

# Clock Reaction Race

## Reaction Pathways

## Introduction

Use this dramatic iodine clock reaction to demonstrate the effect of concentration, temperature, and a catalyst on the rate of a chemical reaction.

## Concepts

- Kinetics/Catalysts
- Clock Reactions

## Materials

Potassium iodate solution, 0.2 M,  $\text{KIO}_3$ , 325 mL  
Sodium metabisulfite solution, 0.2 M,  $\text{Na}_2\text{S}_2\text{O}_5$ , 60 mL<sup>†</sup>  
Starch solution, 2%, 180 mL<sup>†</sup>  
Sulfuric acid solution, 0.1 M,  $\text{H}_2\text{SO}_4$ , 10 mL  
Distilled or deionized water, 1105 mL  
Graduated cylinders, 10-mL, 2  
Graduated cylinders, 50-mL, 2  
Graduated cylinders, 100-mL, 2  
watch

<sup>†</sup>These solutions must be prepared fresh.

Beakers, 250-mL, 6  
Beakers, 400-mL, 6  
Bucket or utility pan (for ice bath)  
Flask, Erlenmeyer, 500-mL  
Hot Plate  
Ice  
Thermometer  
Timer or stop-

## Safety Precautions

*Potassium iodate is an oxidizer. It is moderately toxic by ingestion and a body tissue irritant. Sodium metabisulfite is a skin and tissue irritant. Sulfuric acid solution is corrosive to eyes, skin and other tissues. Wear chemical splash goggles, a chemical-resistant apron and chemical-resistant gloves. Please review current Material Safety Data Sheets for additional safety, handling, and disposal information.*



## Preparation

### Preparation of “A” Solutions:

1. Label six 400-mL beakers as 1A, 2A, 3A, 4A, 5A, and 6A.
2. Pour 50 mL of the 0.2 M potassium iodate solution plus 150 mL of distilled or deionized water into beakers 1A, 4A, and 5A.
3. Pour 100 mL of the 0.2 M potassium iodate solution plus 100 mL of distilled or deionized water into beaker 2A.
4. Pour 25 mL of the 0.2 M potassium iodate solution plus 175 mL of distilled or deionized water into beaker 3A.
5. Place beaker 4A on a hot plate and warm it to about 45 °C. Place beaker 5A into an ice bath and cool it to about 10 °C.
6. Pour 50 mL of the 0.2 M potassium iodate solution, 140 mL of distilled or deionized water, and 10 mL of the 0.1 M sulfuric acid solution into beaker 6A.

## Clock Reaction Race *continued*

7. Check to make sure each beaker contains 200 mL of solution. Use the table below as a guide.

	Beaker 1A	Beaker 2A	Beaker 3A	Beaker 4A	Beaker 5A	Beaker 6A
Potassium Iodate Solution, 0.2 M	50 mL	100 mL	25 mL	50 mL	50 mL	50 mL
Distilled or Deionized Water	150 mL	100 mL	175 mL	150 mL	150 mL	140 mL
Sulfuric Acid Solution, 0.1 M	—	—	—	—	—	10 mL
[KIO <sub>3</sub> ] before mixing with Solution B	0.05 M	0.1 M	0.025 M	0.05 M	0.05 M	0.05 M
[KIO <sub>3</sub> ] after mixing with Solution B	0.04 M	0.07 M	0.02 M	0.04 M	0.04 M	0.04 M

### Preparation of “B” Solutions:

1. Add 420 mL of distilled or deionized water to a 500-mL flask that contains 16 g of solid sodium metabisulfite. Stopper the flask and shake well until the solid is dissolved. This is now a 0.2 M sodium metabisulfite solution. Prepare this solution fresh.
2. Prepare 1300 mL of 2% starch solution by making a smooth paste of 26 g soluble starch and 130 mL distilled or deionized water. Pour the paste into 1170 mL of boiling water while stirring. Cool to room temperature before using. Starch solution has a poor shelf life and will form mold if kept for too long. Prepare this solution fresh.
3. Label six 250-mL beakers as 1B, 2B, 3B, 4B, 5B, and 6B.
4. Pour 10 mL of the 0.2 M sodium metabisulfite solution, 30 mL of the starch solution, and 40 mL of distilled or deionized water into each of these beakers. Stir each solution.

	Beaker 1A	Beaker 2A	Beaker 3A	Beaker 4A	Beaker 5A	Beaker 6A
Sodium metabisulfite Solution, 0.20 M	10 mL	10 mL	10 mL	10 mL	10 mL	10 mL
Starch Solution	30 mL	30 mL	30 mL	30 mL	30 mL	30 mL
Distilled or Deionized Water	40 mL	40 mL	40 mL	40 mL	40 mL	40 mL
[Na <sub>2</sub> S <sub>2</sub> O <sub>5</sub> ] before mixing with Solution A	0.025 M	0.025 M	0.025 M	0.025 M	0.025 M	0.025 M
[Na <sub>2</sub> S <sub>2</sub> O <sub>5</sub> ] after mixing with Solution A	0.007 M	0.007 M	0.007 M	0.007 M	0.007 M	0.007 M

## Procedure

1. Make a data table on the chalkboard or on an overhead like the following one.

	<b>1 (Control)</b>	<b>2</b>	<b>3</b>	<b>4</b>	<b>5</b>	<b>6</b>
[KIO <sub>3</sub> ]	0.04 M	0.07 M	0.02 M	0.04 M	0.04 M	0.04 M
[Na <sub>2</sub> S <sub>2</sub> O <sub>5</sub> ]	0.007 M	0.007 M	0.007 M	0.007 M	0.007 M	0.007 M
Temperature	Room Temp	Room Temp	Room Temp	Warm	Cool	Room Temp
Catalyst Added	No	No	No	No	No	Yes
Time Until the Appearance of the Blue Color (sec)						

### *Control Reaction:*

2. Place the 400-mL beakers containing the “A” Solutions on the demonstration table in order.
3. Pour Solution 1B into Solution 1A. Carefully time the reaction with a stopwatch or timer. Measure the time from when the two solutions are mixed until the appearance of the blue color. Record the time in column 1 of the data table. The rate of this reaction will be the rate against which the others are compared. This is the control.

### *Effect of Concentration on the Rate:*

4. Pour Solution 2B into Solution 2A. Measure the time until the appearance of the blue color. Solution 2A is twice as concentrated as Solution 1A, so this rate should be about twice that of the rate in the control reaction (the time will be half that of the control). Record the time in column 2 of the data table.
5. Pour Solution 3B into Solution 3A. Measure the time until the appearance of the blue color. Solution 3A is half as concentrated as Solution 1A, so this rate should be about half that of the rate in the control reaction (the time should be double that of the control). Record the time in column 3 of the data table.

### *Effect of Temperature on the Rate:*

6. Carefully remove Solution 4A from the hot plate. Measure the temperature of solution 4A. Immediately pour Solution 4B into Solution 4A. Measure the time until the appearance of the blue color. Because Solution 4A was warmed, the rate of reaction should be faster than that of the control reaction (the time should be less than that of the control). Record the temperature of Solution 4A and the time in column 4 of the data table.
7. Remove Solution 5A from the ice bath. Measure the temperature of Solution 5A. Immediately pour Solution 5B into Solution 5A. Measure the time until the appearance of the blue color. Because Solution 5A was cooled, the rate of reaction should be slower than that of the control reaction (the time should be more than that of the control). Record the temperature of Solution 5A and the time in column 5 of the data table.

### *Effect of a Catalyst on the Rate:*

8. Pour Solution 6B into Solution 6A. Measure the time until the appearance of the blue color. The sulfuric acid solution added to Solution 6B serves as a catalyst in the reaction, so the rate of this reaction should be faster than that of the control reaction (the time should be less than that of the control). Record the time in column 6 of the data table.

## Results

<b>Beaker</b>	<b>1</b>	<b>2</b>	<b>3</b>	<b>4</b>	<b>5</b>	<b>6</b>
Time Until the Appearance of the Blue Color (sec)	6 sec	3 sec	12 sec	4 sec	8 sec	2 sec

## Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. The final solutions may be reduced with sodium thio-sulfate solution according to Flinn Suggested Disposal Method #12a. Add just enough reducing agent to decolorize the blue color of the starch-iodine complex.

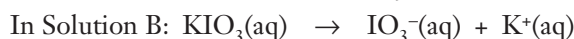
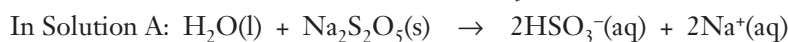
## Tips

- This demonstration may be more dramatic if six student volunteers are recruited to help. They can each be given prepared solutions and asked to mix the solutions simultaneously in front of the class. This twist requires the use of six stop watches or timers.
- All concentrations/dilutions in the *Preparation* section were calculated using the equation  $M_1V_1 = M_2V_2$ . Concentrations before and after mixing are different because combining the solutions dilutes each reactant.

## Discussion

In this reaction, potassium iodate and sodium metabisulfite react to form iodine. The starch solution serves as an indicator of the end of the reaction, forming a dark-blue colored starch-iodine complex in the presence of iodine. The chemical pathway for the formation of iodine is complicated and not completely understood, but the following mechanism serves as an outline.

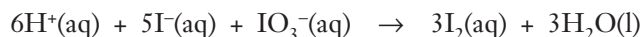
Step 1: Solution A and Solution B contribute hydrogen sulfite ions,  $\text{HSO}_3^-(\text{aq})$ , and iodate ions,  $\text{IO}_3^-(\text{aq})$ , to the solution.



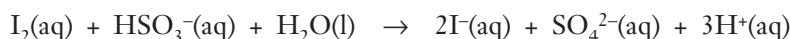
Step 2: The iodate ions react with the hydrogen sulfite ions to produce iodide ions,  $\text{I}^-(\text{aq})$ .



Step 3: In the presence of hydrogen ions,  $\text{H}^+$ , the iodide ions react with excess iodate ions to produce iodine,  $\text{I}_2(\text{aq})$ .



Step 4: Before the iodine can react with the starch to produce a dark-blue colored complex, it immediately reacts with any hydrogen sulfite ions still present to form iodide ions.



Step 5: Once all of the hydrogen sulfite ions have reacted, the iodine is then free to react with the starch to form the familiar dark-blue colored complex.

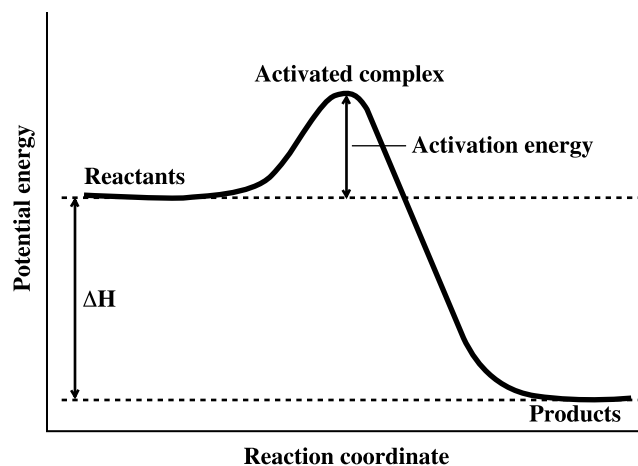


The dark-blue color of the complex is due to the presence of the pentaiodide anion,  $\text{I}_5^-(\text{aq})$ . By itself, the pentaiodide ion is unstable; however, it is stabilized by forming a complex with the starch.

The appearance of the dark-blue color in solution indicates that all of the reactants have been used up and the reaction has gone to completion. Therefore, the rate of reaction can be measured by recording the time to the appearance of the dark-blue color.

In general, the effect of concentration, temperature, and a catalyst on the reaction rate can be understood by looking at the energy profile for a given reaction.

In an energy profile diagram, the left side of the diagram represents reactants, while the right side represents products. In the diagram above, the products are lower in energy than the reactants. In terms of thermodynamics, this reaction is exothermic and should occur spontaneously. However, not all collisions between reactants will produce products. The collision energy for a particular collision must exceed a critical energy for products to be formed. This critical energy is called the *activation energy* and is represented by the hump in the energy profile diagram.



Why must reactant molecules overcome this activation energy, or get over the hump, to reach products? As the reactant molecules approach each other, their atoms interact causing distortion in the bonds of both molecules. This distortion reaches a maximum as the reactants form an *activated complex*, or transition state. The activated complex is a hybrid species formed as the reactant molecules come together and trade atoms to become products. Only those colliding molecules that have enough kinetic energy to reach this distorted intermediate will produce products. As is evident from the energy profile diagram above, the potential energy of this distorted transition state determines the activation energy, or height of the barrier, for a particular reaction. If the barrier is low, almost all colliding molecules will have sufficient energy to reach and overcome the barrier. These reactions will occur spontaneously. If the barrier is high, only a small percentage of collisions will occur with sufficient energy to reach and overcome the barrier and go on to form products. These reactions occur much more slowly than those with a low barrier. In general, as the height of the barrier increases, the rate of the reaction decreases. Therefore, the rate of a reaction depends on the height of the barrier, or the activation energy.

The above description of the energy profile assumes the reaction occurs in a single step. This theory can be applied to multi-step mechanisms, such as the one in this iodine clock reaction, by assuming that one of the steps in the mechanism is much slower than the other steps. This step then determines the rate of the reaction and is called the *rate-determining step*. It is generally a good approximation to say that the energy profile of a reaction describes the energy profile of the rate-determining step.

To increase the rate of a reaction, one of two things must occur: (1) more molecules with sufficient kinetic energy to overcome the barrier must be involved in the reaction to produce a higher number of successful collisions, or (2) the activation energy must be decreased.

One way to obtain a higher number of successful collisions is to increase the concentration of reactant molecules. In beaker #1, the ratio of potassium iodate molecules to sodium metabisulfite molecules is 5:1. This ratio climbs to 10:1 in beaker #2, while it drops to 2.5:1 in beaker #3. The number of collisions is proportional to the concentration so beaker #2 will have twice as many collisions as beaker #1, while beaker #3 will have half as many collisions as beaker #1. In each case, the same fraction of these collisions will possess sufficient energy to overcome the barrier as before. So, since twice as many collisions are occurring in beaker #2, the rate at which a given concentration of B is converted to products will double. Because half as many collisions are occurring in beaker #3, the rate at which a given concentration of B is converted to products is cut in half.

Another way to obtain more molecules with sufficient energy to overcome the barrier is to increase the temperature. The strong temperature dependence of reaction rates can be understood by looking at the relationship between temperature and energy. The average kinetic energy (KE) of a sample is directly proportional to the temperature (T) of the sample according to the following equation:

$$KE = 3/2 k T$$

where  $k$  is Boltzmann's constant. As the temperature is increased, the average kinetic energy of the sample is increased providing a sample with more molecules that possess enough kinetic energy to reach and overcome the barrier.

To lower the activation energy, a catalyst may be added to the reaction mixture. A catalyst is a substance that, when added to the reaction mixture, participates in the reaction and speeds it up, but is not itself consumed in the reaction. In general, a catalyst provides a modified or new mechanism for the reaction that is faster than the original mechanism. The rate of the catalyzed reaction is faster because the activated complex in the catalyzed mechanism is of lower energy than the activated complex in the original mechanism. Hence the barrier to products is lower in the catalyzed reaction. A greater percentage of reactant molecules will possess the needed energy to successfully collide and overcome the barrier. Therefore, the rate of the reaction is increased.

Although thermodynamics determines whether reactants or products are energetically favored, it is kinetics that determines how fast even an exothermic reaction will occur. The speed at which different reactions occur varies widely, ranging from instantaneous to so slow that it appears as if no reaction occurs at all. The use of catalysts in industrial applications can turn costly (slow) reactions into more efficient (fast) processes. Every year more than a trillion dollars worth of goods are manufactured with the help of catalysts. Without them, fertilizers, pharmaceuticals, plastics, fuels, solvents, and surfactants would be in short supply.

## Connecting to the National Standards

This laboratory activity relates to the following National Science Education Standards (1996):

### *Unifying Concepts and Processes: Grades K–12*

- Systems, order, and organization
- Evidence, models, and explanation
- Constancy, change, and measurement

### *Content Standards: Grades 5–8*

- Content Standard A: Science as Inquiry
- Content Standard B: Physical Science, properties and changes of properties in matter, transfer of energy

### *Content Standards: Grades 9–12*

- Content Standard A: Science as Inquiry
- Content Standard B: Physical Science, chemical reactions, conservation of energy and increase in disorder

## Answers to Worksheet

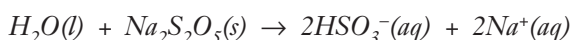
### Sample Results Table

	1 (Control)	2	3	4	5	6
<b>Contents of Beaker A</b>	50 mL 0.2 M $\text{KIO}_3$ , 150 mL water	100 mL $\text{KIO}_3$ , 100 mL water	25 mL $\text{KIO}_3$ , 175 mL water	50 mL $\text{KIO}_3$ , 150 mL water	50 mL $\text{KIO}_3$ , 150 mL water	50 mL $\text{KIO}_3$ , 140 mL water, 10 mL 0.1 M sulfuric acid
<b>Contents of Beaker B</b>	10 mL 0.2 M $\text{Na}_2\text{S}_2\text{O}_5$ , 30 mL starch solution, 40 mL water	10 mL $\text{Na}_2\text{S}_2\text{O}_5$ , 30 mL starch solution, 40 mL water	10 mL 0.2 M $\text{Na}_2\text{S}_2\text{O}_5$ , 30 mL starch solution, 40 mL water	10 mL 0.2 M $\text{Na}_2\text{S}_2\text{O}_5$ , 30 mL starch solution, 40 mL water	10 mL 0.2 M $\text{Na}_2\text{S}_2\text{O}_5$ , 30 mL starch solution, 40 mL water	10 mL 0.2 M $\text{Na}_2\text{S}_2\text{O}_5$ , 30 mL starch solution, 40 mL water
<b><math>[\text{KIO}_3]</math> after mixing solutions</b>	0.04 M	0.07 M	0.02 M	0.04 M	0.04 M	0.04 M
<b><math>[\text{Na}_2\text{S}_2\text{O}_5]</math> after mixing solutions</b>	0.007 M	0.007 M	0.007 M	0.007 M	0.007 M	0.007 M
<b>Temperature</b>	Room temperature	Room temperature	Room temperature	Warm	Cool	Room temperature
<b>Catalyst present?</b>	No	No	No	No	No	Yes
<b>Reaction Time</b>	6 sec	3 sec	12 sec	4 sec	8 sec	2 sec

## Discussion Questions

1. Write the chemical equation for each of the steps in this reaction mechanism.

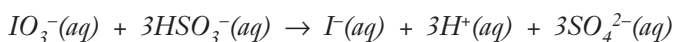
a. Sodium metabisulfite produces hydrogen sulfite ions in water



b. Potassium iodate decomposes

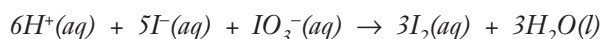


c. Iodate ions react with hydrogen sulfite ions



## Clock Reaction Race *continued*

d. Iodide ions react with iodate ions in the presence of water



In order for a reaction to occur and products to be formed, reactant molecules need to reach a transition state, or activated complex. Only molecules with enough kinetic energy, or activation energy, to reach this state can produce products. The reaction rate depends on the activation energy necessary for molecules to form an activated complex. There are two ways to increase the rate of a reaction: more molecules with sufficient kinetic energy must be present, or the activation energy must be decreased.

2. How does the concentration of reactant molecules affect the reaction rate? Explain.

*The higher the concentration of reactant molecules, the faster the rate at which the reaction will occur. This is because an increase in concentration results in more successful collisions between different reactant molecules. Therefore, the likelihood of these collisions having enough kinetic energy to form an activated complex is also increased.*

3. How does temperature affect the reaction rate? Explain.

*Kinetic energy is directly proportion to temperature. Thus, when temperature is increased, the average kinetic energy of the sample also increases. This results in the presence of more molecules with enough kinetic energy to reach the transition state.*

4. What is a catalyst? What was the catalyst used in this demonstration?

*A catalyst is a substance that speeds up a reaction by lowering the activation energy, but is not consumed in the reaction. The catalyst in this demonstration was sulfuric acid.*

## Reference

Cardillo, C.; *Micro Action Chemistry*; Flinn Scientific: Batavia, IL, 1998; pp 85–87.

## Flinn Scientific—Teaching Chemistry™ eLearning Video Series

A video of the *Clock Reaction Race* activity, presented by Jamie Benigna, is available in *Reaction Pathways*, part of the Flinn Scientific—Teaching Chemistry eLearning Video Series.

## Materials for *Clock Reaction Race* are available from Flinn Scientific, Inc.

Materials required to perform this activity are available in the *Iodine Clock Reaction—Effect of Concentration, Temperature, and a Catalyst on Reaction Rate* available from Flinn Scientific. Materials may also be purchased separately.

Catalog No.	Description
AP4601	Iodine Clock Reaction—Effect of Concentration, Temperature, and a Catalyst on Reaction Rate
P0168	Potassium Iodate Solution, 0.2 M, 500 mL
S0317	Sodium Metabisulfite, 100 g
S0161	Sulfuric Acid Solution, 0.1 M, 500 mL
S0122	Starch, Soluble Potato, 100 g
GP2040	Cylinder, Graduated, 10-mL
GP2044	Cylinder, Graduated, 50-mL
GP2046	Cylinder, Graduated, 100-mL
AP1452	Thermometer, Non-Mercury
AP8386	Flinn Hot Plate, 70 × 70
AP8874	Timer, Large Display
AP9176	Utility Pan, Plastic
W0001	Distilled Water, 1 gallon

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.

# Clock Reaction Race Worksheet

## Results Table

	1 (Control)	2	3	4	5	6
Contents of Beaker A						
Contents of Beaker B						
[KIO <sub>3</sub> ] after mixing solutions						
[Na <sub>2</sub> S <sub>2</sub> O <sub>5</sub> ] after mixing solutions						
Temperature						
Catalyst present?						
Reaction Time						

## Discussion Questions

1. Write the chemical equation for each of the steps in this reaction mechanism.

*a.* Sodium metabisulfite produces hydrogen sulfite ions in water

*b.* Potassium iodate decomposes

*c.* Iodate ions react with hydrogen sulfite ions

*d.* Iodide ions react with iodate ions in the presence of water



## Clock Reaction Race *continued*

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In order for a reaction to occur and products to be formed, reactant molecules need to reach a transition state, or activated complex. Only molecules with enough kinetic energy, or activation energy, to reach this state can produce products. The reaction rate depends on the activation energy necessary for molecules to form an activated complex. There are two ways to increase the rate of a reaction: more molecules with sufficient kinetic energy must be present, or the activation energy must be decreased.

2. How does the concentration of reactant molecules affect the reaction rate? Explain.

3. How does temperature affect the reaction rate? Explain.

4. What is a catalyst? What was the catalyst used in this demonstration?