Worksheet 19 - Le Chatelier's Principle

Le Chatelier's Principle states that if a stress is applied to a system at equilibrium, the system will adjust, to partially offset the stress and will reach a new state of equilibrium.

The "stresses" that can be applied to the system include changes in concentration, pressure, volume and temperature.

- Changes in concentration of either reactants or products will change the value of Q, the reaction quotient.

  Adding more reactant will drive the forward reaction.
  Adding more product will drive the reverse reaction.

  Removal of reactants or products will shift the equilibrium in the direction needed to produce more of the substance that was removed.

Solids and pure liquids can not change concentration. Changing the amount of these present in the system will have no effect on equilibrium, unless they are removed entirely.

Catalysts are species that speed up the rate of a reaction. However, they speed up both the forward and reverse reactions, leaving K unchanged.

- Changes in pressure and volume will affect the equilibrium of reactions involving gases. In the previous worksheet, you calculated the effect of pressure on gaseous systems.

  Increasing the volume of the system (lowering the pressure) drives the equilibrium toward the state with the larger number of moles of gas.

  Increasing the pressure of the system by adding an inert gas has no effect on the equilibrium of the system.

- Changes in temperature change the value of the equilibrium constant, K. This is a fairly complex process, but can be thought of in simple terms, using Le Chatelier's Principle.

  Heat can be treated as a product in exothermic reactions (ΔH < 0) and as a reactant in endothermic reactions (ΔH > 0).

  Raising the temperature of a reaction can be thought of as adding heat. In endothermic reactions (heat = reactant) this will drive the forward reaction. In exothermic reactions (heat = product) and raising the temperature will drive the reverse reaction.
1. Predict what will happen when the reaction volume is decreased in each of the following, after balancing the reactions.

   a) \( \frac{6}{2} \text{ CO}_2 (g) + \frac{6}{2} \text{ H}_2\text{O} (l) \rightleftharpoons \frac{1}{6} \text{ C}_6\text{H}_{12}\text{O}_6 (s) + \frac{6}{6} \text{ O}_2 (g) \)  

   b) \( \frac{1}{1} \text{ PCl}_3 (g) \rightleftharpoons \frac{1}{1} \text{ PCl}_3 (g) + \frac{1}{1} \text{ Cl}_2(g) \)  

   c) \( \frac{1}{1} \text{ H}_2 (g) + \frac{1}{1} \text{ CO}_2 (g) \rightleftharpoons \frac{1}{1} \text{ H}_2\text{O} (g) + \frac{1}{1} \text{ CO (g)} \)  

2. Balance the exothermic reaction below.

\[
2 \text{ NO}_2 (g) \rightleftharpoons \frac{1}{1} \text{ N}_2\text{O}_4 (g) \quad \Delta H = -58.0 \text{ kJ}
\]

Predict the effect of each of the following changes on this system at equilibrium (drive forward reaction, drive reverse reaction, no effect).

   a) add \( \text{N}_2\text{O}_4 \)  
   b) remove \( \text{NO}_2 \)  
   c) increase the volume  
   d) decrease the temperature  
   e) Add \( \text{N}_2 \) no effect

3. The equilibrium constant, \( K_p \), for the reaction

\[
\frac{1}{1} \text{ H}_2 (g) + \frac{1}{1} \text{ I}_2 (s) \rightleftharpoons \frac{2}{1} \text{ HI (g)} \quad K_p = 0.35
\]

is 0.35 at 25°C. Decide if each of the following mixtures is at equilibrium, at 25°C. If it is not at equilibrium, decide which way the reaction will proceed to reach equilibrium.

   a) \( P_{\text{H}_2} = 0.10 \text{ atm}, P_{\text{HI}} = 0.90 \text{ atm} \) and there is solid \( \text{I}_2 \) present \( Q > K_p \)  

\[
(0.10)^2 \frac{1}{0.1} = 8.1
\]

   b) \( P_{\text{H}_2} = 0.55 \text{ atm}, P_{\text{HI}} = 0.44 \text{ atm} \), and there is solid \( \text{I}_2 \) present \( Q = K_p \) at equil.  

\[
(0.55)^2 \frac{1}{0.44} = 1.35
\]

   c) \( P_{\text{H}_2} = 0.25 \text{ atm}, P_{\text{HI}} = 0.15 \text{ atm} \) and there is solid \( \text{I}_2 \) present \( Q < K_p \)  

\[
(0.25)^2 \frac{1}{0.15} = 0.92
\]
4. 10.0 g of the weakly soluble solid Ag₂CO₃ is added to 1.00 L of water and the following equilibrium is established:

\[ \text{Ag}_2\text{CO}_3 (s) \rightleftharpoons 2 \text{Ag}^+ (aq) + \text{CO}_3^{2-} (aq) \]

For this reaction, at 25° C, \( K = 8.2 \times 10^{-12} \). This is usually called \( K_{\text{sp}} \) or the solubility constant, which is a measure of how much of an "insoluble" solid will dissolve.

Write the expression for the solubility constant in terms of [reactants] and [products].

\[ K_{\text{sp}} = \frac{[\text{CO}_3^{2-}][\text{Ag}^+]^2}{[\text{Ag}_2\text{CO}_3]} = 8.2 \times 10^{-12} \]

Make an ICE table for this reaction and calculate the equilibrium concentrations of Ag⁺ and CO₃²⁻.

<table>
<thead>
<tr>
<th></th>
<th>[Ag⁺] (M)</th>
<th>[CO₃²⁻] (M)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>+ 2x</td>
<td>+ x</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>2x</td>
<td>x</td>
</tr>
</tbody>
</table>

Substitute these values into the expression for \( K_{\text{sp}} \) (above) and solve for:

\[ [\text{Ag}^+]_e = 2.54 \times 10^{-4} \text{ M} \quad [\text{CO}_3^{2-}]_e = 1.27 \times 10^{-5} \text{ M} \]

In which direction will each of the following changes force the equilibrium?

a) adding 1.00 L of water (Hint: calculate a value for Q)

\[ Q = \left( \frac{2.54 \times 10^{-4} \text{ M}}{2 \text{ L}} \right)^2 \left( \frac{1.27 \times 10^{-5} \text{ M}}{2 \text{ L}} \right) = 1.02 \times 10^{-12} \quad Q < K \]

b) adding 1.00 moles of NaCl (Hint: remember the solubility rules)

\[ \text{Ag}^+ + \text{Cl}^- \rightarrow \text{AgCl(s)} \]

c) adding 10.0 g of Ag₂CO₃(s)  

no effect, it doesn't change [Ag₂CO₃]

d) adding 0.5 moles of AgNO₃(s) (Hint: more solubility rules)

soluble salt in excess [Ag⁺]
5. 0.100 moles of HCN, hydrocyanic acid are added to 1.00 L of water. The following reaction takes place:

$$\text{HCN (aq) + H}_2\text{O} \rightleftharpoons \text{H}^+ \text{(aq) + CN}^- \text{(aq)}$$

For this reaction at 25°C, \( K = 6.2 \times 10^{-10} \).

Write the expression for \( K \) in terms of [reactants] and [products].

$$K = \frac{[\text{H}^+][\text{CN}^-]}{[\text{HCN}]} = 6.2 \times 10^{-10} = \frac{\chi^2}{100 - \chi} = \frac{7.87 \times 10^{-6}}{97} = \frac{[\text{H}^+]}{[\text{CN}^-]}$$

Make an ICE table for this reaction.

<table>
<thead>
<tr>
<th></th>
<th>[HCN] (M)</th>
<th>[H(^+)] (M)</th>
<th>[CN(^-)] (M)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>( \frac{1}{100} )</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>( -\chi )</td>
<td>( +\chi )</td>
<td>( +\chi )</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>( \frac{1}{100 - \chi} )</td>
<td>( \chi )</td>
<td>( \chi )</td>
</tr>
</tbody>
</table>

Since \( K \) is so small, the [HCN] will not change appreciably (5% rule).

Put the equilibrium concentration values into the expression for \( K \) (above) and solve for:

$$[\text{H}^+]_e = 7.87 \times 10^{-6} \quad [\text{CN}^-]_e = 7.87 \times 10^{-6} \quad [\text{HCN}]_e = 0.100 \text{ M}$$

In which direction will each of the following changes force the equilibrium?

a) adding 1.00 L of water (Hint: calculate a value for \( Q \))

$$Q = \frac{(3.94 \times 10^{-6})^2}{0.05} = 1.6 \times 10^{-10} < K$$

b) adding 1.00 moles of HCl (aq)

increase [H\(^+\)]

$$Q = 3.09 \times 10^{-10} < K$$

b) adding 0.1 moles of NaCN (s), a soluble salt

increase [CN\(^-\)]

$$Q = 3.09 \times 10^{-10} < K$$