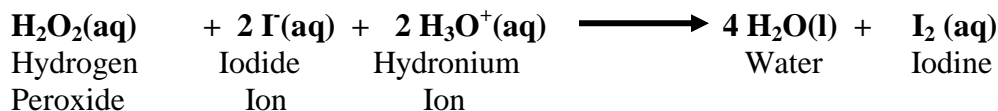


EXPERIMENT 1  
THE RATE LAW

**PURPOSE:**

To determine the Rate Law for the following chemical reaction:

**PRINCIPLES:****Concept of the experiment**

The rate of a chemical reaction is a measure of how fast a chemical reaction occurs.

The Reaction Rate can be determined experimentally by measuring the change in concentration of the reactants or products, divided by the change in time.

During the course of the reaction, the reactants are used up to produce products.

As a result, the **concentration of the reactants ( $\text{H}_2\text{O}_2$ ,  $\text{I}^-$  and  $\text{H}_3\text{O}^+$ ) decreases** and the **concentration of the products ( $\text{H}_2\text{O}$  and  $\text{I}_2$ ) increases** accordingly.

In this experiment the Reaction Rate will be calculated by dividing the experimentally determined increase in concentration of one of the products (elemental iodine,  $\text{I}_2$ ), by the corresponding time interval:

$$\text{Rate} = \frac{\Delta \text{I}_2}{\Delta t}$$

The experimental determination of the increase in concentration of iodine ( $\text{I}_2$ ), during a corresponding time interval, can be easily monitored, since the presence of even small amounts of iodine can be detected by virtue of the intensely blue colored complex formed between iodine and starch.

One creative way of measuring the rate of formation of iodine is to couple the reaction, in which the iodine is formed (Reaction 1), with a much faster reaction that consumes all of the iodine (Reaction 2). This is achieved by adding a small amount of sodium thiosulfate ( $\text{Na}_2\text{S}_2\text{O}_3$ ) and starch indicator is to the reaction mixture. The thiosulfate ion ( $\text{S}_2\text{O}_3^{2-}$ ) does not react with any of the reactants, but it does react rapidly with iodine according to the Reaction 2.



Reaction 2 immediately consumes the  $\text{I}_2$  generated in the first reaction, until all of the  $\text{S}_2\text{O}_3^{2-}$  (thiosulfate ion) is used up. When all of the  $\text{S}_2\text{O}_3^{2-}$  is consumed,  $\text{I}_2$  builds up and reacts with starch to form the deep blue Starch-Iodine Complex.

EXPERIMENT 1  
THE RATE LAW

The appearance of the deep-blue complex tells us that at this point in time ( $t_{\text{color}}$ ), sufficient  $\text{I}_2$  has been produced by Reaction 1 to use up all of the  $\text{S}_2\text{O}_3^{2-}$  (thiosulfate ion) originally added. From this observation, we can calculate the concentration of  $\text{I}_2$  (Iodine) produced by Reaction 1, by noting that according to the stoichiometry of Reaction 2:

**1 mole of  $\text{I}_2$  reacts with 2 moles of  $\text{S}_2\text{O}_3^{2-}$  (thiosulfate), or**  
 **$\frac{1}{2}$  mole of  $\text{I}_2$  reacts with 1 mole of  $\text{S}_2\text{O}_3^{2-}$  (thiosulfate).**

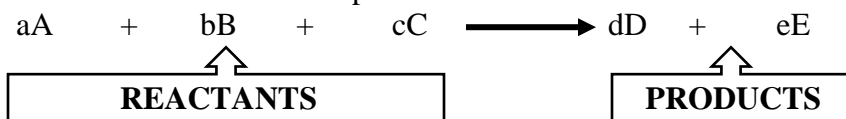
It follows that at the time the deep blue color ( $t_{\text{color}}$ ) appears:

$\text{Reaction Rate} = \frac{\Delta [\text{I}_2]}{\Delta t} = \frac{\Delta [\text{S}_2\text{O}_3^{2-}] \text{ (originally added and used up)}}{2 \Delta t}$
--

**PART I: REACTION ORDERS**

The initial rate of a reaction depends on the initial concentration of one or more of the Reactants, raised to a power commonly referred to as “Reaction Order”

Consider a reaction with a pattern similar to the reaction studied in this experiment.



As long as the reverse reaction is negligibly slow, the relationship between the Rate of Reaction and the Concentration of Reactants can be expressed by a mathematical expression called the Rate Law:

$$\text{Rate} = k [\text{A}]^m [\text{B}]^n [\text{C}]^p$$

where:

“**k**” is a proportionality constant called the Rate Constant,

“**m**” is the Reaction Order with respect to Reactant A,

“**n**” is the Reaction Order with respect to Reactant B, and

“**p**” is the Reaction Order with respect to Reactant C.

The values of the reaction orders (“**m**”, “**n**” and “**p**”) determine the dependence of the reaction rate on the concentrations of the respective reactants. Reaction orders commonly have one of the following values: -1, 0, 0.5, 1 and 2. Reaction orders (such as “**m**”, “**n**” and “**p**”) can be determined only experimentally and they are NOT related to the coefficients of the balanced chemical equation (such as “**a**”, “**b**” and “**c**”). The examples below illustrate how the Reaction Orders (“**m**”, “**n**” and “**p**”) can be determined experimentally for a reaction involving three reactants (A, B and C), such as the reaction studied in this experiment.

The reaction orders (“**m**”, “**n**” and “**p**”) with respect to the three reactants (**A, B and C**) are determined by measuring the initial rate for several reaction runs with varying concentrations of one reactant (for example **A**) independently of the concentration of the other reactants (**B and C**). This allows us to determine the dependence of the rate on the concentration of [A] and the numerical value of the Reaction Order with respect to reactant A (“**m**”).

EXPERIMENT 1  
THE RATE LAW

Temperature	Reaction Run	[A] M	[B] M	[C] M	Initial Rate M/s
Room Temperature	1	<b>0.0600</b>	0.0450	0.0200	<b><math>9.66 \times 10^{-6}</math></b>
Room Temperature	2	<b>0.0300</b>	0.0450	0.0200	<b><math>4.83 \times 10^{-6}</math></b>
		x 1/2	Constant	Constant	x 1/2

Between the first two experiments (1 & 2), the concentration of [A] is halved, the Reaction Rate is also halved, while the concentrations of [B] and [C] stay constant. It follows that the initial rate is directly proportional to the initial concentration of [A].

**The reaction is therefore of the first order with respect to [A] and  $m = 1$ .**

Experimental data seldom provide such obvious and easy to interpret numbers.

Reaction orders calculated from experimental data can be calculated by substituting the initial concentrations of the reactants and the corresponding initial rates into a ratio of the rate laws.

To calculate “m”

$$\frac{\text{Rate 1}}{\text{Rate 2}} = \frac{k [A]^m [B]^n [C]^p}{k [A]^m [B]^n [C]^p}$$

$$\frac{9.66 \times 10^{-6} \text{ M/s}}{4.83 \times 10^{-6} \text{ M/s}} = \frac{k [0.0600 \text{ M}]^m [0.0450 \text{ M}]^n [0.0200 \text{ M}]^p}{k [0.0300 \text{ M}]^m [0.0450 \text{ M}]^n [0.0200 \text{ M}]^p}$$

Canceling out similar terms and doing the calculations yields:  **$2.00 = 2.00^m$**

To find “m”, we take the log of both sides of the equation and solve for “m”

$$\log 2.00 = \log (2.00^m) \quad \log 2.00 = m \log 2.00 \quad m = \frac{\log 2.00}{\log 2.00} = \frac{0.301}{0.301} = 1$$

EXPERIMENT 1  
THE RATE LAW

The Reaction Order with respect to [B], (“n”) can be calculated in a similar manner by using the data obtained for another set of Reaction Runs (For example 1 & 3).

Temperature	Reaction Run	[A] M	[B] M	[C] M	Initial Rate M/s
Room Temperature	1	0.0500	<b>0.0150</b>	0.0200	<b><math>1.33 \times 10^{-6}</math></b>
Room Temperature	3	0.0500	<b>0.0300</b>	0.0200	<b><math>2.66 \times 10^{-6}</math></b>
		Constant	x 2	Constant	x 2

Between experiment 1 and experiment 3, the concentration of [B] is doubled, the Reaction Rate is also doubled, while the concentrations of [A] and [C] stay constant. It follows that the initial rate is directly proportional to the initial concentration of [B]. The reaction is therefore of the first order with respect to [B] and  $n = 1$ . It follows that for this reaction the Rate Law is:

$$\text{Rate Law} = k [A]^1[B]^1[C]^p \quad \text{where “p” is still unknown.}$$

It follows that in order to determine the Reaction Orders (“m”, “n” and “p”) the **(1) Initial Concentrations of the Reactants** and the **(2) Initial Reaction Rates** must be known..

1. **Initial Concentrations of the Reactants.**

As explained above, In order to determine the Rate Law for the reaction being studied, the initial concentrations of the 3 reactants must be known.

a. The Initial Concentrations of  $I^-(aq)$  and  $H_3O^+(aq)$  can be calculated since the exact molarity of the reagents (stock solutions) used is known.

- **KI (0.050 M  $I^-$ )**
- **$H_3O^+$  (0.050 M  $HC_2H_3O_2/Na C_2H_3O_2$  buffer and 0.30 M  $HC_2H_3O_2$ )**

b. The Initial Concentration of  $H_2O_2$  cannot be calculated since the exact molarity of the reagent (stock solution) used, is known only approximately (between 0.8 M and 0.9 M). This is so because the  $H_2O_2$  stock solution cannot be kept at a previously determined accurate concentration, since its concentration changes in time, due to decomposition, when exposed to the environment.

In order to obtain the accurate concentration of the  $H_2O_2$  stock solution, its concentration must be determined experimentally.

This is done by a procedure referred to as “**standardization**”, described in the procedure section of this experiment.

EXPERIMENT 1  
THE RATE LAW

## 2. Initial Reaction Rates

If we know the initial concentration of the thiosulfate ion ( $S_2O_3^{2-}$ ), that is the same for each experiment, and remember that it is all used up when the color of the solution changes, then we know that half the amount of  $I_2$  was also consumed in Reaction 2. This means that the change in the  $I_2$  concentration is equal to half the initial concentration of the thiosulfate ion ( $S_2O_3^{2-}$ ) and it remains constant throughout all the experiments. In essence, the same amount of  $I_2$  is produced in each experiment (Reaction Run) at the time the color changes, but it takes varying times for this to occur, since the  $t_{\text{color}}$  depends on the reaction conditions (**concentration of reactants** and **temperature**).

Please note that in this experiment all Reaction Runs are performed at Room Temperature (R.T.). It follows that the time the deep blue color appears ( $t_{\text{color}}$ ) and, consequently the calculated Initial Reaction Rates, depend solely on the Initial Concentration of the three Reactants.

The time of the color change ( $t_{\text{color}}$ ) is also the time that passed during the reaction ( $\Delta t$ ). It follows that the rate of any of the reactions can be calculated as:

$$\text{Rate} = \frac{\Delta [I_2]}{\Delta t} = \frac{[I_2] \text{ produced at } t_{\text{color}}}{t_{\text{color}}}$$

## PART II: THE RATE CONSTANT

The Rate Constant “k” is characteristic for every reaction and it is independent of the concentrations of the reactants. However, “k” is temperature dependent.

The rate constant, “k” can be calculated, after the reaction orders and the initial reaction rates have been determined experimentally, from any of the reactions that have been run at the same temperature (for example at room temperature, commonly abbreviated R.T.).

In this experiment data for four reaction runs with different initial concentrations of reactants but at the same temperature (R.T.) provide the experimental data needed for the calculation of the Rate Constant “k” at R.T. In order to obtain an accurate value for the rate constant, “k” will be calculated for every one of the four experiments (Reaction Runs) and the average value of the rate constant will be reported in the final expression of the Rate Law. An example of calculation of “k” at room temperature is given below for a reaction with a pattern similar to the reaction studied in this experiment.



To calculate “k”:

- First solve the Rate Law for k:

$$\text{Rate} = k [A]^m [B]^n [C]^p$$

Solve for k  $\Longrightarrow$

$$k = \frac{\text{Rate}}{[A]^m [B]^n [C]^p}$$

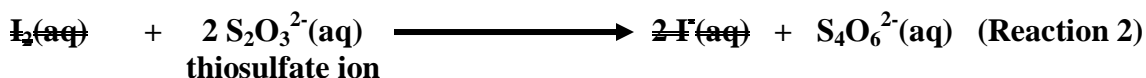
- Next, calculate “k” by substituting the **Initial Concentrations of the Reactants**, the experimentally determined **Reaction Orders** and the **Initial Rate of Reaction**

EXPERIMENT 1  
THE RATE LAW

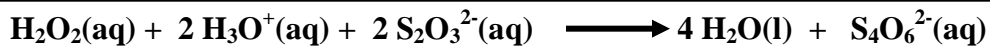
**PROCEDURE:****PART A: STANDARDIZATION OF THE H<sub>2</sub>O<sub>2</sub> STOCK SOLUTION**

In order to obtain the accurate concentration of the H<sub>2</sub>O<sub>2</sub> stock solution, the solution needs to be standardized, by reacting it with a known amount of another reagent. The entire procedure by which the molarity of a solution of one substance (H<sub>2</sub>O<sub>2</sub>) is obtained from an accurately known amount of another substance (Na<sub>2</sub>S<sub>2</sub>O<sub>3</sub>, sodium thiosulfate) is called **standardization**.

The standardization is based on the sequence of the two reactions, discussed on page 1:



Summing up and cancelling out identical terms yields the following **NET REACTION**:



The procedure used here involves the titration of 1.00 mL of H<sub>2</sub>O<sub>2</sub> solution with a solution Na<sub>2</sub>S<sub>2</sub>O<sub>3</sub> of known concentration (0.050 M).

By determining the exact volume of 0.050 M Na<sub>2</sub>S<sub>2</sub>O<sub>3</sub>(aq) required to consume all of the H<sub>2</sub>O<sub>2</sub> contained in 1.00 mL solution of H<sub>2</sub>O<sub>2</sub> of unknown concentration, the molarity of the H<sub>2</sub>O<sub>2</sub> can be calculated. Note that according to the stoichiometry of the Net Reaction, **1 mole of H<sub>2</sub>O<sub>2</sub>** reacts with **2 moles S<sub>2</sub>O<sub>3</sub><sup>2-</sup>**

$$\text{Number of moles of H}_2\text{O}_2 = \frac{\text{Number of moles of S}_2\text{O}_3^{2-}}{2}$$

Each student performs one titration and combines his/her titration data with the data obtained by his/her partner.

The experimental procedure involves the following steps:

**1. Preparation of the buret for the titration**

- a. Obtain a 50.0 mL buret and clean it thoroughly with deionized water.
- b. In order to ensure that the buret is in good working condition follow the following steps:
  - fill up the buret with D.I. water above the 0.00 mL mark,
  - open the stopcock
  - drain all of the D.I. water while checking the stopcock for leaks.
- c. Rinse the buret with three portions of about 5 mL of the 0.050 M Na<sub>2</sub>S<sub>2</sub>O<sub>3</sub> solution, coating the barrel each time before emptying out the solution.
- d. Fill the buret with the 0.050 M Na<sub>2</sub>S<sub>2</sub>O<sub>3</sub> solution a little above the “0.00 mL” line.
- e. Use a buret clamp to clamp the buret
- f. Open the stopcock and drain the 0.050 M Na<sub>2</sub>S<sub>2</sub>O<sub>3</sub> solution in order to completely fill the tip of the buret and flush out any air bubbles caught in the tip.

*\* If you have difficulty getting rid of the air bubbles, ask your instructor to assist you.*

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EXPERIMENT 1  
THE RATE LAW

- g. Set the level of the titrant (0.050 M  $\text{Na}_2\text{S}_2\text{O}_3$  solution) at the 0.00 mL mark.
- Record the initial buret reading. It is not absolutely necessary to refill it to exactly 0.00 mL; however it is necessary to record exactly the starting volume to the nearest  $\pm 0.01$  mL.

**2. Preparation of the  $\text{H}_2\text{O}_2$  stock solution for standardization**

- Add 25 mL of Deionized Water to a clean (does not need to be dry) 125 mL Erlenmeyer flask.
- A buret containing the  $\text{H}_2\text{O}_2$  stock solution is set up for classroom use. Deliver exactly 1.00 mL of  $\text{H}_2\text{O}_2$  stock solution from this buret into the 125 mL Erlenmeyer flask.
- Add 10 mL of 2 M  $\text{H}_2\text{SO}_4$  solution to the solution contained in the flask. (Use a 10 mL graduated cylinder to measure out the of 2 M  $\text{H}_2\text{SO}_4$ )  
After this addition, immediately wash out your 10 mL graduated cylinder (tap water, followed with D.I Water), so that you have it available for Step 3.c. that follows below.
- Weigh out on the centigram balance 1 g of solid KI and add it to the solution contained in the flask
- Add 3 drops of 3% ammonium molybdate catalyst to the solution.

**3. The titration**

- Swirl the solution until the solid KI dissolves.
- Immediately titrate the brown iodine solution that forms with 0.050 M  $\text{Na}_2\text{S}_2\text{O}_3$  (sodium thiosulfate) solution until the brown color begins to fade to yellow.
- Add 2 mL of 0.1% starch indicator and titrate to the disappearance of the blue color of the starch-iodine complex.  
(Use a 10 mL graduated cylinder to measure out the 2 mL of 0.1% starch indicator )
- Continue titrating with 0.50 M  $\text{Na}_2\text{S}_2\text{O}_3$  until you reach the end point.  
The end point is signaled by the disappearance of the blue color of the starch-iodine complex) and it should require 32 to 36 mL of titrant (0.050 M  $\text{Na}_2\text{S}_2\text{O}_3$  , sodium thiosulfate) solution.
- If any of the titrations was faulty, repeat it.

EXPERIMENT 1  
THE RATE LAW

**PART B: REACTION RATE MEASUREMENTS**

The table below summarizes the volumes of stock solutions to be used in making the reaction mixtures for four reactions run at **room temperature**

**TABLE I**

Run Nr.	Temp ( <sup>o</sup> C)	50 mL Flask	250 mL Reaction Flask					
		Reactant <u>A</u> ≈ 0.8 M H <sub>2</sub> O <sub>2</sub>	Reactant <u>B</u> 0.050 M KI	Reactant <u>C</u> [????????] M H <sub>3</sub> O <sup>+</sup>		D.I. Water	0.1% Starch Indicator	0.050 M Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub>
		(mL)	(mL)	Reactant C <sub>1</sub> 0.30 M HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub> (mL)	Reactant C <sub>2</sub> 0.050 M HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub> / NaC <sub>2</sub> H <sub>3</sub> O <sub>2</sub> buffer (mL)	(mL)	(mL)	(mL)
1	R.T.	10	25	0	30	75	5	5
2	R.T.	5	25	0	30	80	5	5
3	R.T.	10	50	0	30	50	5	5
4	R.T.	10	25	45	30	30	5	5

Note that the total volume of all four reaction mixtures is **150 mL**.

**1. Preparing the delivery of the components for the four reaction runs**

- **Solutions to be delivered to the 250 mL Reaction Flask from graduated cylinders**  
Each stock solution should be delivered from the same **clean** (washed, rinsed with D.I water and drained) and **labeled** graduated cylinder, designated for the delivery of that specific stock solution. Label the graduated cylinders as indicated below:

50 mL Flask	250 mL Reaction Flask					
0.80 M H <sub>2</sub> O <sub>2</sub> Reactant A	0.050 M KI Reactant B	0.30 M HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub> Reactant C <sub>1</sub>	0.050 M Buffer Reactant C <sub>2</sub>	D.I. Water	0.1% Starch	0.050 M Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub>
Delivered From buret	50 mL Graduated Cylinder	50 mL Graduated Cylinder	100 mL Graduated Cylinder	100 mL Graduated Cylinder	10 mL Graduated Cylinder	10 mL Graduated Cylinder



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EXPERIMENT 1  
THE RATE LAW

- **Solution to be delivered to the 50 mL Flask from a buret ( $\text{H}_2\text{O}_2$ )**
  - Obtain a 50.0 mL buret and clean it thoroughly with deionized water.
  - Rinse the buret with three portions of about 5 mL of the  $\text{H}_2\text{O}_2$  stock solution, coating the barrel each time before emptying out the solution.
  - Fill the buret with the  $\text{H}_2\text{O}_2$  stock solution a little above the “0” line.
  - Use a buret clamp to clamp the buret
  - Open the stopcock and drain the  $\text{H}_2\text{O}_2$  stock solution in order to completely fill the tip of the buret and flush out any of the air bubbles caught in the tip.
  - Check the stopcock for leaks.
  - Set the level of the  $\text{H}_2\text{O}_2$  stock solution at the 0.00 mL mark.

**2. The actual procedure for carrying out each reaction run is the same for all runs and it is described below for Reaction Run Nr. 1**

- **Filling and stirring the solution in the 250 mL Reaction Flask**
  - Deliver all the stock solutions listed in Table I, under the heading “250 mL Reaction Flask”, into this flask, by using the appropriately labeled graduated cylinders for volume measurement and delivery.
  - Stirring the solution:
    - Place a magnetic stirring bar in the 250 mL flask
    - Place the flask on a magnetic stirring plate and center it.
    - Set the magnetic stirring plate so, as to provide a gentle and uniform mixing of the solution in the flask
    - Stir the solution with the stirring magnet for about three minutes.  
Note that improper and/or insufficient mixing is the major source of error in this experiment. Do not spill the contents of the flask.  
If you do, discard the solution and make a new solution.
- **Filling the 50 mL Flask**
  - Wash, rinse with D.I. water and drain the D.I. water from a 50 mL flask.
  - Deliver the accurately measured appropriate volume of previously standardized  $\text{H}_2\text{O}_2$  stock solution, from the buret into the 50 mL flask.

EXPERIMENT 1  
THE RATE LAW

- **Mixing and Timing**

Before, during and after the addition of the solution from the 50 mL flask into the 250 mL flask, make sure that the 250 mL flask is kept on the magnetic stirring plate (well centered) and its contents are continuously being stirred (**Figures on next page**).

**Please note that the magnetic stirring plate is NOT illustrated in the figures below.**

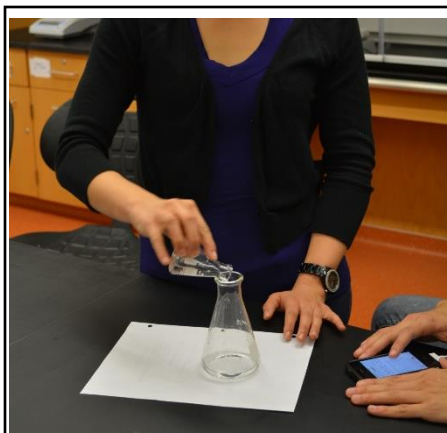


Figure 1

Have a timer available. Pour the solution from the 50 – mL flask into the 250 mL reaction flask, while continuously stirring the reaction mixture in the 250 - mL flask



Figure 2

While you continue swirling the reaction mixture, your partner should start the timer at the time the solution from the 50 mL flask drained completely into the 250 mL reaction flask. You should leave the 50 – mL flask over the mouth of the reaction flask, as shown above and continue swirling the reaction mixture.



Figure 3

Continue swirling the solution until the blue color appears. Note and record the time at which the blue color appears, to the nearest second.

- **Pour the contents of the 250 mL Reaction Flask, through a funnel, into the appropriate waste container. In this manner, the magnetic stirring bar will be caught in the funnel rather than accidentally being dumped in the waste container.**

- Rinse the 250 mL Reaction Flask and the magnetic stirring bar.
- There is no need to rinse the 50 mL flask since you will be using it for the same solution ( $\text{H}_2\text{O}_2$  stock solution)

**3. Repeat the experiment with the other three mixtures in the table**

- The 250 mL Reaction Flask should be rinsed with tap water followed by deionized water and drained between experiments.
- Read and record the time it took for the blue color to appear for every reaction to the nearest second.

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EXPERIMENT 1  
THE RATE LAW

REPORT FORM

NAME: \_\_\_\_\_ Date: \_\_\_\_\_ Partner: \_\_\_\_\_

**PART I: STANDARDIZATION OF THE H<sub>2</sub>O<sub>2</sub> SOLUTION**

**DATA TABLE I**  
**Standardization of the H<sub>2</sub>O<sub>2</sub> Solution**

Molarity of the Na<sub>2</sub>S<sub>2</sub>O<sub>3</sub> solution: \_\_\_\_\_ M

	Team Member 1	Team Member 2
<b>Name:</b>		
Final Buret Reading (mL)		
Initial Buret Reading (mL)		
mL of Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub> added		
L of Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub> added		
Moles of Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub> added		
Moles of H <sub>2</sub> O <sub>2</sub> used		
Volume of H <sub>2</sub> O <sub>2</sub> (mL) used		
Volume of H <sub>2</sub> O <sub>2</sub> (L) used		
Molarity of H <sub>2</sub> O <sub>2</sub> (mol/L)		
Average Molarity of H <sub>2</sub> O <sub>2</sub> (mol/L)		

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EXPERIMENT 1  
THE RATE LAW

**PART II: INITIAL CONCENTRATION OF REACTANTS**

**1. Concentration calculations for H<sub>2</sub>O<sub>2</sub> and I<sup>-</sup>**

For each kinetic run, calculate the initial concentration of the reactants:

**Reactant A: H<sub>2</sub>O<sub>2</sub>**

**Reactant B: I<sup>-</sup>**

**Reactant C: H<sub>3</sub>O<sup>+</sup>**

Since the reaction takes place in a total volume of 150.00 mL, this volume must be taken into account in calculating the initial concentration of the two reactants.

For example, in Run 1, since the 10.00 mL of  $\approx 0.8 \text{ M H}_2\text{O}_2$  added reacts in a total volume of 150.00 mL, the initial concentration of  $[\text{H}_2\text{O}_2]_0$  can be calculated as follows:

$$[\text{H}_2\text{O}_2]_0 = \text{???? M H}_2\text{O}_2 \frac{10.00 \text{ mL}}{150.0 \text{ mL}} = \text{???? M H}_2\text{O}_2$$

Use the molarity of the H<sub>2</sub>O<sub>2</sub> stock solution, obtained by standardization

Similarly, in Run 1, the initial concentration of  $[\text{I}^-]_0$  is calculated as follows:

$$[\text{I}^-]_0 = 0.050 \text{ M I}^- \frac{25.0 \text{ mL}}{150.0 \text{ mL}} = 0.0083 \text{ M I}^-$$

On the next page, carry out similar calculations for all other initial values of these two reactants and complete the appropriate columns in DATA TABLE III.

You are required to:

- Show all your calculations
- Include units in your calculations, and
- Express all measured quantities (including your answer) in the appropriate number of significant figures.

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EXPERIMENT 1  
THE RATE LAW

**DATA TABLE II**  
**Concentration Calculations for Initial Concentration of Reactants**

	Reactant A: $[\text{H}_2\text{O}_2]_0$	Reactant B: $[\text{I}]_0$
Run 1		
Run 2		Same as in Run 1
Run 3	Same as in Run 1	
Run 4	Same as in Run 1	Same as in Run 1

For the concentration of the  $\text{H}_2\text{O}_2$  stock solution use the molarity that has been determined by the standardization procedure of this solution.  
(Data Table I, page 11)

EXPERIMENT 1  
THE RATE LAW

2. **Concentration calculation for  $[H_3O^+]$**

**(Reactant C = Reactant  $C_1$  + Reactant  $C_2$ )**

The source of  $H_3O^+$  ions is the weak Acetic Acid ( $HC_2H_3O_2$ ) whose partial ionization yields a small percentage of  $H_3O^+$  that depends on the concentration of the weak acid. In order to calculate the concentration of the  $H_3O^+$  ions, we first must calculate the concentration of the Acetic Acid ( $HC_2H_3O_2$ ) in the Run 1, 2 & 3 (identical) and Run 4 (different), by taking into account the dilution of the acetic acid to 150.0 mL

• **Concentration calculation for Acetic Acid solution for Runs 1, 2 & 3**

The only source of the Acetic Acid ( $HC_2H_3O_2$ ) in Runs 1, 2 & 3 is the **buffer solution with an Acetic Acid ( $HC_2H_3O_2$ ) concentration of 0.050 M.**

Calculate the concentration of the Acetic Acid ( $HC_2H_3O_2$ ) in these runs:

$$[HC_2H_3O_2]_0 =$$

➤  **$[H_3O^+]_0$  Concentration calculation for Runs 1, 2 & 3**

$$(C = C_1 + C_2 = 0 + C_2 = C_2)$$

The Percent Ionization of the Acetic Acid ( $HC_2H_3O_2$ ) in this buffer solution is 0.18 %. Calculate the  $[H_3O^+]$  by using the following formula:

$$[H_3O^+] = \text{Concentration of } [HC_2H_3O_2] \times \frac{\text{Percent Ionization}}{100}$$

$$[H_3O^+]_0 =$$

• **Concentration calculations for Acetic Acid solution for Run 4**

There are two sources of Acetic Acid ( $HC_2H_3O_2$ ) in the reaction mixture for Run 4. In order to calculate the total Concentration of  $HC_2H_3O_2$ , we need to consider both these sources and account for the respective dilution factors.

➤ Calculate the number of moles of Acetic Acid provided by the buffer solution.

$$? \text{ moles } HC_2H_3O_2 =$$

➤ Calculate the number of moles of Acetic Acid provided by the Acetic Acid solution.

$$? \text{ moles } HC_2H_3O_2 =$$

---

EXPERIMENT 1  
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- What is the total number of moles in this reaction mixture?

# moles =

- What is the molarity of the  $\text{HC}_2\text{O}_3\text{O}_2$  in this reaction mixture?

**? moles**  

---

**L solution** =

- **$[\text{H}_3\text{O}^+]_0$  Concentration calculation for Run 4**

**$(C = C_1 + C_2)$**

The Percent Ionization of the  $[\text{H}_3\text{O}^+]$  in this reaction mixture is 0.18 %  
Calculate the  $[\text{H}_3\text{O}^+]$  in this reaction mixture:

**$[\text{H}_3\text{O}]_0 =$**

EXPERIMENT 1  
THE RATE LAW

**PART III: DETERMINATION OF REACTION RATES**

**DATA TABLE III**  
Concentrations of reagents involved in the determination of Reaction Rates

	$[\text{S}_2\text{O}_3^{2-}]_{\text{added}}$ (show calculations)	$[\text{I}_2]_{\text{produced}} = \frac{[\text{S}_2\text{O}_3^{2-}]}{2}$ (show calculations)
All Runs		

**DATA TABLE IV**  
Reaction Rate Measurements and Calculation

Run Nr.	[H <sub>2</sub> O <sub>2</sub> ] <sub>0</sub> Reactant A  (M)	[I <sup>-</sup> ] <sub>0</sub> Reactant B  (M)	[H <sub>3</sub> O <sup>+</sup> ] <sub>0</sub> Reactant C  (M)	TIME ( <u>t<sub>color</sub></u> )  (s)	[I <sub>2</sub> ] produced  M	REACTION RATE  $\frac{[\text{I}_2]_{\text{produced}}}{t_{\text{color}}}$  (M x s <sup>-1</sup> )
1						
2						
3						
4						

Express Reaction Rates as (**A x 10<sup>-6</sup>**) throughout the entire experiment



---

EXPERIMENT 1  
THE RATE LAW

**PART IV:****DETERMINATION OF REACTION ORDERS WITH RESPECT TO REACTANTS**

The general formula for the Rate Law for the reaction studied is:



When performing your calculations you are required to:

- Show all calculations neatly, in a well-organized manner and in detail.
- Round off your answer to an integer.

**1. Reaction Order with respect to [H<sub>2</sub>O<sub>2</sub>]**

From Run 1 and Run 2

**Rate 1**

\_\_\_\_\_ = \_\_\_\_\_ = \_\_\_\_\_

**Rate 2**

Value of “m” rounded off to an integer =

**2. Reaction Order with respect to [I<sup>-</sup>]**

From Run 1 and Run 3

**Rate 3**

\_\_\_\_\_ = \_\_\_\_\_ = \_\_\_\_\_

**Rate 1**

Value of “n” rounded off to an integer =

**3. Reaction Order with respect to [H<sub>3</sub>O<sup>+</sup>]**

From Run 1 and Run 4

**Rate 1**

\_\_\_\_\_ = \_\_\_\_\_ = \_\_\_\_\_

**Rate 4**

Value of “p” rounded off to an integer =

EXPERIMENT 1  
THE RATE LAW

**PART V:****DETERMINATION OF THE “RATE CONSTANT, “k” AT ROOM TEMPERATURE**

The Rate Constant, “k” can be calculated by substituting the known values of reactant concentrations, the reaction orders and the corresponding reaction rates in the formula of the Rate Law. Refer to the example on page 6.

Please express all quantities in the appropriate number of significant figures.

**DATA TABLE V**  
**Summary of Reaction Orders**

Reactant	H <sub>2</sub> O <sub>2</sub>	I <sup>-</sup>	H <sub>3</sub> O <sup>+</sup>
Reaction Order	m = ____	n = ____	p = ____

**DATA TABLE VI**  
**Rate Constant Calculations for “k” at Room Temperature**

	[H <sub>2</sub> O <sub>2</sub> ] <sup>m</sup> Reactant A	[I] <sup>n</sup> Reactant B	[H <sub>3</sub> O <sup>+</sup> ] <sup>p</sup> Reactant C	Reaction Rate (M · s <sup>-1</sup> )	Rate Constant “k” (M <sup>-1</sup> · s <sup>-1</sup> )
Run 1					
Run 2					
Run 3					
Run 4					
<b>AVERAGE</b>					

**CONCLUSION:**

**Write the complete form of the Rate Law:**

- Include the **formulas** of all Reactants. Do not include numerical values.
- Include the respective Reaction Orders for all Reactants.
- Include the experimentally determined numerical value of “k”, expressed in the correct units.

**RATE =**

**Bibliography:**

1. Nivaldo J. Tro, “Chemistry: A Molecular Approach”, Third Edition
2. R.A.D. Wentworth “Experiments in General Chemistry”, Sixth Edition
3. James M. Postma & all, “Chemistry in the Laboratory”, Seventh Edition