Iodine Clock and an Introduction to Kinetics

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Introduction:

This introduction to kinetics lab makes use of the iodine clock reaction to instruct students at an early undergraduate chemistry level with a set of mini-experiments carried out in groups of 4-5 students. At standard conditions, the iodine clock reaction takes ~40 seconds to complete and ends with a stark transition from a colorless solution to dark blue/purple. Through modifying the temperature and reactant ratios, students can qualitatively observe the Arrhenius Equation and reaction orders of the reactants. The lab is designed for 2 lab sessions of (2-3) hours and contains several exercises designed for entire class collaboration.

Materials (per Group):

2x 50mL beaker
2x 25mL graduated cylinder
1x 100 mL beaker
1x Red+blue coloring dye
1x Thermometer
2x 500 mL beaker
2x 1L beaker
1x Hot plate
1x Stir bar

PART ONE: Observable Kinetics**

The instructor may want to perform the activities of part one in front of the class. The endpoint of the coloring dye experiments is subjective, and the instructor may want to define reaction completion. Secondly, this expedites the first two activities for more advanced lab groups.

Activity 1: Procedure

Fill a 50 mL beaker with deionized water.

Prepare stopwatch, or watch a clock.

Add 1 drop of coloring dye.

Record time when the dye has dispersed into the solution. (The lab should agree upon an endpoint. A tutorial is advisable.)

Time: ________________________

Fill the 100 mL beaker with deionized water.

Prepare stopwatch, or watch a clock.

Add 1 drop of coloring dye.

Record time when the dye has dispersed into the solution.

Time: ________________________

<u>Activity 1: Questions</u>

For all questions on the WS, please work together and do not hesitate to ask for help.

What was the difference in timescale between the two experiments?

Why wasn't/was there a difference in the timescale?

Would you classify this reaction as a physical or chemical reaction? Explain.

What would be required for this to be considered a chemical reaction?

Explain why the coloring dye dispersed in solution.

<u>Activity 2: Procedure</u>

Make sure to rinse off all glassware with deionized water to avoid cross-contamination. Use one of the 1L beakers as aqueous waste.

Pour 300 mL Stock Solution A into a 500 mL beaker, label the beaker.

Pour 300 mL Stock Solution B into the other 500 mL beaker, label the beaker.
If you run out of A or B refill with the stock solutions.

Using the graduated cylinder, pour 25 mL of A and B into separate 50 mL beakers (label these too)

Using the clock or a stopwatch, start recording the time as another group member pours the two solutions into the 100 mL beaker.

Record the time when the solution becomes purple-black.
Time: ________________________

<u>Activity 2: Questions</u>

Was this a physical or chemical reaction? Explain.

If I mixed red and blue coloring dye I could make a purple solution. Is this a chemical reaction?

Why does the solution go from clear to spontaneously purple after ~40 seconds? What does this tell you about the chemical reaction.
Hint- look at Appendix A at the back

PART TWO: Temperature Effects
Use the thermometer to check the temperature.

Activity 1: Procedure

Create an ice bath with the other 1L beaker.
  Fill about ½ way with ice.
  Fill with water to create a slurry.
  If available add salt.

Fill one of the 50 mL beakers with deionized water.

Place this beaker into the ice bath.

Fill another 50 mL beaker with deionized water.

Place this beaker on the hot plate and heat to about 30-40 °C

Remove both beakers from ice bath/hot plate and recording with stopwatch/clock add 1 drop of red/blue dye and record the time for it to diffuse.

Time Cold: ____________________________

Time Hot: ________________________

Time (from Part One): ________________________

Activity 1: Questions

What did you observe between the different rates of diffusion at different temperatures?

What do you expect to happen if you continue to increase/decrease the temperature? Are there limits?

What would happen if the beaker was heating or cooling during the diffusion process? Explain.

<u>Activity 2: Procedures</u>

For this reaction we will be using 25 mL of A and a 25 mL 2:3 dilution of B to deionized water.
<u>Explain to a grad student how you will make the diluted B solution before you start heating or cooling.</u>

Select and write on the board 4 temperatures between 51 °C and 0 °C that you want to examine for the Iodine Clock reaction.

Measure out the 25 mL solution of A and the diluted 25 mL solution of B in graduated cylinders.

Pour into separate 50 mL beakers.

Heat or cool both beakers to the desired temperature.
Monitor with the thermometer.

Prepare to start recording the time with a stopwatch or clock.

Pour beakers of solutions A and B into the 100 mL beaker.

Temperature One: ___________________________
Time: ___________________________

Temperature Two: ___________________________
Time: ___________________________

Temperature Three: ___________________________
Time: ___________________________

Temperature Four: ___________________________
Time: ___________________________

Record the times and temperatures on the x-y plot on the whiteboard.

Activity 2: Questions

Draw the graph (Temperature v.s. Time) onto the worksheet.

Arrhenius gives the following equation for temperature dependance of rate:

$$k = Ae^{-\frac{E_a}{RT}} \quad ,$$

where k is the rate constant, A is an exponential prefactor (constant), e is the mathematical value 2.718…, E_a is the reaction energy barrier, R is the ideal gas constant, and T is the temperature.

From this equation, how do you expect temperature to effect the rate constant? Does this behavior match the graph?

Examine the other terms of the Arrhenius Equation, explain them in detail.

Arrhenius Equation aside, can you describe why temperature influences rate at a particle level?

What happens at above 50 °C? Why do you think this happens?
Hint- See Appendix A for more information.

Some chemical processes are exothermic. What does exothermic mean? How might this influence the rate of reaction and why might this be a concern?

PART THREE: Concentration Effect on Kinetics

Procedure:

In this experiment, you will measure the time of reaction until color change. Procedure will be repeated four times, and the reaction mixture will be diluted with a different amount of water each time.

Choose 4 dilutions to perform, and write them below in the format A : B : H_2O.

Volumes of solutions A and B will stay the same, but you will vary the amount of water. Choose volumes of water between 0 mL and 50 mL.

Measure 25 mL of A into the 50 mL beaker using a graduated cylinder.

Measure 25 mL of B into the other 50 mL beaker using a graduated cylinder.

Measure desired amount of water into the 100 mL beaker. Record this volume below in the dilution ratio format.

Prepare to measure the time of reaction.

Simultaneously pour the contents of both beakers into the 100 mL beaker containing water. Start the timer/stopwatch/smartphone.

Record time when the solution becomes purple-blue.

Dilution One: ___________________________
Time: ___________________________

Dilution Two: ___________________________
Time: ___________________________

Dilution Three: ___________________________
Time: ___________________________

Dilution Four: ___________________________
Time: ___________________________

Record your dilution factors and reaction times on the graph.

Questions

Draw the graph (Concentration v.s. Time) here:

Explain the trend observed with the reaction rate.

Why does the reaction rate depend on the concentration of the reactants?

Imagine the following reaction:

$$A + B \rightarrow AB$$

How would the rate change if:

(a) I increased the amount of A and B uniformly?
(b) I decreased the amount of A and B uniformly?
(c) I increase the amount of A?
(d) I increase the amount A and decrease the amount of B uniformly?
(e) I increase the amount of AB?

Explain each response.

PART FOUR: Temperature and Concentration Together

<u>Procedure:</u>

In this part of the lab you will ***design your own experiment****. Choose to* ***hold either the temperature or the dilution of B constant*** *and then* ***vary the other parameter****. For example, hold at 30 °C and do 5 different dilutions of B.*

Create 5 data points and construct a graph on the whiteboard.

Choose which parameter you will hold constant:

Temperature / Dilution Factor

Data Point One: ___________________________
Time: ___________________________

Data Point Two: ___________________________
Time: ___________________________

Data Point Three: ___________________________
Time: ___________________________

Data Point Four: ___________________________
Time: ___________________________

Data Point Five: ___________________________
Time: ___________________________

<u>Questions:</u>

What were you attempting to learn from your data set? What did you learn?

How would you optimize the reaction to go as fast as possible? What are your limits?

Can you think of some real-life applications of these sorts of kinetics experiments? Why might scientist do these?

There is unavoidably some error in your results. Explain possible sources of error, and how (given unlimited funding) to remove them. Is there some error that you cannot control?

APPENDIX A: The Iodine Clock Reaction

For the majority of the experiments it is much easier (and still insightful) to think of the reaction studied as an A + B -> AB. However, the actual mechanism is more complicated.

When you mix solutions A and B together the following sets of reactions occur. Note that solution A contains the potassium iodate and solution B contains everything else.

First all the salts are dissolved by water into aqueous species.

For example:
Potassium Iodate dissociates into the corresponding ions: $KIO_3 \rightarrow K^+ + IO_3^-$
Sodium Bisulfide dissociates into the corresponding ions: $NaHSO_3 \rightarrow Na^+ + HSO_3^-$

The first reaction (below) is the **rate determining step**. In this reaction, the Iodate anion (IO_3^-) reacts with the Bisulfite (HSO_3^-) to produce Iodide (I^-):

$$IO_3^- + 3HSO_3^- \rightarrow I^- + 3HSO_4^-$$

The excess Iodate (IO_3^-) then rapidly oxidizes iodide (I^-) to iodine (I_2):

$$IO_3^- + 5I^- + 5H^+ \rightarrow 3I_2 + 3HSO_4^-$$

Iodine (I_2) is then reduced by residual Bisulfite (HSO_3^-) back to Iodide (I^-):

$$I_2 + HSO_3^- + H_2O \rightarrow 2I^- + HSO_4^- + 2H^+$$

This cyclic process of the oxidation by excess Iodate and reduction by excess Bisulfite continues until one of the two species is used up.

What makes the blue color in solution is a final step where a non-bonding interaction between Iodine (I_2) and starch. This reaction step is much slower than the oxidation by Bisulfite, and can only occur once the Bisulfite is extinguished. Iodide (I^-) cannot form this complex with starch.

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-Thank You-