{A} (g) \rightleftharpoons 5 \text{B} (g) + 2 \text{C} (g)
\]

What is the value of the equilibrium constant, $K_c$, at the same temperature?

Setup:

\[
\begin{align*}
\text{Initial conc.} & = 0.924 \\
\text{Final conc.} & = 0.359
\end{align*}
\]

\[
\begin{align*}
\text{Balanced} & = \frac{5 \text{B} (g) + 2 \text{C} (g)}{3 \text{A} (g)} \\
0.359 \text{ mole A} & = 0.239 \text{ mole B} + 0.598 \text{ mole C}
\end{align*}
\]

\[
\begin{align*}
\frac{0.359 \text{ mole A}}{3 \text{ mole A}} & = \frac{0.239 \text{ mole B}}{3 \text{ mole A}} + \frac{0.598 \text{ mole C}}{3 \text{ mole A}}
\end{align*}
\]

[Ans: 0.0241]
The equilibrium constant for the reaction:
\[ 2 \text{NO (g)} \rightleftharpoons \text{N}_2 (g) + \text{O}_2 (g) \]
is \(2.60 \times 10^{-3}\) at 1100 °C. If 0.820 mole of NO (g) and 0.223 mole each of N\(_2\) (g) and O\(_2\) (g) are mixed in a 1.00 liter container at 1100 °C, what are the concentrations of NO (g), N\(_2\)(g), and O\(_2\) (g) at equilibrium?

Setup:
\[
Q = \frac{[N_2][O_2]}{[NO]^2} = \frac{(0.223)(0.223)}{(0.820)^2} = 7.40 \times 10^{-2}
\]

\(Q > K\); the reaction is proceeding spontaneously to the left.

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

<table>
<thead>
<tr>
<th>Initial conc</th>
<th>0.820</th>
<th>0.223</th>
<th>0.223</th>
</tr>
</thead>
<tbody>
<tr>
<td>Change conc</td>
<td>+ 2x</td>
<td>-x</td>
<td>-x</td>
</tr>
<tr>
<td>Equi conc</td>
<td>0.820 + 2x</td>
<td>0.223 - x</td>
<td>0.223 - x</td>
</tr>
</tbody>
</table>

\[
K = \frac{[N_2][O_2]}{[NO]^2}
\]

\[
\sqrt{2.60 \times 10^{-3}} = \sqrt{\frac{(0.223 - x)(0.223 - x)}{(0.820 + 2x)^2}}
\]

\[
0.0510 = \frac{(0.223 - x)}{(0.820 + 2x)}
\]

\[
x = 0.164 \text{ M}
\]

\[
[N_2] = 0.223 - x = 0.223 - 0.164 = 0.058 \text{ M}
\]

\[
[NO] = 0.820 + 2x = 0.820 + 2(0.164) = 1.15 \text{ M}
\]

**Ans:**

\[
[NO] = 1.15 \text{ M}
\]

\[
[N_2] = 0.058 \text{ M}
\]

\[
[Cl_2] = 0.058 \text{ M}
\]
SET B:

1) A mixture of 1.16 mole of A, 1.35 mole of B and 0.641 mole of C is placed in a one-liter container at a certain temperature. The reaction was allowed to reach equilibrium. At equilibrium, the number of moles of A is 1.95. Calculate the equilibrium constant, \( K_c \), for the reaction:

\[
\begin{array}{c|c|c|c}
\text{Setup} & 2 \text{A(g)} & \rightleftharpoons & 2 \text{B(g)} + \text{C(g)} \\
\hline
\text{Initial conc.} & 1.16 & 1.35 & 0.641 \\
\text{Change in conc.} & +0.79 & -0.79 & -0.179 \\
\text{Equil. conc.} & 1.95 & (1.35-0.79) & (0.641-0.179) \\
& & 0.56 & 0.466 \\
\end{array}
\]

\[
K_c = \frac{[B]^2[C]}{[A]^2} = \frac{(0.56)^2(0.466)}{(1.95)^2}
\]

\[
= 0.020
\]

Ans: 0.020

2) 0.822 mole of \( \text{SO}_3 \) (g) is placed in a 1.00 liter container at 600 K. 36.7% of the \( \text{SO}_3 \) (g) are decomposed when equilibrium is established.

\[
2 \text{SO}_3 \rightleftharpoons 2 \text{SO}_2 + \text{O}_2
\]

What is the value of the equilibrium constant, \( K_c \), at the same temperature?

\[
\begin{array}{c|c|c|c|c}
\text{Setup} & 2 \text{SO}_3 \text{(g)} & \rightleftharpoons & 2 \text{SO}_2 \text{(g)} + \text{O}_2 \text{(g)} \\
\hline
\text{Initial conc.} & 0.822 & 0 & +0.8 \\
\text{Change in conc.} & -0.302 & +0.302 & +0.302/2 \\
\text{Equil. conc.} & 0.822-0.302 & 0.302 \text{formed} & \frac{(0.302)}{2} \text{formed} \\
\text{initial} & \text{final} & \text{formed} & \text{formed} \\
& & & \\
\end{array}
\]

\[
K = \frac{[\text{SO}_2]^2[\text{O}_2]}{[\text{SO}_3]^2} = \frac{(0.302)^2(0.151)}{(0.520)^2}
\]

\[
= 5.05 \times 10^{-2}
\]

Ans: 5.05 \times 10^{-2}
3) For the equilibrium:
\[ 2 \text{BrCl (g)} \rightleftharpoons \text{Br}_2 (g) + \text{Cl}_2 (g) \]
at 205 °C, the equilibrium constant, \( K_C \), is 0.143. If 1.34 moles each of \( \text{Br}_2 (g) \) and \( \text{Cl}_2 (g) \) are introduced in a container which has a volume of 11.0 liters and allowed to reach equilibrium at 205 °C, what would be the concentrations of \( \text{Br}_2 (g) \), \( \text{Cl}_2 (g) \), and \( \text{BrCl (g)} \) at equilibrium?

Setup:
\[
\begin{array}{c|c|c}
\text{Initial conc.} & \text{Change in conc.} & \text{Equi conc.} \\
\hline
\text{BrCl (g)} & 0 & +2x \\
\text{Br}_2 (g) & \frac{1.34 \text{mol}}{11.0 \text{L}} = 0.122 & \frac{1.34 \text{mol}}{11.0 \text{L}} - x \\
\text{Cl}_2 (g) & 0.122 - x & 0.122 - x \\
\end{array}
\]

\[ K = \frac{[\text{Br}_2][\text{Cl}_2]}{[\text{BrCl}]^2} \]
\[ \sqrt{0.143} = \sqrt{\frac{(0.122-x)(0.122-x)}{(2x)^2}} \]
\[ x = 0.0693 \]
\[ [\text{Br}_2] = [\text{Cl}_2] = 0.122 - x = 0.122 - 0.0693 = 0.0529 \text{M} \]
\[ [\text{BrCl}] = 2x = 2(0.0693) = 0.139 \text{M} \]
Ans:
\[ [\text{Br}_2] = 0.0529 \text{M} \]
\[ [\text{Cl}_2] = 0.0529 \text{M} \]
\[ [\text{BrCl}] = 0.139 \text{M} \]

4) a) What is the numerical value of the equilibrium constant, \( K_C \), for the reaction:
\[ 3 \text{N}_2 (g) + 3 \text{O}_2 (g) \rightleftharpoons 6 \text{NO} (g) \]
if the equilibrium constant for the reaction:
\[ 2 \text{NO} (g) \rightleftharpoons \text{N}_2 (g) + \text{O}_2 (g) \text{ is } 3.5 \times 10^{-6} \]
Setup:
\[ K = \left( \frac{1}{3.5 \times 10^{-6}} \right)^3 = (0.286 \times 10^6)^3 = 2.3 \times 10^{16} \]
Ans:
\[ K = 2.3 \times 10^{16} \]

b) What is the equilibrium constant expression for the reaction:
\[ 2 \text{Ni} (s) + 2 \text{CO}_2 (g) \rightleftharpoons 2 \text{CO} (g) + 2 \text{NiO} (s) \]
Ans:
\[ K = \frac{[\text{CO}]^2}{[\text{CO}_2]^2} \]

5) For the reaction:
\[ 2 \text{NO}_2 (g) + \text{O}_2 (g) \rightleftharpoons 2 \text{NO}_3 (g) \]
at 923 °C, \( K_C \) is 42.5. If 0.0500 mole of \( \text{NO}_2 (g) \), 0.122 mole of \( \text{O}_2 (g) \) and 0.300 mole of \( \text{NO}_3 (g) \) are mixed in a 1.00 liter container at 923 °C, in what direction will the reaction proceed? (Show your calculation to prove that your answer is not a guess.)

Setup:
\[ 2 \text{NO}_2 (g) + \text{O}_2 (g) \rightleftharpoons 2 \text{NO}_3 (g) \]
\[ Q = \frac{(0.300)^2}{(0.0500)^2(0.122)} = 295 \]
\[ Q > K ; \text{The reaction will proceed spontaneously to the left} \]

Answer: \( Q > K_c \) (The reaction will proceed spontaneously to the left)
SET C :

1) For the equilibrium:

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

at 300 K, the equilibrium constant, \( K_c \), is 0.185. If 1.45 moles each of \( \text{N}_2 \) (g) and \( \text{O}_2 \) (g) are introduced in a container that has a volume of 6.00 liters and allowed to reach equilibrium at 300 K, what are the concentrations of \( \text{N}_2 \) (g), \( \text{O}_2 \)(g), and NO (g) at equilibrium?

Setup:

\[ \begin{array}{c|c|c|c}
\text{Initial conc.} & 0 & \frac{1.45}{6.0} = 0.24 & \frac{1.45}{6.0} = 0.24 \\
\text{Change in conc.} & +2x & -x & -x \\
\text{Equi conc.} & 2x & 0.242-x & 0.242-x \\
\end{array} \]

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

\[ \sqrt{0.185} = \sqrt{\frac{(0.242-x)(0.242-x)}{(2x)^2}} \]

\[ x = 0.130 \]

\[ [\text{N}_2] = [\text{O}_2] = 0.242 - x = 0.242 - 0.130 = 0.112 \text{ M} \]

\[ [\text{NO}] = 2x = 2(0.130) = 0.260 \text{ M} \]

\[ \text{Ans:} \ [\text{N}_2] = 0.112 \text{ M} \]

\[ [\text{O}_2] = 0.112 \text{ M} \]

\[ [\text{NO}] = 0.260 \text{ M} \]

2) 1.87 mole of \( \text{A (g)} \) are placed in a 1.00 liter container at 700 K. When the reaction reached equilibrium:

\[ 2 \text{A (g)} \rightleftharpoons 3 \text{B (g)} + \text{C (g)} \]

What is the value of the equilibrium constant, \( K_c \), at the same temperature?

Setup:

\[ \begin{array}{c|c|c|c}
\text{Initial conc.} & 1.87 & 0 & 0 \\
\text{Change in conc.} & -0.80 & +1.21 & +0.405 \\
\text{Equi conc.} & 1.07 & 1.21 & 0.405 \\
\end{array} \]

\[ \frac{43.3}{100} \times 1.87 \text{mole} \text{A} = 0.810 \text{mole} \text{A} \text{ reacted} \]

0.810 mole "A" reacting \( \frac{3 \text{mole B}}{2 \text{mole A}} = 1.21 \text{mole B} \) formed

0.810 mole "A" reacting \( \frac{1 \text{mole C}}{2 \text{mole A}} = 0.405 \text{mole C} \) formed

\[ K = \frac{[\text{B}]^3 [\text{C}]}{[\text{A}]^2} = \frac{(1.21)^3(0.405)}{1.07} \]

\[ = 0.639 \text{ M} \]

\[ \text{Ans: } 0.639 \]
3) a) What is the equilibrium constant expression for the reaction:
\[ 2 \text{Fe} (s) + 2 \text{NO}_2 (g) \rightleftharpoons 2 \text{NO} (g) + 2 \text{FeO} (s) \]

\[ \text{Ans: } K = \frac{[\text{NO}]^2}{[\text{NO}_2]^2} \]

b) What is the numerical value of the equilibrium constant, \( K \), for the reaction:
\[ 2 \text{SO}_3 (g) \rightleftharpoons 2 \text{SO}_2 (g) + \text{O}_2 (g) \]
if the equilibrium constant for the reaction:
\[ \frac{1}{2} \text{O}_2 (g) + \text{SO}_2 (g) \rightleftharpoons \text{SO}_3 (g) \]
is \( 2.7 \times 10^{-3} \)?

\[ \text{Setup: } \]
\[ K = \frac{1}{(2.7 \times 10^{-3})^2} = 1.4 \times 10^5 \]

\[ \text{Ans: } 1.4 \times 10^5 \]

4) A mixture of 3.31 moles of A, 4.33 moles of B, and 5.95 moles of C is placed in a one-liter container at a certain temperature. The reaction was allowed to reach equilibrium. At equilibrium, the number of moles of B is 6.16. Calculate the equilibrium constant, \( K \), for the reaction:
\[ \text{A} (g) \rightleftharpoons 3 \text{B} (g) + 2 \text{C} (g) \]

\[ \text{Setup: } \]
\[ \text{Initial conc. } \begin{array}{c} 3.31 \ \ \ 4.33 \ \ \ 5.95 \\ \text{Equi conc. } \begin{array}{c} \? \ \ \ 6.16\ M \ \ \ \? \end{array} \end{array} \]

\[ 1.83 \text{ moles B formed} \left( \frac{2 \text{ moles C}}{3 \text{ moles B}} \right) = 1.22 \text{ moles C formed} \]

\[ 1.83 \text{ moles B formed} \left( \frac{1 \text{ mole A}}{3 \text{ moles B}} \right) = 0.61 \text{ mole A react} \]

\[ \text{A} (g) \rightleftharpoons 3 \text{B} (g) + 2 \text{C} (g) \]

\[ \begin{array}{c|c|c|c}
\text{Initial conc.} & 3.31 & 4.33 & 5.95 \\
\text{change in conc.} & -0.61 & +1.83 & +1.22 \\
\text{Equil conc.} & 2.70 & 6.16 & 7.17 \\
\end{array} \]

\[ K = \frac{[\text{B}]^3 [\text{C}]^2}{[\text{A}]^3} = \frac{(6.16)^3 (7.17)^2}{2.7} \approx 4.5 \times 10^3 \]

\[ \text{Ans: } 4.5 \times 10^3 \]
For the reaction:

\[ 2 \text{A (g)} \rightleftharpoons \text{B (g)} + \text{C (g)} \]

at 900 \degree C, \( K_c \) is \( 1.40 \times 10^{-3} \). If 0.780 mole of A (g) and 0.244 mole each of B (g) and C (g) are mixed in a 1.00 liter container at 900 \degree C, what are the concentrations of A, B, and C at equilibrium?

Setup:

\[
Q = \frac{[B][C]}{[A]^2} = \frac{(0.244)(0.244)}{(0.780)^2} = 0.0978
\]

\[ Q < K \]

The reaction will proceed spontaneously to the left in order to reach equilibrium.

\[
2 \text{A (g)} \rightleftharpoons \text{B (g)} + \text{C (g)}
\]

<table>
<thead>
<tr>
<th>Initial conc</th>
<th>0.780</th>
<th>0.244</th>
<th>0.244</th>
</tr>
</thead>
<tbody>
<tr>
<td>Change in conc</td>
<td>+2x</td>
<td>-x</td>
<td>-x</td>
</tr>
<tr>
<td>Equil conc.</td>
<td>0.780 + 2x</td>
<td>0.244 - x</td>
<td>0.244 - x</td>
</tr>
</tbody>
</table>

\[
K = \frac{[B][C]}{[A]^2}
\]

\[ \sqrt{1.40 \times 10^{-3}} = \sqrt{\frac{(0.244-x)(0.244-x)}{(0.780+2x)^2}} \]

\[ 3.74 \times 10^{-2} = \frac{(0.244-x)}{(0.780+2x)} \]

\[ x = 0.200 \]

\[
[B] = [C] = 0.244 - x = 0.244 - 0.200 = 0.044 \text{ M}
\]

\[
[A] = 0.780 + 2x = 0.780 + 2(0.200) = 1.180 \text{ M}
\]

Ans:

<table>
<thead>
<tr>
<th>[A]</th>
<th>[B]</th>
<th>[C]</th>
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<tbody>
<tr>
<td>1.180 M</td>
<td>0.044 M</td>
<td>0.044 M</td>
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