IODINE CLOCK REACTION KINETICS

PRELAB ASSIGNMENT

Calculate the concentration of H\(_2\)O\(_2\) immediately after mixing together the contents of vessels A, B and C as specified for RUN 1 on page 2. The volume of 2 drops of starch is negligible.

Answer_________________

INTRODUCTION

In this experiment we will determine the effect of a reactant concentration and temperature on the rate of a chemical reaction. A reaction is chosen which proceeds conveniently slowly near room temperature and which can be measured easily by a dramatic color change. This reaction is the oxidation of iodide ion (I\(^-\)) to molecular iodine (I\(_2\)) by hydrogen peroxide (H\(_2\)O\(_2\)):

\[
\text{H}_2\text{O}_2 + 2 \text{H}^+ + 2 \text{I}^- \xrightarrow{\text{slow}} \text{I}_2 + 2 \text{H}_2\text{O}
\]  

(1)

As this reaction proceeds, the colorless reactants gradually develop a brown color due to the product I\(_2\). Because of the difficulty of timing the appearance of the I\(_2\), we make use of another much faster reaction in the same solution to mark the progress of the slow reaction:

\[
\text{I}_2 + 2 \text{S}_2\text{O}_3^{2-} \xrightarrow{\text{fast}} 2 \text{I}^- + \text{S}_4\text{O}_6^{2-}
\]

(2)

Reaction 2 is so fast that I\(_2\) produced by reaction 1 is consumed instantaneously by the thiosulfate (S\(_2\)O\(_3^{2-}\)), so that the I\(_2\) color cannot develop. Because both S\(_2\)O\(_3^{2-}\) and S\(_4\)O\(_6^{2-}\) are colorless, the solution remains colorless. However, we do not add enough thiosulfate to react with all the I\(_2\) that will be formed from reaction 1. By this device the reaction solution stays colorless until the instant at which all the thiosulfate is consumed, and then free I\(_2\) begins to appear. We time the reaction from the initial mixing until the appearance of I\(_2\). In order to help see this appearance we add starch indicator which forms an intensely colored dark complex with I\(_2\) and signals the appearance of I\(_2\) by a dramatic color change.

We know the initial concentration of S\(_2\)O\(_3^{2-}\) and measure the time interval for this S\(_2\)O\(_3^{2-}\) to be consumed. These quantities determine the rate of the slow reaction 1. By repeating the experiment with different concentrations of H\(_2\)O\(_2\) and different temperatures, we can determine the effect of H\(_2\)O\(_2\) concentration and temperature on the rate of reaction 1.
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IN THE LABORATORY

- Thoroughly clean and label glassware as follows:
  - Two 125 mL Erlenmeyer flasks, both labeled “A reaction Mixture”
  - Two small beakers, both labeled “B Peroxide”
  - Two small beakers, both labeled “C Thiosulfate”
  - One 25 or 50 mL graduated cylinder for water

Here is a summary of the setups for the five runs. Instructions for each are given below.

<table>
<thead>
<tr>
<th></th>
<th>Run 1</th>
<th>Run 2</th>
<th>Run 3</th>
<th>Run 4</th>
<th>Run 5</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.10 M HI into flask A, mL</td>
<td>50</td>
<td>50</td>
<td>50</td>
<td>50</td>
<td>50</td>
</tr>
<tr>
<td>0.20 M H$_2$O$_2$ into beaker B, mL</td>
<td>20</td>
<td>30</td>
<td>40</td>
<td>30</td>
<td>30</td>
</tr>
<tr>
<td>Deionized water into beaker B, mL</td>
<td>20</td>
<td>10</td>
<td>0</td>
<td>10</td>
<td>10</td>
</tr>
<tr>
<td>0.15 M S$_2$O$_3^{2-}$ into beaker C, mL</td>
<td>10</td>
<td>10</td>
<td>10</td>
<td>10</td>
<td>10</td>
</tr>
<tr>
<td>Starch into beaker C, drops</td>
<td>2</td>
<td>2</td>
<td>2</td>
<td>2</td>
<td>2</td>
</tr>
<tr>
<td>Temperature</td>
<td>room</td>
<td>room</td>
<td>room</td>
<td>hot</td>
<td>cold</td>
</tr>
</tbody>
</table>

In order to avoid contaminating reagents use only the graduated cylinder placed near each reagent bottle to measure that reagent. Leave those cylinders in place.

PRELUDE TO RUN 5

- Set up an ice bath.

- Dispense reagents into one set of glassware A, B, and C as shown above for Run 5. Set these in the ice bath for later use.
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RUN 1

- Dispense reagents for Run 1 into another set of glassware.
- Have a clock or watch with a second hand ready. Then in quick succession:
  - Add the contents of both beakers B and C into flask A.
  - Swirl the mixture in A.
  - Start timing as soon as beaker B is emptied. This is time $t_i$.
- Place a thermometer in flask A but DO NOT STIR WITH IT.
- Swirl occasionally, and perhaps two minutes or more later the reaction mixture will suddenly turn dark. Note the time as the final time $t_f$. Record the temperature and the time interval $\Delta t = t_f - t_i$ in seconds.

RUN 2

- Thoroughly rinse flask A just used, and then dispense reagents.
- Follow the procedure for Run 1. This one is faster.

RUN 3

- Thoroughly rinse flask A, and then dispense reagents.
- Follow the procedure for Run 1. This is even faster than 2.

RUN 4

- Thoroughly rinse flask A, and dispense reagents.
- Place flask A and beaker C (but not B) on a hot plate and warm these up to between 45 °C and 50 °C. Use a thermometer to monitor the temperature in flask A only.
- When flask A is over 45 °C, remove both from the hot plate. Mix B and C into A and swirl.
- Leave the thermometer in the reaction mixture and record the temperature when the color turns. This one is very fast.

RUN 5

The temperature of solution A in the ice bath should be below 10 °C.
**Iodine Clock**

- Leave flask A in the ice bath. Pour beakers B and C into A. Swirl to mix. Note the temperature.

- Swirl flask A from time to time while the reaction is proceeding. Record the time when the color turns. This one is very slow.

- If the final temperature is different than the initial temperature, record the average.

On the data sheet, fill in the following quantities in the table for each of the five runs:

Row 1: \( \Delta t \) is your measured time interval in seconds.

Row 2: Temp, \( ^\circ\text{C} \) is the temperature of the reaction mixture.

Row 3: \([\text{H}_2\text{O}_2]\) \(_i\) is the \( \text{H}_2\text{O}_2 \) molarity in flask A at time \( t_i \).

Calculate the total volume ignoring the 2 drops of starch.

Row 4: \([\text{S}_2\text{O}_3^{2-}]\) \(_i\) is the thiosulfate molarity in flask A at time \( t_i \).

Row 5: \( \Delta \text{[H}_2\text{O}_2] \) is the change in peroxide concentration during \( \Delta t \).

This calculation is illustrated below.

\[
\begin{align*}
\text{[H}_2\text{O}_2]_i & \quad \uparrow \quad \Delta \text{[H}_2\text{O}_2] \quad \downarrow \\
\text{[H}_2\text{O}_2] & \\
\text{[S}_2\text{O}_3^{2-}]_i & \\
0 & \quad t_i \quad \Delta t \quad t_f \\
\text{[I}_2] & \quad \text{time}
\end{align*}
\]

Between \( t_i \) and \( t_f \) the following changes occur:

- Thiosulfate molarity drops from \([\text{S}_2\text{O}_3^{2-}]\) \(_i\) to zero.

- The \( \text{I}_2 \) moles per liter consumed is \([\text{S}_2\text{O}_3^{2-}]\) \(_i\)/2 (see the balancing coefficients in Reaction 2).

- The \( \text{I}_2 \) consumed in Reaction 2 is produced in Reaction 1.
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- For each $\text{I}_2$ mole produced, a mole of $\text{H}_2\text{O}_2$ is consumed (see balancing coefficients in Reaction 1).

- Thus the $\text{H}_2\text{O}_2$ loss, $\Delta[\text{H}_2\text{O}_2]$, is $[\text{S}_2\text{O}_3^{2-}]/2$.

Row 6: $Rate_i$ is the initial rate of the reaction: $\Delta[\text{H}_2\text{O}_2]/\Delta t$

Row 7: $k$ is the specific rate constant in the first-order rate law

$$Rate = k \ [\text{H}_2\text{O}_2]$$

Use the initial rate and $[\text{H}_2\text{O}_2]_i$ to calculate a $k$ value for

$$k = Rate_i / [\text{H}_2\text{O}_2]_i$$

Row 8: Calculate log $k$.

Row 9: Calculate $1/T$ from the temperature. $T$ is the absolute temperature in Kelvin, where $T$ (K) = $T$ (°C) + 273.15.

- Plot log $k$ vs. $1/T$ on the attached graph paper. Select proper scales in order to maximize space utilization.

- Draw a straight line coming as close as possible to most of the five points.

- Calculate the slope of the line. Considering that log $k$ values are dimensionless, what are the units of the slope?

The increase in the rate constant with temperature is described by the Arrhenius equation

$$\log k = \left(\frac{-E_a}{19.2}\right) \left(\frac{1}{T}\right) + b$$

Here the constant 19.2 is $2.303 \ R$ with units of J mol$^{-1}$ K$^{-1}$, $E_a$ is the activation energy in J mol$^{-1}$ and $b$ is an irrelevant constant.

- Use your line slope to calculate $E_a$.

Your report consists of the data sheet and the graph.
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Name: ______________________________________  Section ______________________

Reagents Concentrations: 0.10 M HI, 0.20 M H₂O₂, 0.15 M S₃O₅²⁻

**IODINE CLOCK REACTION KINETICS**

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</tr>
</thead>
<tbody>
<tr>
<td>1. ( t ), seconds</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2. Temp, °C</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3. ([H₂O₂]_i), M</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>4. ([S₃O₅²⁻]_i), M</td>
<td></td>
<td></td>
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<td></td>
<td></td>
</tr>
<tr>
<td>5. ( \Delta[H₂O₂] ), M</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>6. ( Rate_i ), M/s</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>7. ( k ), s⁻¹</td>
<td></td>
<td></td>
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<td></td>
</tr>
<tr>
<td>8. log ( k )</td>
<td></td>
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<td></td>
<td></td>
</tr>
<tr>
<td>9. ( 1/T ), K⁻¹</td>
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<td></td>
</tr>
</tbody>
</table>

The slope of your log \( k \) vs. 1/T straight line is ____________.

The units of this slope are ________________.

The activation energy \( E_a \) is ________________.