

The Iodine Clock: A Study in Kinetics

Cautions

These solutions may be toxic, irritants and corrosive. Avoid skin contact from all chemicals and wash thoroughly if contact occurs.

Purpose

The purpose of this experiment is to measure the effect of reactant concentration upon the rate of the reaction between the peroxydisulfate ion ($S_4O_6^{2-}$) and iodide ion (I^-), to experimentally determine the order of the reaction with respect to the reactant concentration, and to obtain the rate law for the chemical reaction.

Introduction

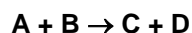
This experiment will investigate a reaction occurring at different rates as a result of varying the concentration of reactants. Two examples illustrating the difference in the rate of reaction include the rusting of iron as a reasonably slow reaction while the decomposition of TNT is an extremely fast reaction. Kinetics is the area of chemistry that deals with the studying of the rates of reactions. Four main factors that control the rates of homogeneous reactions in solution include the nature of the reactants, the concentration of the reactants, the temperature of reaction, and the use of catalysts. Reactions will only occur when reactant molecules obtain enough energy to interact or collide with other reactant molecules. Each one of the factors above affects how reactant molecules interact or collide with each other.

As the concentration of a solution changes, the number of particles per unit volume changes. This can be advantageous because an increase in the number of particles per unit volume will increase the probability of collisions between reacting species and therefore the rate of the reaction will increase.

In a simplistic model, temperature can be used as a measure of the average kinetic energy of a reaction. As the temperature of a reaction increases, the kinetic energy of reacting species increases. As a consequence of increased kinetic energy, the number of collisions between reacting species is increased and the rate of reaction is increased.

Catalysts are chemical species added to a reaction to increase the reaction rate. The mechanisms by which a catalyst increases the rate of reaction are numerous. An added catalyst can decrease the energy at which collisions occur, therefore allowing more collisions to occur and increasing the rate of reaction. A catalyst can also assist in the correct geometrical arrangement of reacting species necessary for collision thereby increasing the number of collisions and the rate of reaction. A major advantage of using a catalyst is that it remains chemically unchanged as a result of the reaction and can be recovered and/or reused. Catalysis is a branch of chemistry that is very vital in the production of many products used in our life such as plastics, petroleum, rubber and pharmaceutical products just to name a few.

Consider the general reaction scheme:



The rate of the reaction can be measured by observing the rate at which the reactants A and B disappear or the rate at which the products C and D appear. Experimentally, the change in concentration of A, B, C, and D with time can be measured.

The rate of reaction is expressed mathematically as follows:

$$\begin{aligned} \text{Rate of Disappearance of A} &= \frac{\text{Change in Concentration of A}}{\text{Time Required for Change}} = \frac{-\Delta[A]}{\Delta t} \\ \text{Rate of Appearance of C} &= \frac{\text{Change in Concentration of C}}{\text{Time Required for Change}} = \frac{+\Delta[C]}{\Delta t} \end{aligned}$$

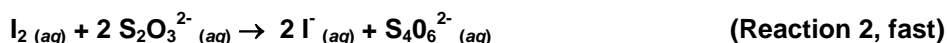
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The rates of the reaction will depend on the concentration of the reactants and is proportional to $1/\text{time}$. Therefore in this experiment, **$1/\text{time}$ values will be used as a substitute for rate of reaction.**

The rate law expression for the general reaction scheme is given as “rate = $k[A]^x[B]^y$ ”. In the expression [A] and [B] are the molar concentrations of the chemical species, A and B, x and y are the powers to which the respective concentrations must be raised to describe the rate, and k is the specific rate constant.

One of the objectives of chemical kinetics is to determine the rate law. The orders of the reaction are the powers to which the concentrations in the rate law are raised. Adding the individual exponents ($x+y$) together will give the overall order of the reaction. One method for determining the order of the reaction is to observe the changes of the reaction as reagent concentrations are varied. The stoichiometry and order of the reaction have no dependence on each other and usually are different. The specific rate constant, k , has a definite value for a given reaction and is independent of the concentration. The specific rate constant only depends on temperature and may be calculated once the rate of the reaction is known.

The reaction studied in this experiment is referred to as the “Iodine Clock Reaction” and involves the following two sequential chemical reactions:



Reaction 2 is instantaneous when compared to Reaction 1 and converts all the $\text{I}_{2(aq)}$ produced by Reaction 1, but only when some of $\text{S}_2\text{O}_3^{2-}_{(aq)}$ ion is present. When all the $\text{S}_2\text{O}_3^{2-}_{(aq)}$ ion has reacted, the level of the $\text{I}_{2(aq)}$ will increase as fast as Reaction 1 proceeds. $\text{I}_{2(aq)}$ is a yellow compound, however in low concentrations as in this experiment it appears almost colorless, like the ions in the two reactions.

In order to study the kinetics of Reaction 1, an indicator will be used to detect the $\text{I}_{2(aq)}$ molecule. Starch reacts with $\text{I}_{2(aq)}$ to produce a dark blue color. The reaction of starch with the $\text{I}_{2(aq)}$ molecule results in a convenient “clock” to measure the rate at which Reaction 1 proceeds. The rate of Reaction 1 depends on the concentrations of I^- , BrO_3^- and H^+ in solution. The concentration of reactants will be changed and the rate of the resulting reaction will be investigated in this experiment.

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Procedure

Part A: Preparation of Reaction Solutions

- A1) Obtain approximately 100 mL of KI, $\text{Na}_2\text{S}_2\text{O}_3$, KBrO_3 , and HCl in individual clean and dry beakers. Record exact concentrations from the reagent bottles in Table 1 of the data sheet.
- A2) Label the containers by either writing the chemical name on one end of a paper towel and setting the container on the other end or using a wax pencil. These reagents will be used to prepare the reaction mixture.
- A3) Set up a buret held by using a buret clamp on a ring stand and fill the buret with KI solution. The buret will be used to measure accurate volumes of the KI solutions.
- A4) Use a 10 mL graduated cylinder to measure the other reagent volumes given in table below.

Trial #	Flask 1			Flask 2		
	KI	$\text{Na}_2\text{S}_2\text{O}_3$	H_2O	KBrO_3	HCl	Starch
1	10.00 mL	10.0 mL	10.0 mL	10.0 mL	10.0 mL	3-4 Drops
2	20.00 mL	10.0 mL	0.0 mL	10.0 mL	10.0 mL	3-4 Drops
3	10.00 mL	10.0 mL	0.0 mL	20.0 mL	10.0 mL	3-4 Drops
4	10.00 mL	10.0 mL	0.0 mL	10.0 mL	20.0 mL	3-4 Drops

NOTE: To save time and reduce the possibility of contamination, use one 10 mL graduated cylinder for all the solutions in flask #1 and a second 10 mL cylinder for the solutions in flask #2. Be sure to rinse the containers thoroughly with distilled water between the different trials.

- A5) Using two 250 mL Erlenmeyer flasks, prepare the solutions according to the above table and record actual volumes on the data sheet in Table 1.
- A6) Obtain a digital timer.
- A7) Record room temperature.

Part B: Reaction and Observations

- B1) To begin the first reaction, pour the contents of flask #1 into flask #2 and begin timing the reaction as soon as the reagents are poured together.
- B2) Swirl the solution continually and gently to mix.
- B3) Record the time when the solution turns blue to the nearest second in Table 2 on the data sheet.
- B4) Repeat steps **B1** through **B3** for all 4 trials.

Disposal

Dispose of all solutions down the sink and flush with an excess of water.

Clean-Up

Wash all glassware with soap then rinse 3 times with tap water, and once with deionized water.

Clean your work area with water and dry with paper towels. Wash your hands before leaving the laboratory.

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Data Sheet & Calculations

Name: _____

Lab Partner: _____

*Show all work for calculations on page 7.***Part A:****Table 1: Stock Solution Concentrations**

Chemical Formula	Concentration (M)	Exact Concentration (M)
KI	0.010 M	
Na ₂ S ₂ O ₃	0.0010	
KBrO ₃	0.040	
HCl	0.10	

Table 2: Exact Solution Volumes (mL)

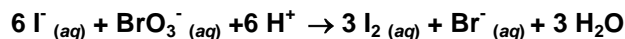
Trial #	KI	Na ₂ S ₂ O ₃	H ₂ O	KBrO ₃	HCl	Starch	Total Volume*
1							
2							
3							
4							

* Disregard the starch volume in the total volume.

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Part B:

Rate of Reaction Dependence on Concentration



$$\text{rate} = k[\text{I}^-]^m[\text{BrO}_3^-]^n[\text{H}^+]^p = -\frac{\Delta[\text{BrO}_3^-]}{\Delta t}$$

The color of the reaction changes when the available number of moles of BrO_3^- has been consumed in the reaction allowing for the time of the reaction to be measured. The rate of the reaction is determined by the time required for a color change to occur. The relative rate of each reaction will be inversely proportional to the time for a color change to occur. Remember from introduction: **1/time values will be used as a substitute for rate of reaction.**

Table 3: Rate of Reaction Dependence on Concentration (units in parenthesis)

Trial #	Temperature (°C)	Time (s)	1/ time (1/s)	[I ⁻] (M) *	[BrO ₃ ⁻] (M) *	[H ⁺] (M) *
1						
2						
3						
4						

* Reactant concentrations in the reaction mixture will not be the same as the stock solution concentrations in Table 1. Use dilution calculations ($M_1V_1 = M_2V_2$) to calculate each reactant concentration in Table 2. The KI stock solution is the source of I^- , and therefore the concentration of KI is used to calculate the concentration of I^- , [I⁻]. Likewise, KBrO_3 is the source of BrO_3^- , and HCl is the source of H^+ .

Determining the Orders of the Reaction

- Use the experimental data collected to calculate the exponents of the rate law expression, and report each exponent to the nearest whole number value. Attach to the data sheets the necessary calculations for determining the exponents of "m", "n", and "p" from the data collected.

Rate Law: $\text{rate} = k[\text{I}^-]^m[\text{BrO}_3^-]^n[\text{H}^+]^p$

m = _____

n = _____

p = _____

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- Calculate the pseudo-rate constant, pseudo-k, for each trial by substituting the concentrations of I^- , BrO_3^- , and H^+ and the values of m, n, and p into the rate law expression. (Since $1/\text{time}$ is being used as the rate of the reaction, the “pseudo-rate” constant is being calculated instead of the “real” rate constant.) Also, calculate the average pseudo-rate constant.

Table 4: Pseudo-Rate Constant Values

Trial	1	2	3	4	Average	Unit
pseudo-k						

Calculations:

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Post-lab Assignment

Name: _____

1. Write out the complete rate law expression for the iodine clock reaction.
2. What is the overall order of the iodine clock reaction?
3. If you wanted to double the rate of the iodine clock reaction, what experimental changes must be made?
4. Should all the rate constants for the different trials be similar? Why or Why not? Do your values confirm your prediction? Why or Why not?

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Pre-lab Assignment

Name: _____

1. What is a catalyst?
2. How will an increase in temperature affect the rate of reaction? Explain your reasoning using examples.
3. Name two variables other than temperature that effect reaction rate. Describe how each influences reaction rate.
4. From the data below calculate the rate law expression for reaction of $A + B \rightarrow C + D$. Also, give the overall order of the reaction and the rate constant. {As with today's experiment $1/\text{time}$ will be used as a substitute for rate}

Experiment #	[A] (M)	[B] (M)	Time (s)	Rate (1/s)
1	0.10	0.10	60	
2	0.20	0.10	15	
3	0.10	0.20	30	

Rate Law Expression: _____

Overall Order of Reaction = _____

k = _____