EXPERIMENT

22

FACTORS AFFECTING REACTION RATES

OBJECTIVE

- To study factors that affect the rates of chemical reactions

INTRODUCTION

Chemical kinetics is the study of chemical reactions rates, how reaction rates are controlled, and the pathway or mechanism by which a reaction proceeds from its reactants to its products.

The rate of a chemical reactions is expressed as a *change* in the concentration of a reactant (or product) as a function of time (e.g., per second)—the greater the change in the concentration per unit of time, the faster the rate of the reaction.

In this experiment we will investigate five factors that can be used and controlled to affect the rate of a chemical reaction.

- the nature of the reactants
- the surface area of the reactants
- the temperature of the chemical system
- the presence of a catalyst
- the concentration of the reactants

Nature of the Reactants

Some substances are naturally more reactive than others and, therefore, undergo rapid chemical changes. For examples, the reaction of sodium metal and water is a very rapid, exothermic reaction (see Experiment 9, Part D), whereas no reaction occurs between gold and water; hydrogen reacts explosively with fluorine at room temperature, whereas the reaction between hydrogen and iodine is extremely slow, even at elevated temperatures.

Surface Area of the Reactants

Generally speaking, the greater the exposed surface area of the reactant, the greater is the reaction rate. For examples, a large piece of coal burns very slowly, but coal *dust* burns rapidly; solid potassium iodide reacts very slowly with solid lead nitrate, but when both are dissolved in solution, the formation of lead iodide is instantaneous.

Temperature of the Reaction

A **rule of thumb**: a 10°C rise in temperature doubles (increases by a factor of two) the rate of a chemical reaction. The added heat not only increases the collision frequency[1] between reactant molecules, but also, and more importantly, increases their kinetic energy. Upon collision of reactant molecules, this kinetic energy is converted into internal energy, which, in turn, is distributed throughout the collision system. This increased internal energy increases the probability for the weaker bonds to be broken and the new bonds to be formed.

Presence of a Catalyst

A **catalyst** increases the rate of a chemical reaction without undergoing any *net* chemical change. Some catalysts increase the rate of only one specific chemical reaction without affecting similar reactions. Others are more general and affect an entire set of similar reactions. Catalysts generally reroute the pathway of a chemical reaction so that this "alternate" path, while perhaps more circuitous, requires less energy to produce the products.

Concentration of Reactants

An increase in the concentration of a reactant generally increases the reaction rate. On occasion, such an increase may have no effect or may even decrease the reaction rate. A quantitative investigation of concentration changes on reaction rate is undertaken in Experiment 23.

TECHNIQUES

Review:	Technique 1, page 9	Inserting glass tubing through a rubber stopper
	Technique 2, page 9	Cleaning glassware
	Technique 5, page 11	Weighing on a balance
	Technique 6b, page 13	Collecting gases
	Technique 8, page 15	Transferring liquids
	Technique 11d, page 21	Heating liquids
	Technique 14a, b, page 25	Quantitative transfer of liquids
	Appendix C	Graphing data

[1] A 10°C temperature rise only increases the collision frequency between reactant molecules by factor of 1.02—no where near the factor of two that is normally experienced.

EXPERIMENTAL PROCEDURE

Ask your instructor which parts of the Experimental Procedure you are to complete.

Caution: *A number of strong acids are used. Handle with care, do not allow them to touch the skin or clothing.*

A. Nature of the Reactants

1. **Different Acids Affect Reaction Rates.** Into four numbered 75-mm test tubes (Figure 22.1) containing approximately 1 mL of 3 M H_2SO_4, 6 M HCl, 6 M HNO_3, and 6 M H_3PO_4 (**Caution:** *avoid skin contact with the acid*), place a 1-cm *polished* strip of magnesium ribbon. Compare the reaction rates and record your observations. ***CAUTION***

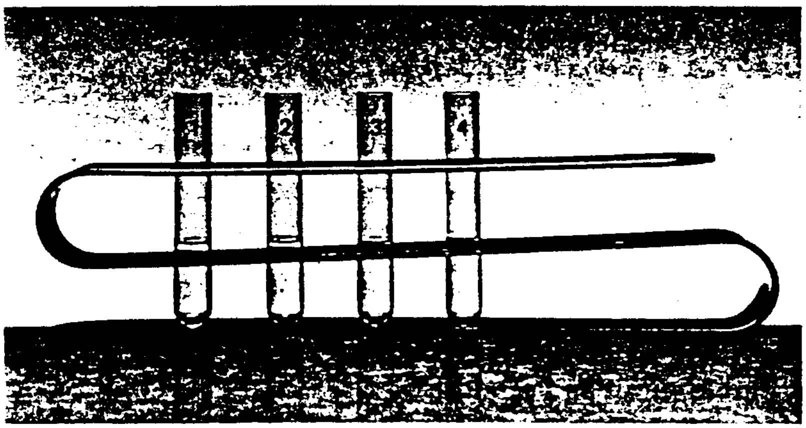


Figure 22.1. *Number test tubes for identification*

2. **Different Metals Affect Reaction Rates.** Place approximately 1 mL of 6 M HCl into each of three numbered 75-mm test tubes. Add *polished* strips of each of the following: zinc to the first, lead to the second, and copper to the third. Compare the reaction rates and record.

B. Surface Area of Reactants

1. **Assemble the Apparatus.** Assemble the apparatus shown in Figure 22.2. Fill a 200-mm test tube with water and invert it (Figure 18.2) over the collection port. **T 1 T 6b**

2. **Time for Generation of CO_2.** Place about 5 g of marble chips, $CaCO_3$, in the generator and cover with 50 mL of water. Extend the thistle tube *below the water level* in the generator and add through it 10 mL of 6 M HCl. Record the time required to fill the test tube. **T 5**

3. **Repeat the Generation of CO_2.** Using a mortar and pestle, crush about 5 g of marble chips. Place the crushed marble chips in the generator and repeat the experiment. Compare the results of the two tests.

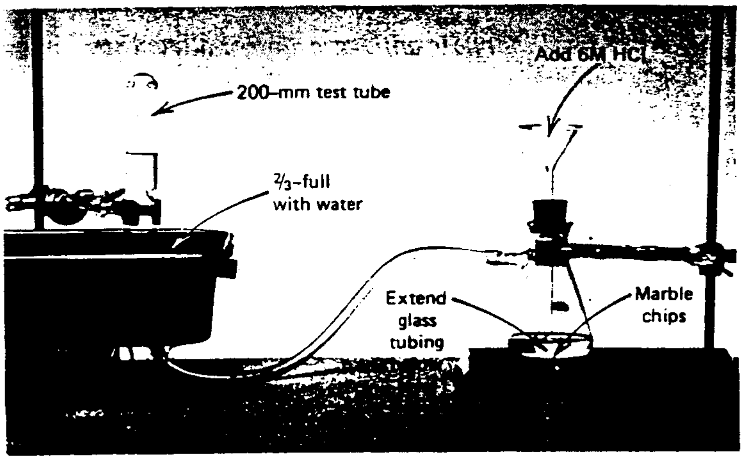


Figure 22.2. *Apparatus for the collection of CO_2 gas*

C. Temperature of the Reaction: Hydrochloric Acid-Sodium Thiosulfate Reaction System

Ask your instructor to determine if *both* Parts C and D are to be completed. You should perform the experiment with a partner; as one student pours the test solutions together, the other notes the time.

The oxidation-reduction reaction that occurs between hydrochloric acid and sodium thiosulfate, $Na_2S_2O_3$, produces insoluble sulfur as a product.

$$2\,HCl(aq) + Na_2S_2O_3\,(aq) \rightarrow S(s) + SO_2(g) + 2\,NaCl(aq) + H_2O \qquad (22.1)$$

The time required for the cloudiness of sulfur to appear is a measure of the reaction rate. Measure each volume (±0.1 mL) with separate graduated pipets.

1. **Prepare the Solutions.** Pipet 2 mL of 0.1 M $Na_2S_2O_3$ into a set of three 150-mm, *clean* test tubes. Into a second set of three 150-mm test tubes pipet 2 mL of 0.1 M HCl. T 2 T 14a T 14b

2. **Time for Reaction at the "Lower" Temperature.** Place a $Na_2S_2O_3$-HCl pair of test tubes in a salt-ice-water bath until thermal equilibrium is established (approximately 5 minutes). Pour the HCl solution into the $Na_2S_2O_3$ solution, agitate the mixture for several seconds, and return the reaction mixture to the ice bath. Record the time for the cloudiness of the sulfur to appear. Record the temperature (±0.1°C) of the ice bath. T 8

3. **Time for Reaction at the "Higher" Temperature.** Place a second $Na_2S_2O_3$-HCl pair of test tubes in a hot water (≈ 70°C) bath until thermal equilibrium is established (approximately 5 minutes). Pour the HCl solution into the $Na_2S_2O_3$ solution as before, agitate the mixture for several seconds, and return the reaction mixture to the hot water bath. Record the time for the cloudiness of the sulfur to appear. Record the temperature (±0.1°C) of the bath. T 11d T 8

4. **Time for Reaction at "Room" Temperature.** Combine the remaining set of $Na_2S_2O_3$-HCl test solutions at room temperature and proceed as in Parts C.2 and C.3. Record the appropriate data. **T 8**

5. **Plot the Data.** Construct a temperature (y-axis) vs. time (x-axis) on linear graph paper. Plot the three data points. Have the instructor approve your graph. **App C**

D. Temperature of the Reaction: Oxalic Acid-Potassium Permanganate Reaction System

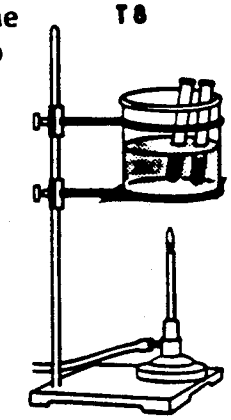
The reaction rate for the oxidation-reduction reaction between oxalic acid, $H_2C_2O_4$, and potassium permanganate, $KMnO_4$, is measured by recording the time elapsed for the (purple) color of the permanganate ion, MnO_4^-, to disappear.

$$5\,H_2C_2O_4(aq) + 2\,KMnO_4(aq) + 3\,H_2SO_4(aq) \rightarrow 10\,CO_2(g) + 2\,MnSO_4(aq) + K_2SO_4(aq) + 8\,H_2O \qquad (22.2)$$

Measure the volume (±0.1 mL) of each solution with separate, *clean* graduated pipets. As one student pours the test solutions together, the other notes the time.

1. **Prepare the Solutions.** Into a 150-mm test tube pipet 1 mL of 0.01 M $KMnO_4$ (in 3 M H_2SO_4) and 4 mL of 3 M H_2SO_4. Pipet 5 mL of 0.33 M $H_2C_2O_4$ in a second, clean 150-mm test tube. **T 14a T 14b T 2**

2. **Time for Reaction at "Room" Temperature.** Pour the $H_2C_2O_4$ solution into the $KMnO_4$ solution. Record the time for the purple color of the permanganate ion to disappear. Record "room" temperature. **T 8**



3. **Time for Reaction at "Higher" Temperature.** Place a second $KMnO_4$-$H_2C_2O_4$ pair of test tubes in a warm water (≈ 40°C) bath until thermal equilibrium is established (approximately 5 minutes). Pour the $H_2C_2O_4$ solution into the $KMnO_4$ solution as before, agitate the mixture for several seconds, and return the reaction mixture to the warm water bath. Record the time for the disappearance of the purple color. Record the temperature (±0.1°C) of the bath. **T 11d T 8**

4. **Time for Reaction at the "Highest" Temperature.** Repeat Parts D.1 and D.3, but increase the temperature of the bath to about 80°C. Record the appropriate data. **T 11d T 8**

5. **Plot the Data.** Construct a temperature (y-axis) vs. time (x-axis) on linear graph paper. Plot the three data points. Have the instructor approve your graph. **App C**

E. Presence of a Catalyst

Hydrogen peroxide is relatively stable, but readily decomposes in the presence of a catalyst.

1. **Add a Catalyst.** Place 5 mL of a 3% H_2O_2 solution in a *clean* 200-mm test tube. Add a "pinch" of MnO_2 to the solution and observe. Note its instability. **T 2 T 14a**

F. Concentration of Reactants: Magnesium-Hydrochloric Acid System

Ask your instructor for advice in completing *both* Parts F and G.

1. **Prepare the Reactants.** Pipet 1 mL of 6 M HCl, 3 M HCl, 1 M HCl, and 0.1 M HCl successively into separate, clean 75-mm test tubes.[2] Weigh (±0.01 g) separately (for identification) four 5-mm strips of *polished* magnesium. **T 14a T 14b T 5**

2. **Time for the Completion of the Reaction.** Add a preweighed magnesium strip to the 6 M HCl solution and record the time for all traces of the magnesium strip to disappear. Repeat the experiment with the 3 M HCl, 1 M HCl, and 0.1 M HCl solutions.

3. **Plot the Data.** Plot $\left(\frac{\text{mol HCl}}{\text{mol Mg}}\right)$ (y-axis) vs. time in seconds (x-axis) for the four tests on linear graph paper. Have the instructor approve your graph. **App C**

G. Concentration of Reactants: Iodic Acid-Sulfurous Acid System

A series of interrelated oxidation-reduction reactions occur between iodic acid, HIO_3, and sulfurous acid, H_2SO_3, that ultimately lead to the formation of iodine, I_2, and sulfuric acid as the final products.

$$2\,HIO_3(aq) + 5\,H_2SO_3(aq) \rightarrow I_2(aq) + 5\,H_2SO_4(aq) + H_2O \text{ (net equation)} \qquad (22.3)$$

The free iodine, appears *only* after all of the sulfurous acid is consumed in the reaction. Once the iodine forms, its presence is detected by its reaction with starch, forming a deep-blue complex.

$$I_2(aq) + \text{starch} \rightarrow I_2\bullet\text{starch}(aq) \qquad (22.4)$$

1. **Prepare the Test Solutions.** Prepare the test solutions according to Table 22.1. Measure the volumes (±0.02 mL) of the 0.1 M HIO_3, starch, and 0.1 M H_2SO_3 solutions with *clean*, 10-mL graduated cylinders[3] and the volume of the water (±0.5 mL) with a 50-mL graduated cylinder. **T 14a T 2**

Table 22.1. Reactant Concentration and Reaction Rate

	Beaker (A), 150-mL Beaker			Beaker (B), 150-mL Beaker	
Trial	0.1 M HIO_3	starch	H_2O	0.1 M H_2SO_3	H_2O
1	3.0 mL	5.0 mL	92.0 mL	10.0 mL	90.0 mL
2	6.0 mL	5.0 mL	89.0 mL	10.0 mL	90.0 mL
3	10.0 mL	5.0 mL	85.0 mL	10.0 mL	90.0 mL

[2]Remember to properly rinse the pipet with the appropriate solution before dispensing it into the test tube.

[3]Be careful! If only one 10-mL graduated cylinder is available and it is *not* cleaned between volume measurements, the HIO_3 and H_2SO_3 solutions will react before the time measurement begins. This is a significant error in technique.

FACTORS AFFECTING REACTION RATES

Date_______ Lab Sec. __ Name____________________________________ Desk No.___

1. List five factors that can affect the rate of a chemical reaction.

   a.

   b.

   c.

   d.

   e.

2. Identify the factor from Question 1 that accounts for the following chemical observations.

   a. Enzymes accelerate certain biochemical reactions, but are not consumed.

   b. Wooden sticks and twigs burn more rapidly than logs.

   c. Rubber tires deteriorate more rapidly in smog-laden areas than in urban areas.

   d. Meat spoils more rapidly when it is not refrigerated.

   e. Gold and silver are used for jewelry, but iron and sodium (also shiny, lustrous metals) are not.

   f. Aquatic life grows more rapidly near the cooling water outfall from a power plant.

3. Assuming that the rate of a chemical reactions doubles for every 10°C temperature increase, by what factor would a chemical reaction increase if the temperature were increased over a 40°C range?

4. a. What is the proper labeling for the y-axis in Part G?

b. What is an appropriate title for the graph prepared in Part G?

5. Reactions in Parts B, C, D, F, and G are timed. Indicate the "signal" to stop timing in each reaction.

Part B.______________________________

Part C. ______________________________

Part D. ______________________________

Part F.______________________________

Part G. ______________________________

6. Techniques.
   a. What is the subtitle of Technique 11d? Where is this technique used in this experiment?

   b. Technique 14b discusses the pipetting of a solution. How is a pipet properly prepared before dispensing a solution to be used for the experiment?

   c. How is a solution properly delivered from a pipet?

REPORT SHEET-EXPERIMENT 22

FACTORS AFFECTING REACTION RATES

Date_______ Lab Sec. ___ Name____________________________________ Desk No.____

A. Nature of the Reactants

1. List the concentrations of __________ acid in order of decreasing reaction rate with magnesium.

   ________________, ________________, ________________, ________________

   Conclusion.

2. List the metals in order of decreasing rate with 6 M HCl.

   ________________, ________________, ________________

   Conclusion.

B. Surface Area of Reactants

1. Time to collect CO_2 from marble chips (seconds)____________________________

2. Time to collect CO_2 from crushed marble chips (seconds)______________________

3. How does the surface area of the marble chips affect the reaction rate?

C. Temperature of the Reaction: Hydrochloric Acid-Sodium Thiosulfate Reaction System

1. Time for sulfur to appear — Temperature of the reaction

Time for sulfur to appear	Temperature of the reaction
__________ seconds	__________ °C
__________ seconds	__________ °C
__________ seconds	__________ °C

2. Plot the temperature (y-axis) vs. time (x-axis) for the three trials.
   Instructor's approval of graph.____

3. How does temperature affect reaction rate?

4. From your graph, estimate the temperature at which the appearance of sulfur should occur in 60 seconds. Assume no changes in concentration.

D. Temperature of the Reaction: Oxalic Acid-Potassium Permanganate Reaction System

1. Time for permanganate to disappear — Temperature of the reaction

Time for permanganate to disappear	Temperature of the reaction
____________ seconds	____________ °C
____________ seconds	____________ °C
____________ seconds	____________ °C

2. Plot the temperature (y-axis) vs. time (x-axis) for the three trials.
   Instructor's approval of graph. ________________________________

3. How does temperature affect reaction rate?

4. From you graph, estimate the time for the disappearance of the purple permanganate ion at 65°C. Assume no changes in concentration.

E. Presence of a Catalyst

1. What effect does the MnO_2 catalyst have on the rate of evolution of O_2 gas?

2. Write a balanced equation for the decomposition of H_2O_2.

F. Concentration of Reactants: Magnesium-Hydrochloric Acid System

M HCl	mol HCl	mass Mg	mol Mg	$\frac{\text{mol HCl}}{\text{mol Mg}}$	time (sec)
6.0					
3.0					
1.0					
0.10					

1. Plot $\left(\frac{\text{mol HCl}}{\text{mol Mg}}\right)$ (y-axis) vs. time in seconds (x-axis) on linear graph paper.
   Instructor's approval of graph. ________________________________

2. How does a change in the concentration of HCl affect the time required for a known mass of magnesium to react?

3. From your graph, predict the time, in seconds, for 2 mg of Mg to react in 1 mL of 2.0 M HCl.

G. Concentration of Reactants: Iodic Acid-Sulfurous Acid System

Molarity of HIO_3 ________________ Molarity of H_2SO_3 solution ________________

Trial	mL HIO_3	$[HIO_3]_i$[6]	mL H_2SO_3	$[H_2SO_3]_i$	time (sec)
1					
2					
3					

1. Plot $[HIO_3]_i$ (y-axis) vs. time in seconds (x-axis) on linear graph paper.
   Instructor's approval of graph. ________________

2. How does a change in the concentration of HIO_3 affect the time required for the appearance of the deep-blue I_2•starch complex?

3. Estimate the time, in seconds, for the deep-blue I_2•starch complex to form when 8.5 mL of 0.1 M HIO_3 are used. Assume all other conditions remain constant.

[6] See footnote 5.