1. Consider the following reaction

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

Write the equilibrium expression, \( K_c \).

2. Consider the following reaction

\[ \text{CaCO}_3 (s) \rightleftharpoons \text{CaO} (s) + \text{O}_2 (g) \]

Write the equilibrium expression, \( K_c \).

3. Consider the following reaction

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

Write the equilibrium expression, \( K_p \).

4. Consider the following reaction

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

Write the equilibrium expression, \( K_p \).

5. Consider the decomposition of nitrous oxide, laughing gas,

\[ 2\text{N}_2\text{O} (g) \rightleftharpoons 2 \text{N}_2 (g) + \text{O}_2 (g) \]

At 25°C, \( K_c \) is \( 7.3 \times 10^{34} \).

(a) Based on the information given, what can you say about the rate of decomposition of the reaction?

(b) Based on the information given, does nitrous oxide have a tendency to decompose into nitrogen and oxygen?

(c) What is the value of \( K_p \) for the reaction at 25°C?

6. Consider the following reaction

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

Calculate the value of the equilibrium constant, \( K_c \), for the above system, if 0.1908 moles of \( \text{CO}_2 \), 0.0908 moles of \( \text{H}_2 \), 0.0092 moles of \( \text{CO} \), and 0.0092 moles of \( \text{H}_2\text{O} \) vapour were present in a 2.00 L reaction vessel at equilibrium.
7. Consider the following reaction

\[ C_2H_4 (g) + H_2 (g) \rightleftharpoons C_2H_6 (g) \quad K_c = 0.99 \]

What is the concentration for each substance at equilibrium if the initial concentration of ethene, \( C_2H_4 (g) \), is 0.335 M and that of hydrogen is 0.526 M?

8. Consider the following reaction

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

Determine the value of the equilibrium constant, \( K_c \), for the reaction. Initially, a mixture of 0.100 M NO, 0.050 M H\(_2\), 0.100 M H\(_2\)O was allowed to reach equilibrium (initially there was no N\(_2\)). At equilibrium the concentration of NO was found to be 0.062 M.

9. Consider the following reaction

\[ \text{N}_2\text{O}_4 (g) \rightleftharpoons 2 \text{NO}_2 (g) \]

A reaction flask is charged with 3.00 atm of dinitrogen tetroxide gas and 2.00 atm of nitrogen dioxide gas. At 25\(^\circ\)C, the gases are allowed to reach equilibrium. The pressure of the nitrogen dioxide was found to have decreased by 0.952 atm. Estimate the value of \( K_p \) for this system.

10. Consider the following reaction. The initial concentrations are \([\text{HSO}_4^-] = 0.50 \text{ M}, [\text{H}_3\text{O}^+] = 0.020 \text{ M}, [\text{SO}_4^{2-}] = 0.060 \text{ M}.\]

\[ \text{HSO}_4^- (aq) + \text{H}_2\text{O} (l) \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{SO}_4^{2-} (aq) \quad K = 0.012 \]

(a) Which way would the reaction shift to reach equilibrium?
(b) What are the equilibrium concentrations of the products and reactants.
Answers:

1. \( K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]} \)

2. \( K_c = [\text{O}_2] \)

3. \( K_p = \frac{p_{\text{SO}_2}^2}{p_{\text{SO}_3}^2 p_{\text{O}_2}} \)

4. \( K_p = \frac{p_{\text{H}_2} p_{\text{CO}}}{p_{\text{H}_2\text{O}}} \)

5. (a) Based on the information given, you cannot predict the rate of decomposition of nitrous oxide.
   (b) From the value of the \( K_{eq} \), nitrous oxide has a strong tendency to decompose into nitrogen and oxygen.
   (c) \( K_p = 1.8 \times 10^{36} \)

6. \([\text{CO}_2]\) = 0.1908 mol \( \text{CO}_2/2.00 \text{ L} = 0.0954 \text{ M} \)
   \([\text{H}_2]\) = 0.0454 M
   \([\text{CO}]\) = 0.0046 M
   \([\text{H}_2\text{O}]\) = 0.0046 M

   \[ K = \frac{(0.0046)(0.0046)}{(0.0954)(0.0454)} = 0.0049 \text{ or } 4.9 \times 10^{-3} \]

7. 

<table>
<thead>
<tr>
<th></th>
<th>( \text{C}_2\text{H}_4 )</th>
<th>( \text{H}_2 )</th>
<th>( \text{C}_2\text{H}_6 )</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{I} )</td>
<td>0.335</td>
<td>0.526</td>
<td>0</td>
</tr>
<tr>
<td>( \text{C} )</td>
<td>(-x)</td>
<td>(-x)</td>
<td>(+x)</td>
</tr>
<tr>
<td>( \text{E} )</td>
<td>0.335 (-x)</td>
<td>0.526 (-x)</td>
<td>(+x)</td>
</tr>
</tbody>
</table>

\[ K = \frac{x}{(0.335 - x)(0.526 - x)} = 0.0995 \text{ or } 1.77^* \]

* \( x=1.77 \) is not possible because the concentration of \( \text{C}_2\text{H}_4 \) will result in a negative value.

\[ [\text{C}_2\text{H}_4] = 0.236 \text{ M} \]
\[ [\text{H}_2] = 0.526 - x = 0.526 - 0.0995 = 0.427 \text{ M} \]
\[ [\text{C}_2\text{H}_6] = 0.0995 \text{ M} \]
8.

<table>
<thead>
<tr>
<th></th>
<th>NO</th>
<th>H₂</th>
<th>N₂</th>
<th>H₂O</th>
</tr>
</thead>
<tbody>
<tr>
<td>[I]</td>
<td>0.100</td>
<td>0.0500</td>
<td>0</td>
<td>0.100</td>
</tr>
<tr>
<td>[C]</td>
<td>-2x</td>
<td>-2x</td>
<td>+x</td>
<td>+2x</td>
</tr>
<tr>
<td>[E]</td>
<td>0.062</td>
<td>0.038</td>
<td>0.019</td>
<td>0.138</td>
</tr>
</tbody>
</table>

From ICE table

\[
2x = 0.038
\]

Therefore, substitute for \( x \) and calculate \([E]\) for each species:

\[
\begin{align*}
\text{NO} & \quad \text{H}_2 & \quad \text{N}_2 & \quad \text{H}_2\text{O} \\
\text{[I]} & \quad 0.100 & \quad 0.0500 & \quad 0 & \quad 0.100 \\
\text{[C]} & \quad -0.038 & \quad -0.038 & \quad +0.019 & \quad +0.038 \\
\text{[E]} & \quad 0.062 & \quad 0.012 & \quad 0.019 & \quad 0.138
\end{align*}
\]

\[
K = \frac{(0.019)(0.138)^2}{(0.062)^2(0.012)^2} = 6.5 \times 10^2
\]

9.

<table>
<thead>
<tr>
<th></th>
<th>N₂O₄</th>
<th>NO₂</th>
</tr>
</thead>
<tbody>
<tr>
<td>[I]</td>
<td>3.00</td>
<td>2.00</td>
</tr>
<tr>
<td>[C]</td>
<td>+x</td>
<td>-2x = -0.952</td>
</tr>
<tr>
<td>[E]</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

From ICE table

\[
x = 0.952/2
\]

Therefore, substitute for \( x \) and calculate \([E]\) for each species:

\[
\begin{align*}
\text{N}_2\text{O}_4 & \quad \text{NO}_2 \\
\text{[I]} & \quad 3.00 & \quad 2.00 \\
\text{[C]} & \quad +0.476 & \quad -0.952 \\
\text{[E]} & \quad 3.476 & \quad 1.048
\end{align*}
\]

\[
K = \frac{(1.048)^2}{(3.476)} = 0.316
\]
10. (a) Use the trial $K_{eq}$, $Q$, to determine the reaction direction.

$$Q = \frac{(0.020)(0.060)}{0.50} = 0.0024$$

$Q < K_{eq}$, therefore, equilibrium will shift to the right to produce more products.

(b) 

<table>
<thead>
<tr>
<th></th>
<th>HSO$_4^-$</th>
<th>H$_3$O$^+$</th>
<th>SO$_4^{2-}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>[I]</td>
<td>0.50</td>
<td>0.020</td>
<td>0.060</td>
</tr>
<tr>
<td>[C]</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>[E]</td>
<td>0.50 -x</td>
<td>0.020+x</td>
<td>0.060+x</td>
</tr>
</tbody>
</table>

$$K = \frac{(0.020 + x)(0.060 + x)}{(0.50 - x)}$$

To solve, need to use the quadratic equation

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

$x = 0.0372$ or $-0.129$ *

For $x = 0.0372$,

$$[HSO_4^-] = 0.46 \text{ M}; [H_3O^+] = 0.057 \text{ M}; [SO_4^{2-}] = 0.097 \text{ M}$$

* For $x = -0.129$,

$$[HSO_4^-] = 0.63 \text{ M}; [H_3O^+] = -0.109 \text{ M}; [SO_4^{2-}] = -0.069 \text{ M}$$

it yields negative concentrations.

Therefore, the correct equilibrium concentrations are:

$$[HSO_4^-] = 0.46 \text{ M}; [H_3O^+] = 0.057 \text{ M}; [SO_4^{2-}] = 0.097 \text{ M}$$