Lab: Kinetics and Equilibrium

FOR THE TEACHER

Summary
In this lab, students will investigate the reaction of the hydrogen sulfite ion (HSO$_3^-$) and the iodate ion (IO$_3^-$) to determine the effect that changing concentration and temperature has on the reaction rate.

Grade Level
High School

AP Chemistry Curriculum Framework
- **Big Idea 4:** Rates of chemical reactions are determined by details of the molecular collisions.
  - 4.1 The student is able to design and/or interpret the results of an experiment regarding the factors (i.e., temperature, concentration, surface area) that may influence the rate of a reaction.
- **Big Idea 6:** Any bond or intermolecular attraction that can be formed can be broken. These two processes are in a dynamic competition, sensitive to initial conditions and external perturbations.
  - 6.3 The student can connect kinetics to equilibrium by using reasoning about equilibrium, such as Le Chatelier’s principle, to infer the relative rates of the forward and reverse reactions.

NGSS Standards
This activity will help prepare your students to meet the performance expectations in the following standards:
- HS-PS1-5: Apply scientific principles and evidence to provide an explanation about the effects of changing the temperature or concentration of the reacting particles on the rate at which a reaction occurs.
- HS-PS1-6: Refine the design of a chemical system by specifying a change in conditions that would produce increased amounts of products at equilibrium.

Objectives
By the end of this lab, students should be able to
- Explain the relationship between reaction time and rate.
- Understand the relationship between concentration and reaction rate and why it exists.
- Comprehend the relationship between temperature and reaction rate and why it exists.
- Justify whether the process of heat exchange is endothermic or exothermic.
- Interpret the direction of equilibrium shift.
- Apply Hess’ Law to the two-step reaction process to determine the net reaction.
- Write an equilibrium constant expression for the net reaction.

Chemistry Topics
This lab supports students’ understanding of
- Equilibrium
- LeChatelier’s Principle
- Hess’ Law
- Kinetics
- Reaction Time
- Reaction Rate
- Kinetic Molecular Collision Theory
- Effective and ineffective collisions
- Net Reactions
- Volume
- Concentration
- Temperature
- Endothermic Reactions

Submitted by
Lindsay Davis
The Baldwin School
Bryn Mawr, PA

Thanks to:
Dow Chemistry Teacher Summit
Time
Teacher Preparation: 60 minutes
Lesson: 75 minutes

Materials (per lab group)
- 1- 250 mL beaker
- 2- 50 mL beakers
- Stop watch
- 2- 10 mL graduated cylinders
- 2 test tubes
- Thermometer
- Hot Plate
- Ice
- Distilled water
- 52mL Solution A—0.040 M $\text{KIO}_3$
- 60mL Solution B—0.040 M $\text{NaHSO}_3$, 0.010 M $\text{H}_2\text{SO}_4$, and starch solution
- 1L Volumetric Flask

Safety
- Always wear safety goggles when handling chemicals in the lab.
- Students should wash their hands thoroughly before leaving the lab.
- When students complete the lab, instruct them how to clean up their materials and dispose of any chemicals.
- Exercise caution when using a heat source. Hot plates should be turned off and unplugged as soon as they are no longer needed.
- An operational fire extinguisher should be in the classroom.
- When working with acids, if any solution gets on students’ skin, they should immediately alert you and thoroughly flush their skin with water.

Teacher Notes
- This lab is best for honors level chemistry, or AP chemistry but can be adapted as needed for a standard level.
- Solution A is made by dissolving 8.6g of KIO$_3$ in distilled water and diluting to a volume of 1L in a volumetric flask. Each lab group will need 52mL of Solution A.
- Solution B is made by dissolving 8g of soluble starch in 500mL of boiling distilled water. Then, when cooled somewhat, add 4g of NaHSO$_3$ and 2 mL of 1M H$_2$SO$_4$ to the solution and stir until dissolved. Transfer to 1L volumetric flask and dilute to 1L. Each lab group will need 60mL of Solution B.
- You can make the solutions ahead of time as they have a good shelf life.
- Label beakers, test tubes, and graduated cylinders with “A” and “B” so students do not cross contaminate the solutions.
- Remind students to make their Data Tables before coming to the lab.
- Students can be put into groups of two to three.
- Have the students be aware of human reaction time in regard to starting and stopping the timer. The same student should time every trial of the entire lab to increase consistency of data.
- When pouring the solutions back and forth, students should pour over the sink in case of spillage.
- Tell students not to be scared when the color changes as it happens very suddenly. You don’t want the students to drop the beakers of chemicals.
• **Answers to Post-Lab Questions:**

1. Explain on a molecular level how and why the rate of a reaction changes as concentration changes.
   - *As concentration increases, the rate of a reaction increases because there are more particles colliding into each other more often.***

2. Explain on a molecular level how and why the rate of a reaction changes as temperature changes.
   - *As temperature increases, the rate of a reaction increases because the particles are moving more quickly and are colliding into each other more often.***

3. Explain on a molecular level how and why the activation energy of a reaction changes as temperature changes.
   - *As the temperature of the reactants increase, their energy is closer to the energy of the activated complex. So, less activation energy is needed to reach the activated complex and form product.***

4. How is the time of a reaction related to the rate of a reaction? Explain.
   - *As the time of a reaction increases, the rate of a reaction decreases because rate is measured as concentration per unit time.***

5. In order to investigate the effect Solution A’s concentration had on the rate of the reaction, the composition and amount of Solution B was not changed. Explain why.
   - *There must be one control factor and one variable factor. Otherwise, it would not be clear which chemical’s concentration affected the rate of the reaction.***

6. Applying Hess’ Law to the two-step reaction process (by adding the Step 1 & 2 reactions together from the introduction…but you are not manipulating the reactions…you are just adding them together), write the net reaction that occurs in this lab.
   - 2IO$_3^-$ (aq) + 3HSO$_3^-$ (aq) + 3H$^+$ (aq) + 4I$^-$ (aq) $\rightarrow$ 3I$_2$ (aq) + 3SO$_4^{2-}$ (aq) + 3H$_2$O (l)

7. Write the equilibrium constant expression for the net reaction.
   - $K = [I_2] [SO_4^{2-}]^3/[IO_3^-] [HSO_3^-] [H^+] [I^-]^4$

8. Since the solutions in Part B were heated, was this reaction endothermic or exothermic? Justify your answer.
   - Since the reactants were heated on a hot plate and gained/absorbed energy, the reaction was endothermic.

9. Based on your answer from #8, which direction did the equilibrium shift as the solution was heated? Justify your answer using LeChatelier’s Principle.
   - Since the reaction was endothermic, the heat is on the reactant side of the reaction. Therefore, when heat is applied, the reaction shifts forward towards the products.
FOR THE STUDENT
Lesson
Kinetics and Equilibrium Lab

Background
We will investigate the reaction of the hydrogen sulfite ion (HSO$_3^-$) and the iodate ion (IO$_3^-$) to determine the effect of changing concentration and temperature has on the reaction rate.

**The reaction occurs in a two-step process:**
1. The iodide ion is generated by the following reaction between the iodate and bisulfite:
   **Step 1:** IO$_3^-$ (aq) + 3HSO$_3^-$ (aq) → I$^-$ (aq) + 3SO$_4^{2-}$ (aq) + 3H$^+$ (aq)

2. The iodate in excess will oxidize the iodide generated above to form iodine:
   **Step 2:** 5I$^-$ (aq) + 6H$^+$ (aq) + IO$_3^-$ (aq) → 3I$_2$ (aq) + 3H$_2$O (l)

*However, the iodine produced in Step 2 is reduced immediately back to iodide by the bisulfite:

I$_2$ (aq) + HSO$_3^-$ (aq) + H$_2$O (l) → 2I$^-$ (aq) + HS$_2$O$_4^-$ (aq) + 2H$^+$ (aq)

**Finally, when the bisulfite is fully consumed, the iodine will survive (i.e., no reduction by the bisulfite) to form the dark blue complex with starch. This is why a certain amount of time goes by (even after the reactants are mixed) before the dark blue color is formed.

***The presence of I$_2$ (aq) can be detected by adding starch to the solution because the starch changes to a dark blue color in the presence of iodine. We can determine the rate of the reaction by timing how long it takes for the dark blue color to appear.

Objective
To determine the effect of changing concentration and temperature has on the rate of the reaction between the hydrogen sulfite ion (HSO$_3^-$) and the iodate ion (IO$_3^-$).

Materials
- 1- 250 mL beaker
- 2- 50 mL beakers
- Stop watch
- 2 - 10 mL graduated cylinders
- 2 test tubes
- Thermometer
- Hot Plate
- Ice
- Solution A—0.040 M KIO$_3$
- Solution B—0.040 M NaHSO$_3$, 0.010 M H$_2$SO$_4$, and starch solution

Safety
- Always wear safety goggles when handling chemicals in the lab.
- Wash your hands thoroughly before leaving the lab.
- Follow the teacher’s instructions for cleanup of materials and disposal of chemicals.
• Exercise caution when using a heat source. Hot plates should be turned off and unplugged as soon as they are no longer needed.

• When working with acids and bases, if any solution gets on your skin immediately rinse the area with water.

Pre-Lab

• Make Data tables prior to coming to lab. See Data Table section below for details.

Procedure

PART A

1. Using a clean, dry, 10 mL graduated cylinder, measure exactly 10.0 mL of Solution A and pour it into a 50 mL beaker.

2. Using a second 10 mL graduated cylinder, measure exactly 10.0 mL of Solution B and pour it into a second 50-mL beaker. (Make sure that you do not mix up which graduated cylinder goes with which solution.)

3. Prepare to time the reaction. While one lab partner pours Solution A into Solution B, the second partner should immediately start timing the reaction. Pour the solutions back and forth three times from one beaker to the other to ensure thorough mixing. Then allow the mixture to stand. At the instant a color change occurs, the partner timing the reaction should stop the timer and note the elapsed time. Record this in your data table. Rinse and dry the beakers and graduated cylinders.

4. Using a clean graduated cylinder, measure exactly 10.0 mL of Solution B and pour it into one of the beakers. Using another clean graduated cylinder, measure exactly 7.0 mL of Solution A and then add exactly 3.0 mL of distilled water on top of it. Pour this diluted solution into another beaker. Follow the instructions from step 3 for mixing the solutions and timing the reaction. Record the elapsed time in your data table. Rinse and dry the beakers and graduated cylinders.

5. Repeat step 4 one more time, this time using 5.0 mL of Solution A and 5.0 mL of distilled water. The volume of Solution B remains unchanged (again use 10.0 mL). Follow the instructions from step 3 for mixing the solutions and timing the reaction. Record the elapsed time in your data table. Rinse and dry the beakers and graduated cylinders.

PART B

6. Measure 10.0 mL of Solution A into one test tube. Add 10.0 mL of Solution B into a second test tube.

7. Fill a 250 mL beaker approximately halfway with cold tap water. Add ice cubes to the water and place the two test tubes in the ice water bath. Put the thermometer into one of the chemicals to measure its temperature. Let them stand until the solution is somewhere between 5-10°C. You can assume that the temperature of one solution is the same as the temperature of the other. Always wash and dry the thermometer after removing it from a solution to prevent contamination.

8. When the solution temperature is close to 5-10°C, prepare to time the reaction. Record the exact temperature in your data table. One lab partner should start timing the reaction the instant the second partner pours Solution A into Solution B. Quickly pour the mixture back and forth from test tube to test tube three times and return the mixture to the ice-water bath. At the instant a color change occurs, stop the timer and note the time elapsed. Discard the mixture as instructed. Wash and dry the test tubes.

10. Fill a 250 mL beaker approximately halfway with tap water. Place the two test tubes in the water bath. Put the thermometer into one of the chemicals to measure its temperature. Add ice cubes or heat the water bath to reach a temperature of somewhere between 20-30°C. You can assume that the temperature of one solution is the same as the temperature of the other. Repeat step 8 at this new temperature. Record your observations in your data table. Always wash and dry the thermometer after removing it from a solution to prevent contamination.

11. Repeat steps 9 & 10 using a warm water bath to heat the chemicals to somewhere between 40-50°C. Use warm tap water and, if necessary, a hot plate to prepare this warm bath.

12. Wash the test tubes, dry them as best as possible, and then put them in the test tube rack upside down to dry with a paper towel underneath the test tube rack to catch the dripping water.

**Data Tables**
- Make two data tables, one for Part A and one for Part B.
- Part A data table should have three columns (labeled: Volume of Solution A, Volume of Solution B, and Time of Reaction) and three rows for data since you are doing the experiment three times (with varying concentrations).
- Part B data table should have two columns (labeled: Temperature of Chemicals and Time of Reaction) and three rows for data since you are doing the experiment three times (with varying temperatures).

**Post-Lab and Analysis Graphs**
Plot the following data on two separate graphs. Draw a line through the plotted points to produce a best-fit curve showing the effects of concentration of reactants and temperature on reaction rate. Be sure to title each graph and label the axes.
- Time vs. Volume of Solution A (KIO₃)
- Time vs. Temperature

**Post-Lab and Analysis Questions**
1. Explain on a molecular level how and why the rate of a reaction changes as concentration changes.

2. Explain on a molecular level how and why the rate of a reaction changes as temperature changes.

3. Explain on a molecular level how and why the activation energy of a reaction changes as temperature changes.

4. How is the time of a reaction related to the rate of a reaction? Explain.

5. In order to investigate the effect Solution A’s concentration had on the rate of the reaction, the composition and amount of Solution B was not changed. Explain why.

6. Applying Hess’ Law to the two-step reaction process (by adding the Step 1 & 2 reactions together from the introduction…but you are not manipulating the reactions...you are just adding them together), write the net reaction that occurs in this lab.
7. Write the equilibrium constant expression for the net reaction.

8. Since the solutions in Part B were heated, was this reaction endothermic or exothermic? Justify your answer.

9. Based on your answer from #8, which direction did the equilibrium shift as the solution was heated? Justify your answer using LeChatelier’s Principle.

Conclusion
An overall summary of the lab as well as concepts covered in the lab should be included. Sample questions to include in your conclusion: What did you do? What did you find? What do you think your results mean? Connect the hypothesis, procedures, data, and results together with concepts covered in the lab. Someone should be able to read your conclusion and essentially know what you did in the lab and why you did it.

Error Analysis
A minimum of three possible errors is needed. List the errors and how each one affected the outcome/results of your lab to the best of your knowledge. This is the ANALYSIS part of this section and, in grading, is the most important part. Do not include errors which are totally insignificant or that are mathematical, as these will not be accepted.