Determining the Rate of a Reaction Lab
Advanced Chemistry
NIVA International School 2010-2011

Thoroughly read the entire procedure before starting the lab.

Note: This lab will be submitted as a formal lab report. All observations must be properly recorded, in pen, on your data sheet (use the back if needed) and submitted with your formal lab report.

INTRODUCTION

The *initial-rate method* may be used to determine the reaction *order* of each component in a reacting system. In this method, the rate of reaction is measured near the beginning of the reaction, before the concentrations of components in the system have changed significantly. Then a new system is prepared in which the concentration of only one component is changed from that of the original reaction, and the rate is again measured. This procedure is repeated for each reactant.

In this experiment, the reaction of iodide ion with persulfate ion, represented by the following chemical equation, is studied:

$$2\ I^-(aq) + S_2O_8^{2-}(aq) \rightarrow I_2(aq) + 2S0_4^{2-}(aq) \qquad (1)$$

However, it does not follow that the *order* with respect to each component is the same as the stoichiometry. Rather, the rate equation should be written as

$$\text{Rate} = k\,[I^-]^a\,[S_2O_8^{2-}]^b \qquad (2)$$

where the rate constant, *k,* and the reaction order for each component, a and b, must be determined experimentally.

To determine the reaction rate, you must measure the amount of product formed, or reactant removed, in a time interval. In the experiment you will measure the time required for a specific amount of diiodine to be produced.

Diiodine reacts very rapidly with thiosulfate ion, represented by the equation:

$$I_2\ (aq) + 2S_2O_3^{2-}(aq) \rightarrow 2\ I^-(aq) + S_4O_6^{2-}(aq) \qquad (3)$$

As long as any thiosulfate ion is present in the solution, diiodine will react as quickly as it is produced by reaction (1). Not until after the thiosulfate ion has been totally used up will diiodine begin to appear in the solution. This first appearance of the free diiodine is detected by its reaction with starch (already present in the solution), forming a characteristic blue-black starch-iodine complex.

The average rate for the reaction is given by:

$$AverageRate = \tfrac{1}{2}\,\frac{\text{(quantity of } S_2O_3^{2-} \text{ provided) to consume the } I_2}{\text{time required for blue colour to appear}}$$

So that the rate, and rate constant, will be independent of the size of the reacting system, the rate is usually defined as the rate of change of the concentration. The relationship becomes:

$$\text{Average rate} = 0.5\,[S_2O_3^{2-}] / t \qquad (4)$$

To determine the order of reaction with respect to iodide ion, *i.e.* the value of 'a' in equation (2), the concentration of persulfate ion is kept constant and the concentration of iodide ion varied. If 'a' is first order, halving the concentration of iodide ion would double the time required for reaction. If the reaction is second order, then halving the concentration would quadruple the time required for reaction, etc. The value of 'b' may be determined by varying the persulfate ion concentration while keeping the iodide ion concentration constant.

Once the orders have been determined the value of the rate constant, *k,* may be calculated, knowing the concentration of components involved. From equation (2)

$$k = \text{Rate} / [I^-]^a\,[S_2O_8^{2-}]^b \qquad (5)$$

This rate constant should not vary as long as the temperature remains fixed.

However, the presence of a substance which provides for an alternative reaction mechanism may influence the rate of reaction. Such substances are called catalysts if they speed up the reaction but are not themselves any different after the reaction has occurred. The substance you will examine for its potential as a catalyst is copper(II) sulfate.

By examining the reaction at different temperatures you can determine the activation energy for this reaction. The most reliable way of doing this is graphically. You will plot ln k against 1/T, which should produce a straight line with a slope of $-E_a/R$.

$$\frac{\ln k}{1/T} = -E_a/R$$

where k is the rate constant, E_a is the activation energy, R is the gas constant. (8.31 J K^{-1} mol^{-1}) and T is the absolute temperature in Kelvin (K).

The ions in this lab react according to the following equations:

$$2I^- + S_2O_8^{2-} \rightarrow I_2 + 2SO_4^{2-} \quad \text{...slow...(1)}$$

$$\underline{I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-} \quad \text{...fast...(2)}}$$

$$S_2O_8^{2-} + 2S_2O_3^{2-} \rightarrow 2SO_4^{2-} + S_4O_6^{2-} \quad \text{...overall...}$$

The main focus is on the slow step, equation (1) – the rate-determining step. You will measure the different times required to make the same amount of I_2. **I_2 forms a blue colored substance with starch**. Any time variations will be due to the changing concentrations of I^- or $S_2O_8^{2-}$.

Materials: 3 25 mL graduated cylinders
100 mL beaker
Stopwatch
Glass stirring rod

Chemicals: 0.20 M KI, a source of I^-
0.0050 M $Na_2S_2O_3$, a source of $S_2O_3^{2-}$ (this solution also contains some starch)
0.10 M $(NH_4)_2S_2O_8$, a source of $S_2O_8^{2-}$
0.01 M $CuSO_4$ (acts as a catalyst)

PROCEDURE

Part I – Differing Concentrations

1) Obtain 3 clean 25mL graduated cylinders and label them A, B and C. Use them to carefully measure the volumes of the three stock solutions to be used for each mixture (as detailed in Table I below)

TABLE I – Reaction of Mixtures

Mixture	Solution A (mL)	Solution B (mL)	Solution C (mL)	Reaction Time (sec)	Reaction Time with catalyst (sec)	Initial Concentration of I^- (M)	Initial Concentration of $S_2O_8^{2-}$ (M)
1	20.0	10.0	20.0				
2	10.0	10.0	20.0				
3	20.0	5.0	20.0				
4	20.0	10.0	10.0				
5	30.0	10.0	20.0				

Note: This will serve as your data table. Additional mixtures will be provided on the board.

2) Start by preparing mixture 1. Pour 10 mL of solution B into a clean, dry 100 mL beaker.

3) Get the prescribed amounts of solution A and solution C using the graduated cylinders.

4) Start your stopwatch at exactly the same time that you pour the graduated cylinders A and C into the beaker containing B.

5) Stir the solution continuously. When the first sign of blue appears, record this as your reaction time in Table I above.

6) Repeat steps 2) through 5) for mixtures 2, 3, 4, and 5.

Part II – The catalyst.

1) Repeat steps 2) through 5) of Part I, but this time add 4 drops of 0.01 M $CuSO_4$ (solution D) to the graduated cylinder containing solution C before adding it to the beaker.

2) Record the time required to make a color change in Table I. (Note: It should be faster than the initial time)

ANALYSIS

1. The ions react according to the following reactions:

$$2I^- + S_2O_8^{2-} \rightarrow I_2 + 2SO_4^{2-} \quad \ldots\text{slow}\ldots(1)$$

$$\underline{I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-} \quad \ldots\text{fast}\ldots(2)}$$

$$S_2O_8^{2-} + 2S_2O_3^{2-} \rightarrow 2SO_4^{2-} + S_4O_6^{2-} \quad \ldots\text{overall}\ldots$$

   Explain the changes in the rate of the reaction using the framework of the molecular level. **[4 MARKS]**

2. Identify whether increasing the concentration of I^- increases or decreases the rate of the reaction. **[1 MARK]**

3. Identify whether increasing the concentration of $S_2O_8^{2-}$ increases or decreases the rate of the reaction. **[1 MARK]**

4. Explain how increasing or decreasing the concentration affects the reaction rate using evidence that you observed and your knowledge of molecular kinetics. Include the overall equation. **[4 MARKS]**

5. Explain how the catalyst affects the reaction rate using evidence that you observed and your knowledge of catalysts. Include the overall equation involving the catalyst. **[4 MARKS]**