• Calculate the standard free-energy change for the following reaction at 298 K.

\[ 2\text{Au(s)} + 3\text{Ca}^{2+} (1.0 \text{ M}) \rightarrow 2\text{Au}^{3+}(1.0 \text{ M}) + 3\text{Ca(s)} \]

The reduction half cell reactions and $E^0$ values are:

\[ \text{Au}^{3+}(aq) + 3e^- \rightarrow \text{Au}(s) \quad E^0 = +1.50 \text{ V} \]
\[ \text{Ca}^{2+}(aq) + 2e^- \rightarrow \text{Ca}(s) \quad E^0 = -2.87 \text{ V} \]

In the reaction, Au is being oxidized and so the overall cell potential is:

\[ E^0 = (-2.87) - (+1.50) = -4.37 \text{ V} \]

The reaction involves 6 electrons so, using $\Delta G^0 = -nFE^0$:

\[ \Delta G^0 = - (6) \times (96485) \times (-4.37) = +2530000 \text{ J mol}^{-1} = +2.53 \times 10^3 \text{ kJ mol}^{-1} \]

Answer: $+2.53 \times 10^3 \text{ kJ mol}^{-1}$

Complete and balance the following equation for the reaction between iron(II) ions and permanganate ions in an acidic solution.

\[ \text{Fe}^{2+} + \text{MnO}_4^- \rightarrow \text{Fe}^{3+} + \text{Mn}^{2+} \]

In acid, the relevant half cells are:

\[ \text{Fe}^{2+}(aq) \rightarrow \text{Fe}^{3+}(aq) + e^- \]
\[ \text{MnO}_4^-(aq) + 8\text{H}^+(aq) + 5e^- \rightarrow \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l) \]

Giving an overall reaction:

\[ 5\text{Fe}^{2+}(aq) + \text{MnO}_4^-(aq) + 8\text{H}^+(aq) \rightarrow 5\text{Fe}^{3+}(aq) + \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l) \]

ANSWER CONTINUES ON THE NEXT PAGE
What is the value of the equilibrium constant for the following reaction at 298 K?

\[ 2\text{Fe}^{3+}(\text{aq}) + \text{Sn}(s) \rightleftharpoons \text{Sn}^{2+}(\text{aq}) + 2\text{Fe}^{2+}(\text{aq}) \]

The reduction half cell reactions and \( E^0 \) values are:

- \( \text{Fe}^{3+}(\text{aq}) + e^- \rightarrow \text{Fe}^{2+}(\text{aq}) \quad E^0 = +0.77 \text{ V} \)
- \( \text{Sn}^{2+}(\text{aq}) + 2e^- \rightarrow \text{Sn}(s) \quad E^0 = -0.14 \text{ V} \)

In the reaction, Sn is being oxidized and so the overall cell potential is:

\[ E^0 = (+0.77) - (-0.14) = +0.91 \text{ V} \]

The reaction involves 2 electrons so, using \( E^0 = \frac{RT}{nF} \ln K \):

\[ \ln K = E^0 \times \frac{nF}{RT} = (+0.91) \times \left( \frac{2 \times 96485}{8.314 \times 298} \right) = 70.9 \]

\[ K = e^{70.9} = 6.05 \times 10^{30} \]

Answer: \( 6.05 \times 10^{30} \)
Consider the results of the following set of experiments studying the rate of the chemical reaction: 

\[ 2A + B \rightarrow 3C + D \]

<table>
<thead>
<tr>
<th>Experiment #</th>
<th>initial [A] / M</th>
<th>initial [B] / M</th>
<th>Rate / M hr(^{-1})</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.240</td>
<td>0.120</td>
<td>2.00</td>
</tr>
<tr>
<td>2</td>
<td>0.120</td>
<td>0.120</td>
<td>0.500</td>
</tr>
<tr>
<td>3</td>
<td>0.240</td>
<td>0.060</td>
<td>1.00</td>
</tr>
</tbody>
</table>

Write the rate law expression.

Between experiment 1 and 2, [B] is kept constant. [A] is halved which causes the rate to be reduced by a factor of four. The rate is second order with respect to [A].

Between experiment 1 and 3, [A] is kept constant. [B] is halved which causes the rate to halve. The rate is first order with respect to [B]. Thus,

\[
\text{rate} \propto [A]^2[B] = k[A]^2[B]
\]

Calculate the rate constant, \(k\), with units.

Using experiment 1 and rate = \(k[A]^2[B]\):

\[
(2.00) = k \times (0.240)^2 \times (0.120) \quad \text{so } k = 289
\]

\[
(M \text{ hr}^{-1}) = (\text{units of } k) \times (M)^2 \times (M) \quad \text{so the units of } k \text{ are } M^{-2} \text{ hr}^{-1}
\]

\[
k = 289 \text{ M}^{-2} \text{ hr}^{-1}
\]

What is the rate of the reaction when [A] is 0.0140 M and [B] is 1.35 M?

Using rate = \(289[A]^2[B]\), the rate is:

\[
\text{rate} = (289) \times (0.0140)^2 \times (1.35) = 0.0766 \text{ M hr}^{-1} = 7.66 \times 10^{-3} \text{ M hr}^{-1}
\]

Rate = \(7.66 \times 10^{-3} \text{ M hr}^{-1}\)
Consider the results of the following set of experiments studying the rate of the reaction of nitric oxide with hydrogen at 1280 °C.

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

<table>
<thead>
<tr>
<th>Experiment #</th>
<th>[NO] / M</th>
<th>[H(_2)] / M</th>
<th>Initial Rate / M s(^{-1})</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>5.0 \times 10^{-3}</td>
<td>2.0 \times 10^{-3}</td>
<td>1.3 \times 10^{-5}</td>
</tr>
<tr>
<td>2</td>
<td>1.0 \times 10^{-2}</td>
<td>2.0 \times 10^{-3}</td>
<td>5.2 \times 10^{-5}</td>
</tr>
<tr>
<td>3</td>
<td>1.0 \times 10^{-2}</td>
<td>4.0 \times 10^{-3}</td>
<td>1.0 \times 10^{-4}</td>
</tr>
</tbody>
</table>

Write the rate law expression.

Between experiments 1 and 2, [H\(_2\)] is kept constant. Doubling [NO] (from 5.0 \times 10^{-3} to 1.0 \times 10^{-2} M) leads to the rate quadrupling. The reaction is second order with respect to [NO].

Between experiments 2 and 3, [NO] is kept constant. Doubling [H\(_2\)] (from 2.0 \times 10^{-3} to 4.0 \times 10^{-3} M) leads to the rate doubling. The reaction is first order with respect to [H\(_2\)]. Thus,

\[
\text{rate } \propto [\text{NO}]^2[\text{H}_2] = k[\text{NO}]^2[\text{H}_2]
\]

Calculate the rate constant, \(k\). Include units in your answer.

Using experiment 1 and rate = \(k[\text{NO}]^2[\text{H}_2]\):

\[
(1.3 \times 10^{-5}) = k \times (5.0 \times 10^{-3})^2 \times (2.0 \times 10^{-3})
\]

so \(k = 260\) (M s\(^{-1}\)) = (units of k) \times (M)^2 \times (M)

so the units of \(k\) are M\(^{-2}\) s\(^{-1}\)

\[k = 260 \text{ M}^{-2} \text{ s}^{-1}\]

What is the rate of the reaction when [NO] is 1.2 \times 10^{-2} M and [H\(_2\)] is 6.0 \times 10^{-3} M?

Using rate = 260[NO]\(^2\)[H\(_2\)]:

\[
\text{rate} = (260) \times (1.2 \times 10^{-2})^2 \times (6.0 \times 10^{-3}) = 2.2 \times 10^{-4} \text{ M s}^{-1}
\]

Rate = 2.2 \times 10^{-4} M s\(^{-1}\)
What is the value of the equilibrium constant for the following reaction at 298 K?

\[ 2\text{Fe}^{3+}(aq) + \text{Sn(s)} \rightleftharpoons \text{Sn}^{2+}(aq) + 2\text{Fe}^{2+}(aq) \]

The reduction half cell reactions and \( E^0 \) values are:

- \( \text{Fe}^{3+}(aq) + e^- \rightarrow \text{Fe}^{2+}(aq) \quad E^0 = +0.77 \text{ V} \)
- \( \text{Sn}^{2+}(aq) + 2e^- \rightarrow \text{Sn}(s) \quad E^0 = -0.14 \text{ V} \)

In the reaction, Sn is being oxidized and so the overall cell potential is:

\[ E^0 = (+0.77) - (-0.14) = +0.91 \text{ V} \]

The reaction involves 2 electrons so, using \( E^0 = \frac{RT}{nF} \ln K \):

\[ \ln K = E^0 \times \frac{nF}{RT} = (+0.91) \times \left( \frac{2 \times 96485}{8.314 \times 298} \right) = 70.9 \]

\[ K = e^{70.9} = 6.05 \times 10^{30} \]

Answer: \( 6.05 \times 10^{30} \)