

## Lab 1. Kinetics of the Iodine Clock Reaction

### Prelab Assignment

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*Before coming to lab...*

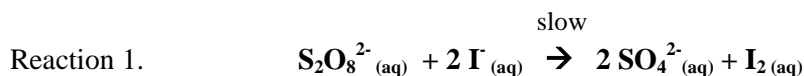
1. Read the lab thoroughly.
2. Follow the guidelines in the "Lab Notebook Policy and Format for Lab Reports" handout and the "Lab Report Checklist" on page 9 to complete in your lab notebook the following sections of the report for this lab exercise: Title, Introduction, Materials/Methods and Data Tables. **Use the data tables provided for Parts 1 – 3.**
3. For the **materials and methods section**, simply cut and paste into your notebook the procedures for parts 1 – 3 on pages 4 – 6, but trim the handout pages so they do not cover the entire notebook page. You will **need two copies** of the procedure, one for the original pages and one for the carbon copy pages. **You will need to devise and record your own procedure and data table for part 4!**
4. Ensure that the table of contents of your lab notebook is current.
5. Answer the pre-lab questions in the spaces provided at the end of this lab exercise—i.e. the last three pages of this lab. Background information can be found in **Chapter 16, especially sections 16.1-16.3** in your textbook (*Silberberg 6<sup>th</sup> ed*).

### Introduction

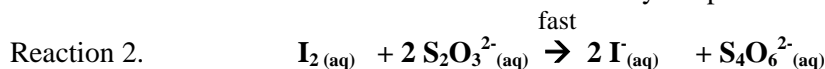
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The major goal of this investigation is to determine the factors, both qualitatively as well as quantitatively, that affect the rate of the iodine clock reaction described below. More specifically, by systematically varying the concentration of the reactants and measuring the initial reaction rate you will determine the rate law and the rate constant of reaction 1, below. Next, by carrying out the reaction at several different temperatures you will determine the activation energy of reaction 1. Lastly, you will investigate the effect of metal ion catalysts on the rate and activation energy of reaction 1.

The iodine clock reaction is a well-known and memorable chemical reaction where colorless solutions are mixed and, after a period of time ranging from seconds to minutes, the solution suddenly turns bluish-black. The reaction you will study is the oxidation-reduction reaction between the **peroxydisulfate ion** and the **iodide ion**, a reaction that is relatively slow at room temperature:



A starch indicator, which turns blue-black in the presence of iodine, has been added to the reaction mixture. However, the color does not show right away because a second reaction (reaction 2) takes place, in which **thiosulfate ions** consume the **iodine molecules** as soon as they are produced in reaction 1.



Since the first reaction is slow, and the second is fast, the iodine molecules are not present long enough to react with the starch and produce the color. Hence, it is only after all of the  $\text{S}_2\text{O}_3^{2-}$  has been used up that iodine accumulates, then combines with the starch, and the solution turns blue-black.

Although three reactions are involved, the reaction between peroxydisulfate and iodide ions (reaction 1) is the one of interest. The second reaction (rxn 2) is only used to delay the reaction between iodine and starch (rxn 3)—in other words, all of the thiosulfate must be consumed before the color changes. Without the thiosulfate ions present, the iodine molecules produced from the 1<sup>st</sup> reaction would slowly build up and the solution would gradually darken, making it difficult to identify a definitive point at which the reaction involving the thiosulfate (rxn 2) is over.

## What is the Rate Law for Reaction 1?

The dependence of the rate of reaction 1 on concentration is given by the rate law:

$$\text{Rate} = k[\text{S}_2\text{O}_8^{2-}]^m[\text{I}^-]^n$$

In this expression **rate** is in units of M/s, **k** is the **rate constant**, and **m** and **n** are (usually) integers that represent the reaction order for each reactant. The units for **k** vary according to the overall order of the reaction. It is important to remember that it is not possible to predict the values of **m** and **n** from the coefficients in the balanced equation—the value for the reaction order for each reactant, **m** and **n**, must be determined experimentally.

The rate of a reaction is the change in concentration of a reactant with time. In this experiment we will monitor the peroxydisulfate concentration, so the rate will be expressed as:

$$\text{Rate} = -\Delta[\text{S}_2\text{O}_8^{2-}]/\Delta t$$

The minus sign indicates that the peroxydisulfate ion,  $\text{S}_2\text{O}_8^{2-}$ , is being used up as the reaction proceeds.

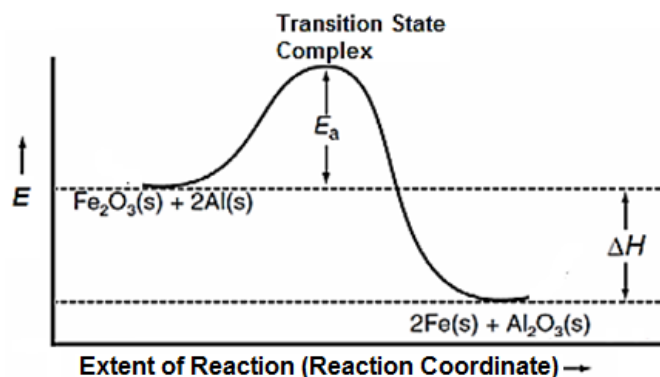
## How does the Rate Constant vary with Temperature and what is the $E_a$ of Reaction 1?

The **rate constant**, **k**, is a constant for a given reaction and varies only with temperature or in the presence of a catalyst. The **Arrhenius equation** describes how **k** varies with temperature (Does the rate constant increase or decrease as temperature increases?):

$$k = A e^{-\frac{E_a}{RT}}$$

$k$  = rate constant (units depend on rate law);      $E_a$  = Activation energy (J/mol)  
 $R$  = gas law constant = 8.314 J/mol K;      $T$  = temperature in Kelvin;  
 $A$  = frequency factor—it relates to the frequency and orientation of the  
the reactant particles as they collide

Another goal of this lab is to determine the **activation energy** for reaction 1. There is an “energy barrier” that reactants must surmount for a reaction to take place. At the top of this energy barrier there is a **transition state complex** that forms before the products are formed. This energy barrier is called the **activation energy**,  $E_a$ , and is the difference in energy between the reactants and the transition state complex. This energy can range from almost zero to hundreds of kJ/mol. Reactants need to possess this amount of energy both to overcome the repulsive electron cloud forces between approaching reactant particles (ions, molecules or atoms) and to break existing bonds within the reactants. Only reactant particles with kinetic energy greater than or equal to the activation energy can react. In general, the higher the activation energy, the slower the reaction.



**Figure 1.** An example of an activation energy reaction coordinate diagram for the thermite reaction, a highly exothermic reaction between iron (III) oxide and aluminum.

Examine the Arrhenius equation, does the rate constant increase or decrease as the activation energy increases? How will this affect the reaction rate? Why?

Measurement of the rate at several temperatures (e.g. room temperature and  $10^\circ\text{C}$  above and below room temperature) will allow you to calculate a value for the **activation energy**,  $E_a$ , for reaction 1.

Taking the natural logarithm of each side of the Arrhenius equation yields an equation of a line:

$$\ln k = \ln A - \frac{E_a}{R} \left( \frac{1}{T} \right)$$

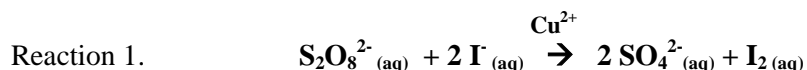
$$y = b + m(x)$$

Hence, a *plot of  $\ln k$  vs.  $1/T$*  should yield a straight line with a negative slope. The slope can then be used to calculate the activation energy of reaction 1. What do you need to multiply the slope by to find the value of the activation energy?

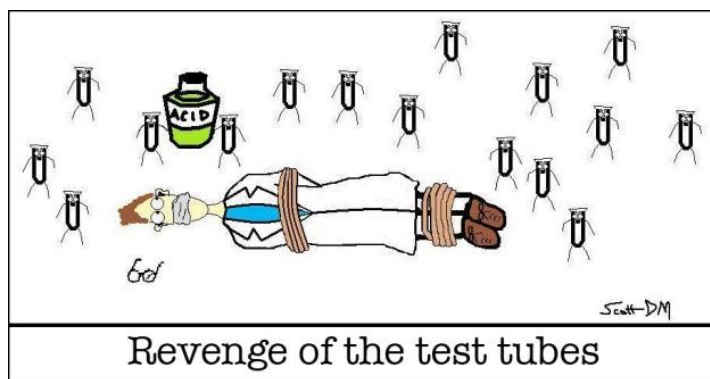
### How does a Catalyst affect the Rate and the $E_a$ of Reaction 1?

The final goal of this lab is to determine the effect that a catalyst will have on the rate of reaction for reaction 1. A *catalyst* is a substance that increases the rate of reaction (both in the forward and reverse directions) without being consumed in the reaction. Catalysts often allow the reactants to react using a reaction pathway (or mechanism) that is lower in activation energy than without the presence of the catalyst. Three metal ion catalysts will be investigated in this lab,  $\text{Cu}^{2+}$  ion from copper (II) sulfate,  $\text{Mg}^{2+}$  ion from magnesium sulfate and  $\text{Cr}^{3+}$  ion from Chromium (III) sulfate.

For example, with the  $\text{Cu}^{2+}$  ion as the catalyst, reaction 1 can be written:



Go to page 4 for the procedure....



## Procedure

### Materials

- Obtain the following items from the lab cart or your drawer:
  - Stopwatch
  - Six 250 mL beakers
  - Six 50 or 100 mL graduated cylinders (use the ones in your drawer plus those on the cart!)
  - Several clean dry 250 mL Erlenmeyer flasks.
- Clean and dry six beakers from your lab drawers for the following amounts of solutions; **label** them so that you don't mix them up! Record the identities and actual concentrations of all solutions in your lab notebook. You will need approximately (*please don't take more than needed*):
  - 250 mL of 0.200 M NaI
  - 200 mL of 0.200 M NaCl
  - 200 mL of 0.00500 M Na<sub>2</sub>S<sub>2</sub>O<sub>3</sub>
  - 100 mL of 1% starch solution
  - 200 mL of 0.100 M K<sub>2</sub>SO<sub>4</sub>
  - 250 mL of 0.100 M K<sub>2</sub>S<sub>2</sub>O<sub>8</sub>
- Each of the graduated cylinders must be used for only one solution. **Label** the six graduated cylinders with the names of the six solutions. Make sure that the cylinders are clean and dry.

### Part 1. How does the Concentration of Reactants affect the Rate of Reaction 1? (Work in teams of 4)

- Record** the temperature of the NaI and the K<sub>2</sub>S<sub>2</sub>O<sub>8</sub> in the beaker. They should be similar.
- Trial 1:
  - Using the exact volumes listed for **trial 1** in **table 1** (measured to the nearest 0.1 mL), add to a 250 mL Erlenmeyer flask all of the solutions except the K<sub>2</sub>S<sub>2</sub>O<sub>8</sub> (in the order listed above) and swirl gently to mix.
  - While one partner monitors the stopwatch, another adds the K<sub>2</sub>S<sub>2</sub>O<sub>8</sub> to the flask all at once. Begin timing at the instant the K<sub>2</sub>S<sub>2</sub>O<sub>8</sub> solution is added. Swirl to get the K<sub>2</sub>S<sub>2</sub>O<sub>8</sub> solution thoroughly mixed. Stop timing the instant the solution turns blue. Do not look away from the flask—if you do you may miss it! It may take up to five minutes for the color to change!
- Before doing trials 2 – 5, below, **go to step 11 of part 2** and do the trials your group is assigned (either trials 6 and 9 (even-numbered groups) or trials 7 and 8 (odd-numbered groups)).
- Trials 2 and 3: Repeat steps 4 and 5, above, using the solution volumes for the 2<sup>nd</sup> and 3<sup>rd</sup> trials as shown in **table 1** below.

**Table 1.** Solution volumes in mL to use in each trial to determine the effect of reactant concentration

Trial	0.200 M NaI	0.200 M NaCl	0.00500 M Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub>	1% Starch	0.100 M K <sub>2</sub> SO <sub>4</sub>	0.100 M K <sub>2</sub> S <sub>2</sub> O <sub>8</sub>	Total Volume	temp (°C)	Time (s)	
									Predicted	Actual
1	20.0	20.0	20.0	10.0	20.0	20.0	110.0			
2	20.0	20.0	20.0	10.0	0.0	40.0	110.0			
3	40.0	0.0	20.0	10.0	20.0	20.0	110.0			
4	?	?	20.0	10.0	?	?	110.0			
5	20.0	20.0	20.0 mL of 0.00250 M	10.0	20.0	20.0	110.0			

8. **Trial 4:** Using volumes of your own choosing such that *both* the NaI and  $K_2S_2O_8$  have different volumes compared to any previous trial. Before carrying out the trial, predict the reaction time. **Record your prediction and the logic behind your reasoning in your lab notebook and have it approved by your instructor, BEFORE carrying out the trial.** Remember that only the NaI and  $K_2S_2O_8$  volumes affect the rate. The other solutions are there either to provide charge balance or as an indicator. When deciding on new solution volumes, use the following criteria:
- Keep the volume of starch at 10.0 mL and the volume of  $Na_2S_2O_3$  at 20.0 mL
  - The volume of NaCl plus NaI should equal 40.0 mL
  - The volume of  $K_2SO_4$  plus  $K_2S_2O_8$  should equal 40.0 mL
  - The purpose of the NaCl and  $K_2SO_4$  solutions is to keep the ionic strength the same for all trials. Ionic strength is a measure of the concentration of ions in a solution.
9. **Trial 5:** Dilute 10.0 mL of the  $Na_2S_2O_3$  to 20.0 mL with DI water, and use this in the fifth trial, leaving all other solutions the same as in trial #1. What do you think will happen to the rate? Why? **Record your prediction and reasoning in your lab notebook, and have it approved by your instructor, before carrying out the trial.**
10. Go to part 3.

### Part 2. How does Temperature affect the Rate of Reaction 1? (Work in teams of 4)

11. **Trials 6 – 9:** Next, each team will repeat trial #1 at *two different assigned temperatures*. Shared class data plus your original trial #1 at room temperature will provide quantitative data for the dependence of the rate on temperature. Use a suitable combination of ice, cold water and warm water to prepare the various temperature baths.

#### **Caution!! Caution!! Caution!! Caution!! Caution!! Caution!!**

All solutions (the mixture in the reaction flask and the  $K_2S_2O_8$  in a graduated cylinder) must equilibrate in the bath until their temperatures are within 1°C of each other before beginning the experiment. Use ring weights to keep the flasks in place in the bath. After the  $K_2S_2O_8$  is added, the reaction flask must be kept in the bath for the entire time until it turns blue, and the temperature of the bath should be kept constant to within 1°C by adding more ice or hot water as necessary.

#### **\*Temperatures to be used:**

- a) Trial 1: Room temperature (enter the data acquired in part 1)
- b) **Even numbered teams do...**
  - Trial 6: 10.0 °C *below* room temperature
  - Trial 9: 20.0 °C *above* room temperature
- c) **Odd numbered teams do...**
  - Trial 7: 20.0 °C *below* room temperature
  - Trial 8: 10.0 °C *above* room temperature

**Table 2.** Solution volumes *in mL* to use in each trial to determine the effect of temperature

Trial	Temp* (°C)	Time (s)		0.200 M NaI	0.200 M NaCl	0.00500 M $Na_2S_2O_3$	1% Starch	0.100 M $K_2SO_4$	0.100 M $K_2S_2O_8$	Total Volume
		Predicted	Actual							
1				20.0	20.0	20.0	10.0	20.0	20.0	110.0
6										
7										
8										
9										

12. **Trials 2 – 5:** *Go back to part 1 and do steps 7 - 9* to complete trials 2 – 5.



**Part 3. Which Metal Ion Catalyst increases the Rate of Reaction 1 the most? (Do in teams of two)**

13. Do in teams of two. In this part of the experiment will study the effect of various *metal ion catalysts* on the rate of the iodine clock reaction.

Obtain the your team's assigned solution (in dropper bottles) from the lab cart

→ 0.10 M copper (II) sulfate (*Used by teams 1, 4, 7, 10 for trials 10 – 11*)

→ 0.10 M magnesium sulfate (*Used by teams 2, 5, 8, 11 for trials 12 – 13*)

→ 0.10 M chromium (III) sulfate (*Used by teams 3, 6, 9, 12 for trials 14 – 15*)

14. Set up a reaction mixture as described for trial 1 in Part 1, but before mixing add **5 drops** of your assigned metal ion catalyst to the 20.0 mL of  $\text{Na}_2\text{S}_2\text{O}_8$ . Mix the solutions well and measure the time for the reaction to turn blue-black.

15. Repeat step 14, above, but use 10 drops of your team's assigned metal ion catalyst.

**Table 3.** Solution volumes in mL to use in each trial to determine the effect of a catalyst

Trial	Metal Ion Catalyst	Drops of Catalyst	Time (s)		0.200 M NaI	0.200 M NaCl	0.00500 M $\text{Na}_2\text{S}_2\text{O}_3$	1% Starch	0.100 M $\text{K}_2\text{SO}_4$	0.100 M $\text{K}_2\text{S}_2\text{O}_8$	Total Volume
			Predicted	Actual							
10	0.10 M $\text{Cu}^{2+}$	5			20.0	20.0	20.0	10.0	20.0	20.0	110.0
11	0.10 M $\text{Cu}^{2+}$	10									
12	0.10 M $\text{Mg}^{2+}$	5									
13	0.10 M $\text{Mg}^{2+}$	10									
14	0.10 M $\text{Cr}^{3+}$	5									
15	0.10 M $\text{Cr}^{3+}$	10									

**Part 4. How does the Presence of a Metal Ion Catalyst affect the Activation Energy of Reaction 1?**

16. The goal now is to design a procedure similar to that used in part 2 that will allow you to determine the activation energy of reaction 1 in the presence of the most effective catalyst from part 3.

Use the following criteria when designing your experiment

- Use the best catalyst from part 3—i.e. the one with the fastest time!
- How many drops of catalyst should you use for each trial? You don't want the reaction to be too fast, especially at higher temperatures!!
- Use the reagent volumes used in trial 1.
- Neatly record your procedure in enough detail so that a competent student could reproduce your results

**Clean up...**

17. Clean and dry all glassware and return those items to the lab cart which you obtained there.

## Calculations

### Part 1. How does the Concentration of Reactants affect the Rate of Reaction 1?

- Excel spreadsheet:** Set up an *Excel spreadsheet* to create a results table which summarizes the following for all trials: Trial#, volumes of NaI, Na<sub>2</sub>S<sub>2</sub>O<sub>3</sub>, and K<sub>2</sub>S<sub>2</sub>O<sub>8</sub>, temperature, initial concentrations of I<sup>-</sup> and S<sub>2</sub>O<sub>8</sub><sup>2-</sup>, time in sec, rate of reaction, and rate constant as calculated in the following steps. Let Excel do the following calculations for you, but you should individually show at least one sample calculation of each type by hand in your notebook. Clearly label all sample calculations in your notebook.

Below is a portion of a sample table that you should use *Excel* to create and do all calculations...

Trial	Volume (mL)							temp (°C)	Time (s)		Initial Concentration (M)		Rate (units?)	Rate Constant, k (units?)
	0.200M NaI	0.200M NaCl	0.00500M Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub>	1% Starch	0.100 M K <sub>2</sub> SO <sub>4</sub>	0.100 M K <sub>2</sub> S <sub>2</sub> O <sub>8</sub>	Total Volume		Predicted	Actual	I <sup>-</sup>	S <sub>2</sub> O <sub>8</sub> <sup>2-</sup>		
	1	20.0	20.0	20.0	10.0	20.0	20.0		110.0					

- Calculate the initial concentration of each reactant, I<sup>-</sup> and S<sub>2</sub>O<sub>8</sub><sup>2-</sup>.** The initial concentration is understood to be the concentration of each reactant in the instant after the solutions are mixed, but before any reaction takes place. Because the concentration of a solution is determined in part by the total solution volume, when two solutions are mixed, the concentration of each solution decreases because the total solution volume increases. Remember to take into account this dilution factor—i.e. use the dilution equation,  $V_{conc} M_{conc} = V_{dil} M_{dil}$
- Calculation of the reaction rate for each trial: Rate =  $-\Delta[S_2O_8^{2-}]/\Delta t$** 
  - Calculate the *initial* concentration of S<sub>2</sub>O<sub>3</sub><sup>2-</sup>.
  - Recall that all of the S<sub>2</sub>O<sub>3</sub><sup>2-</sup> reacts when the solution turns blue, so what is the final concentration of S<sub>2</sub>O<sub>3</sub><sup>2-</sup>? Find the **change in concentration of S<sub>2</sub>O<sub>3</sub><sup>2-</sup>**,  $\Delta[S_2O_3^{2-}]$ , ( $\Delta = \text{final} - \text{initial}$ ) including an appropriate sign.
  - How many moles of S<sub>2</sub>O<sub>8</sub><sup>2-</sup> react for every mole of S<sub>2</sub>O<sub>3</sub><sup>2-</sup> that reacts? Consider the *balanced* chemical equations for reactions 1 and 2 to calculate the **change in concentration of S<sub>2</sub>O<sub>8</sub><sup>2-</sup>**,  $\Delta[S_2O_8^{2-}]$ , in this trial.

In order to determine the reaction rate, ***you will need to determine the number of moles per liter of S<sub>2</sub>O<sub>8</sub><sup>2-</sup> that react in the measured time period.*** Note that when the solution turns blue–black, only a very small percent of the total S<sub>2</sub>O<sub>8</sub><sup>2-</sup> has reacted! Rather, the solution turns color when ***all of the thiosulfate S<sub>2</sub>O<sub>3</sub><sup>2-</sup> has reacted,*** allowing the accumulation of I<sub>2</sub> molecules in the solution. In order to determine  $\Delta[S_2O_8^{2-}]$ , consider that the change in moles of S<sub>2</sub>O<sub>3</sub><sup>2-</sup> = total moles of S<sub>2</sub>O<sub>3</sub><sup>2-</sup> present, and use stoichiometry to determine the corresponding change in moles of S<sub>2</sub>O<sub>8</sub><sup>2-</sup>.

- What **percent of the total S<sub>2</sub>O<sub>8</sub><sup>2-</sup> added actually reacts** in this time interval? Show this calculation in your notebook for Trial 1. You do not need to calculate this for each trial.
- Calculate the Reaction Rate.** Show as sample calculation for *trial 1* using *correct units and sig figs*. The reaction rate is defined as  $-\Delta[S_2O_8^{2-}]/\Delta t$ . It is important to remember that you need the change in concentration to calculate rate. Also note that ***the moles of S<sub>2</sub>O<sub>8</sub><sup>2-</sup> that react and hence  $\Delta[S_2O_8^{2-}]$  are the same for all trials where the initial concentration of S<sub>2</sub>O<sub>3</sub><sup>2-</sup> is the same***—can you explain why? The reaction rate varies because the *reaction time* is different for each trial—the reaction time varies because you are changing the concentration of the reactants [I<sup>-</sup>] or [S<sub>2</sub>O<sub>8</sub><sup>2-</sup>]. The reaction time is also influenced by changes in temperature or the addition of a catalyst.

4. a) **Determine the rate law for Reaction 1.** From the data above, determine the rate law for the reaction, *i.e.*, what are  $m$  and  $n$  in the rate law,  $\text{rate} = k[\text{S}_2\text{O}_8^{2-}]^m[\text{I}^-]^n$ ? Solve for  $m$  and  $n$  precisely using logarithms—See Section 16.3 and/or the model on page 3 of ALE 2 if you need help! Report both the decimal number obtained experimentally and the closest integer value. *Use the integer values in the next step.*
- b) From the above data, write the **rate law** for reaction 1, *i.e.*, what are  $m$  and  $n$  in the rate law,  $\text{rate} = k[\text{S}_2\text{O}_8^{2-}]^m[\text{I}^-]^n$ ?
5. **Calculate the rate constant for Reaction 1 at room temperature.** Assuming the orders are integers, calculate the rate constant,  $k$ , for each trial using your Excel spreadsheet!! As long as the temperature is constant,  $k$  should be the same for the first five trials, so use your Excel spreadsheet to find an average rate constant,  $k$ , at room temperature. Show a sample calculation for trial 1 and be sure to include the correct units and sig figs.

## Part 2. How does Temperature affect the Rate of Reaction 1?

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6. **Make an Arrhenius Plot and Calculate the Activation Energy for Reaction 1.**
- a) Using the trial #1 data at different temperatures (including room temperature), make an **Excel table** of  $k$ ,  $\ln k$ ,  $T$ , and  $1/T$ . Make Excel do these calculations.
- b) Use **Excel** to make an **Arrhenius plot of  $\ln k$  vs  $1/T$** , and include the correlation  $R^2$  and the equation of the linear trendline on your graph. Use the slope of the linear fit to calculate the activation energy  $E_a$  in kJ for the reaction and show this calculation in your lab notebook (individual). Be sure to include the correct units and use correct significant figures.
- c) Paste or tape in the results section of your notebook the Excel table and a properly formatted Arrhenius plot.

## Part 3. Which Metal Ion Catalyst increases the Rate of Reaction 1 the most?

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7. **Calculate the rate of reaction in the presence of each of the three catalysts** at room temperature.
- a) By what factor does each catalyst increase the rate of reaction #1 at room temperature?
- b) What effect does doubling the number of drops of catalyst have on the rate of reaction 1? Cite specific supporting numerical data.
- c) Explain why changing the amount of catalyst has the observed effect.

## Part 4. How does the Presence of a Metal Ion Catalyst affect the Activation Energy of Reaction 1?

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8. **Calculate the rate constant for Reaction 1 in the presence of a catalyst.** Assuming the orders, “ $m$ ” and “ $n$ ,” are integers, calculate the rate constant,  $k$ , for each temperature. Be sure to include the correct units.
9. **Make an Arrhenius Plot and Calculate the Activation Energy for Reaction 1 in the presence of the metal ion catalyst.** Using the data from trial #1 with the best metal ion catalyst at different temperatures (including room temperature), make an **Excel table** of  $k$ ,  $\ln k$ ,  $T$ , and  $1/T$ . Make Excel do these calculations. Use **Excel** to make an **Arrhenius plot of  $\ln k$  vs  $1/T$** , and include the correlation  $R^2$  and the equation of the linear trendline on your graph. Use the slope of the linear fit to calculate the activation energy  $E_a$  in kJ for the reaction and show this calculation in your lab notebook (individual). Be sure to include the correct units and use correct significant figures. Paste or tape in the results section of your notebook the Excel table and a properly formatted Arrhenius plot.
10. Compare the activation energy of reaction 1 with and without catalyst. Explain why they may or may not be different. Cite specific numerical data to support any and all conclusions.



## Analysis of the Results

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11. Start your analysis by summarizing and then discussing the most important results: the orders  $m$  and  $n$  (decimal and integer), the rate law with the average  $k$  value, and the activation energy, with and without catalyst. Discuss your results for the rate law and the Arrhenius plots, the precision (e.g.  $R^2$  for the Arrhenius plot and the standard deviation for the rate constant at room temperature). Discuss the effect the metal ion catalysts have on the rate and on  $E_a$ . Discuss the accuracy of your results—calculate the % error where possible (e.g. for  $m$  and  $n$  and  $E_a$  at room temperature), and as always discuss at least two sources of possible error and explain what effect each would have on the results. In all discussions, cite specific supporting numerical data/results.

## Conclusion

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12. Use bullets to briefly list the major conclusions and likely sources of error.

### WARNING!!!

You will be tested on your ability to use your results (e.g. to predict how long the iodine clock reaction will take a various reactant concentrations and temperatures), so make sure you have all the data and calculations you need recorded neatly in your notebook!

### Lab Report Checklist...

- |                                |  |
|--------------------------------|--|
| <b>Introduction</b>            | <ul style="list-style-type: none"> <li>• Includes <u>all</u> of the major goals of the experiment?</li> <li>• Discusses/explains each of the following?             <ul style="list-style-type: none"> <li>– Reactions 1 – 3</li> <li>– Explains why the solution turns blue and the cause for the delay?</li> <li>– The rate law for reaction 1</li> <li>– Rate = <math>-\Delta[S_2O_8^{2-}]/\Delta t</math></li> <li>– The Arrhenius equation and the natural log version</li> <li>– Activation energy</li> <li>– Plot of <math>\ln k</math> vs. <math>1/T</math> and meaning of slope</li> <li>– Catalysts</li> </ul> </li> </ul>   |
| <b>Materials &amp; Methods</b> | <ul style="list-style-type: none"> <li>• Cuts and pastes neatly onto the lab notebook pages the procedure pages 4 – 6?</li> </ul>  |
| <b>Results</b>                 | <ul style="list-style-type: none"> <li>• Data <u>neatly</u> recorded in a ruled and easy to read table or tables?</li> <li>• Separate data tables for parts 1 – 4?</li> <li>• Tables <u>numbered consecutively</u> and have <u>informative</u> captions?</li> <li>• Correct use of sig figs and units?</li> <li>• Excel table includes the following for all trials and uses correct units and sig figs...             <ul style="list-style-type: none"> <li>– volume &amp; molarity for all solutions</li> <li>– Temperature of the solutions?</li> <li>– Initial concentrations of I. and <math>S_2O_8^{2-}</math></li> <li>– Predicted and actual time for the color to change?</li> <li>– Rate and rate constant?</li> <li>– Excel table cut and pasted to lab notebook page?</li> <li>– Excel table is numbered and has an informative caption?</li> </ul> </li> </ul> |

**Lab Report Checklist continued on the next page!!**

**Lab Report Checklist (cont.)****Calculations**

Neat and orderly sample calculations and each calculation is neatly and clearly labeled?

- 1. Calculation of initial  $[I^-]$  and  $[S_2O_8^{2-}]$**  → consistent use of units and sig figs?
- 2. Calculation of Reaction Rate** → consistent use of units and sig figs?
  - Calculations are labeled with an informative caption
  - Initial  $[S_2O_3^{2-}]$ ,  $\Delta[S_2O_3^{2-}]$ ,
  - $\Delta[S_2O_8^{2-}]$
  - %  $S_2O_8^{2-}$  that reacts
  - Rate =  $\Delta[S_2O_8^{2-}] / \Delta t$
- 3. Determination of Rate law for rxn #1**
  - Clearly labels and work with correct units and sig figs?
  - Included decimal and integer values for m and n?
- 4. Sample calculation of k at room temperature?**
  - Correct use of units and sig figs for k
- 5. Two Excel Plots of  $\ln k$  vs.  $1/T$ : One w/o catalyst, the other with catalyst**
  - Axes labeled and correct use of sig figs? Graph has an informative caption?
  - Includes trendline equation and  $R^2$ ?
  - Calculation of  $E_a$  from slope of line? Uses units and sig figs?
  - Each graph comes with an Excel table containing k,  $\ln k$ , T and  $1/T$  for 5 different temperatures?
  - Compares  $E_a$  with and without catalyst and explains any differences?
- 6. Error Analysis:** discusses precision, accuracy and source(s) of error?
  - Calculates %Error for  $E_a$ ?

**Analysis**

- Summarizes and discusses the most important results: the orders m and n (decimal and integer), the rate law with the average k value, and the activation energy, with and without catalyst?
- Discusses the results for the rate law and the Arrhenius plot?
  - Discusses the *precision* (e.g.  $R^2$  for the Arrhenius plot and the standard deviation for the rate constant at room temperature)?
- Discusses the effect the metal ion catalysts have on the rate and on  $E_a$ ?
- Discusses the accuracy of the results?
  - Calculate the % error where possible? (e.g. for m and n and  $E_a$  at room temperature)
  - Discusses at least two sources of *possible* error and explains what effect each would have on the results
- Cites specific supporting numerical data/results where possible?

**Conclusion**

Uses “bullets” to state concisely the major conclusions and sources of error?

**General Considerations**

- The report is neat and easy to follow?      - Uses the format described in the “*Lab Notebook Policy*” handout?
- Grammar & spelling is at college level?      - Writing is clear, concise and to the point and not “wordy”?

**Acknowledgement:** This lab is adapted from similar labs produced by my present and past colleagues at GRCC, Sue Critchlow, Roger Knutsen and Mary Whitfield.

## Lab 1. Iodine Clock Reaction

Name \_\_\_\_\_

## Prelab Questions

Team Number \_\_\_\_\_ Date \_\_\_\_\_

**Instructions:** Complete the following 10 questions and hand in at the start of your lab period or when instructed by your instructor. For all questions that involve a calculation show your work with using *dimensional analysis/factor-label method* with *correct units* and *significant figures*.

1. What is the limiting reagent for the iodine clock reaction to produce a sudden color change?

*Circle your response(s):* I<sup>-</sup> S<sub>2</sub>O<sub>3</sub><sup>2-</sup> S<sub>2</sub>O<sub>8</sub><sup>2-</sup> Cu<sup>2+</sup> Cl<sup>-</sup> SO<sub>4</sub><sup>2-</sup> Mg<sup>2+</sup> Cr<sup>3+</sup> Starch I<sub>2</sub> S<sub>4</sub>O<sub>6</sub><sup>2-</sup>

Explain your reasoning.

2. Which trials in table 1 are used to determine “m” in the rate law for reaction 1? \_\_\_\_\_

3. Which trials in table 1 are used to determine “n” in the rate law for reaction 1? \_\_\_\_\_

4. Explain how and why an *increase* in temperature will affect the rate of a chemical reaction.

5. Suppose the rate law for a particular reaction between nitrogen dioxide and chlorine is  $\text{Rate} = k [\text{NO}_2]^2 [\text{Cl}_2]$ . How would each of the following affect the rate of the reaction?

a. If the concentration of NO<sub>2</sub> is *doubled* then the reaction rate would \_\_\_\_\_ by a factor of \_\_\_\_\_.

b. The concentration of NO<sub>2</sub> and Cl<sub>2</sub> are both *tripled*, then the reaction rate would \_\_\_\_\_ by a factor of \_\_\_\_\_.

c. The concentration of Cl<sub>2</sub> is halved, then the reaction rate would \_\_\_\_\_ by a factor of \_\_\_\_\_.

**Turn the page for more prelab questions** → → → → → → → → → → → → → → → → →

4. For the reaction  $4 A + 3 B \rightarrow 2 C$  the following data were obtained at constant temperature.

Trial	Initial [A] (mol/l)	Initial [B] (mol/l)	Initial Rate of Formation of C (mol/L · min)
1	0.100	0.100	5.00
2	0.300	0.100	45.0
3	0.100	0.200	10.0
4	0.300	0.200	90.0

- a) Use this data to determine the rate law for the reaction. *Briefly explain how you determined the rate law* or clearly show your work. Note: the rate law can be deduced without doing a detailed calculation. Circle your answer.

Rate Law: \_\_\_\_\_

Reasoning or calculation:

- b) Use the rate law in part (a), above, and the data from trial 3 to calculate,  $k$ , the rate constant—include units for the rate constant. Circle your Answer.

**Table 1.** *Fictitious* data for the iodine clock reaction. The solution volumes used in each are all in mL.

Trial	Time Needed for Color Change (sec)	Rate of Reaction, $-\Delta[S_2O_8^{2-}]/\Delta t$ (Mol L <sup>-1</sup> s <sup>-1</sup> )	Initial Conc. of NaI (M)	Initial Conc. of K <sub>2</sub> S <sub>2</sub> O <sub>8</sub> (M)	0.200 M NaI (mL)	0.0100 M Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub> + 0.1% Starch (mL)	D.I. H <sub>2</sub> O (mL)	0.200 M K <sub>2</sub> S <sub>2</sub> O <sub>8</sub> (mL)	Total Volume (mL)
1	21.2				40.0	20.0	20.0	20.0	100.0

Use the *fictitious data* in Table 1 to answer the questions that follow. Use correct significant figures and units for all calculations

5. Calculate the initial concentration of each reactant in reaction #1, I<sup>-</sup> and S<sub>2</sub>O<sub>8</sub><sup>2-</sup>. See #2 on page 7 for hints. **Enter your results in the table above and in the spaces below**

Answers: [I<sup>-</sup>] = \_\_\_\_\_

[S<sub>2</sub>O<sub>8</sub><sup>2-</sup>] = \_\_\_\_\_

Continued on the next page → → →

6. Calculate the initial concentration of  $\text{S}_2\text{O}_3^{2-}$ . **See #2** on page 7 for hints. Circle your answer and enter it in the space below.

$$[\text{S}_2\text{O}_3^{2-}] = \underline{\hspace{2cm}}$$

7. All of the  $\text{S}_2\text{O}_3^{2-}$  reacts when the solution turns blue, so what is the final concentration of  $\text{S}_2\text{O}_3^{2-}$ ? Find the *change in concentration* of  $\text{S}_2\text{O}_3^{2-}$  (as final – initial) including an appropriate sign. **See #3** on page 7 for hints. Circle your answer and enter it in the space below.

$$\Delta[\text{S}_2\text{O}_3^{2-}] = \underline{\hspace{2cm}}$$

8. How many moles of  $\text{S}_2\text{O}_8^{2-}$  react for every mole of  $\text{S}_2\text{O}_3^{2-}$  that reacts? Consider the *balanced* chemical equations for reactions 1 and 2 to calculate the change in concentration of  $\text{S}_2\text{O}_8^{2-}$ ,  $\Delta[\text{S}_2\text{O}_8^{2-}]$ , in this trial. **See #3** on page 7 for hints. Circle your answer and enter it in the space below.

$$\Delta[\text{S}_2\text{O}_8^{2-}] = \underline{\hspace{2cm}}$$

9. What percent of the total  $\text{S}_2\text{O}_8^{2-}$  added actually reacts in this time interval? **See #3** on page 7 for hints. Circle your answer and enter it in the space below.

$$\% \text{S}_2\text{O}_8^{2-} \text{ that reacts} = \underline{\hspace{2cm}}$$

10. Calculate the rate of reaction for this trial. **See #3** on page 7 for hints. Circle your answer and enter it in the space below.

$$\Delta[\text{S}_2\text{O}_8^{2-}]/\Delta T = \underline{\hspace{2cm}}$$