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 \(= k[A]^2[B]\):

\[
(2.00 \text{ M hr}^{-1}) = k \times (0.240 \text{ M})^2 \times (0.120 \text{ M}) \quad \text{so } k = 289 \text{ M}^2 \text{ hr}^{-1}
\]

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

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

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

Using rate \(= (289 \text{ M}^2 \text{ hr}^{-1})[A]^2[B]\), the rate is:

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

Rate = \(7.66 \times 10^{-3} \text{ M hr}^{-1}\)
• Consider the following balanced redox reaction.

\[ 2\text{In}(s) + 3\text{MnO}_2(s) + 12\text{H}^+(aq) \rightarrow 2\text{In}^{3+}(aq) + 3\text{Mn}^{2+}(aq) + 6\text{H}_2\text{O} \]

If \( E^o = 1.568 \text{ V} \), what would be the measured potential of this cell at 298 K at the following concentrations?

- \([\text{H}^+(aq)] = 0.25 \text{ M}\)
- \([\text{In}^{3+}(aq)] = 0.20 \text{ M}\)
- \([\text{Mn}^{2+}(aq)] = 0.42 \text{ M}\)

\[ Q = \frac{[\text{In}^{3+}(aq)]^2[\text{Mn}^{2+}]^3}{[\text{H}^+(aq)]^{12}} \]

The reaction quotient \( Q \) is given by:

The cell reaction involves a 6e\(^-\) process (2In \( \rightarrow \) 2In(III) and 3Mn(IV) \( \rightarrow \) 3Mn(II)). The cell potential can be calculated using the Nernst equation,

\[ E = E^o - \frac{RT}{nF} \ln Q \]

\[ = (1.568 \text{ V}) - \frac{(8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (298 \text{ K})}{(6) \times (96485 \text{ C mol}^{-1})} \ln \left[ \frac{(0.20)^2(0.42)^3}{(0.25)^{12}} \right] = 1.52 \text{ V} \]

Answer: 1.52 V

• What is the value of the equilibrium constant for the following reaction at 298 K?

\( \text{Cu}^{2+}(aq) + \text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu}(s) \)

Relevant electrode potentials can be found on the data page.

The standard reduction potentials for \( \text{Cu}^{2+}(aq) / \text{Cu}(s) \) and \( \text{Zn}^{2+}(aq) / \text{Zn}(s) \) are +0.34 V and -0.76 V respectively. The \( \text{Zn}^{2+}(aq) / \text{Zn} \) potential is the less positive and is reversed. The cell potential is therefore:

\[ E^o_{\text{cell}} = ((0.34) - (-0.76)) \text{ V} = 1.10 \text{ V} \]

The equilibrium constant for this 2e\(^-\) process can be calculated using \( E^o_{\text{cell}} \)

\[ K = e^{\frac{RT}{nF} \ln K} : \]

\[ \ln K = \frac{(2) \times (96485 \text{ C mol}^{-1})}{(8.314 \text{ JK}^{-1} \text{ mol}^{-1}) \times (298 \text{ K})} \times (1.10) = 85.7 \]

\[ K = e^{85.7} = 1.62 \times 10^{37} \]

Answer: \( 1.62 \times 10^{37} \)
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\textsubscript{2}] / M</th>
<th>Initial Rate / M s\textsuperscript{-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\textsubscript{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\textsubscript{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\textsubscript{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 = \( k[\text{NO}]^2[\text{H}_2] \):

\[
(1.3 \times 10^{-5} \text{ M s}^{-1}) = k \times (5.0 \times 10^{-3} \text{ M})^2 \times (2.0 \times 10^{-3} \text{ M}) \quad \text{so } k = 260 \text{ M}^{-2} \text{s}^{-1}
\]

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

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

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

Using rate = (260 M\textsuperscript{-2} s\textsuperscript{-1})[NO]\textsuperscript{2}[H\textsubscript{2}]:

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

Rate = 2.2 \times 10^{-4} \text{ 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:

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

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

\[ \ln K = E^0 \times \frac{nF}{RT} = (+0.91 \text{ V}) \times \left( \frac{2 \times 96485 \text{ C mol}^{-1}}{8.314 \text{ J K}^{-1} \text{ mol}^{-1} \times 298 \text{ K}} \right) = 70.9 \]

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

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