# Factors Affecting Reaction Rates Experiment

## Assignment & Mini-Report Guidelines

**READING** Experiment – Lab Manual Pages / Handout

*Chemistry*, 5th or 6th ed. by Silberberg – Ch.16 Introduction and Sections 16.1-16.2

**PRE-LAB** *Begin the prelab on a new page of your laboratory notebook.* ***ALL elements of the pre-lab MUST be completed before an experiment is started.*** *The COPY page from your notebook will be collected as you enter the lab. The original pages must stay in your notebook.*

### Heading

- Title of experiment and number, your name, the dates of the experiment.

### Purpose

- Briefly, but specifically explain the purpose of each experiment individually (parts A-D).

### Data Tables

- Starting on a new page, prepare appropriate data tables in the REPORT section of your notebook. These will be checked off *before* the experiment begins. However, DO NOT remove the COPY pages before you perform the experiment.

**MINI LAB REPORT** *Begin the lab report section on a new page of the lab notebook. All of your work for this experiment, including graphs will be done in the lab notebook.*

### Heading

- Title of experiment and number, your name, the dates of the experiment.

### Data / Experimental Analysis

- Data Tables – Record all data into these tables in real-time, meaning as you make the observations.
- ***In your lab notebook***, answer questions and provide calculations and analyses as directed in the lab handout for each part (A-D).

### Post-Laboratory Questions

- If assigned, answer the post-lab questions **directly on the handout** pages and submit with the COPY pages of the mini-report.

# Experiment: Factors Affecting Reaction Rates

**Experiment Background:**

The rate of a chemical reaction is the relationship in the change of the amount of a reactant or product with respect to time.

For the general reaction, 2 A + B → C + D, one way to describe the rate of the reaction would be the change (increase) in the concentration of D with respect to time.

Another way to express it would be the rate of change (decrease) of A with respect to time. These two expressions of the rate can be related through a negative sign and a fraction of ½ based on A disappearing twice as fast as D.

$$\text{rate} = \frac{\Delta[D]}{\Delta t} = \frac{-1}{2}\frac{\Delta[A]}{\Delta t}$$

There are five factors that have a significant influence on the rates of chemical reactions.

1) Nature of the reactants
    - Compounds with high energy bonds or elements with unstable electron configurations are prone to react (Their reactions have a lower activation energy, Ea).
2) Concentration
    - High concentrations lead to more collisions.
3) Surface area
    - For reactions involving the surface of a solid, the greater the surface area, the more collisions that can occur.
4) Temperature
    - The higher the temperature, the greater the kinetic energy of molecules, leading to more effective collisions.
5) Catalysts
    - Catalysts increase the rate by lowering the activation energy (Ea) of a reaction.
    - A catalyst is not changed by the overall reaction.

**System A: Presence of a Catalyst**

*Background:*

Hydrogen peroxide is reasonably stable at room temperature. It degrades very slowly in the dark, faster when exposed to light (half-life on the order of days), and very rapidly with the catalyst, $MnO_2$ (observable in the lab). Hydrogen peroxide degraded to water and oxygen gas under all of these conditions.

***System A (continued)***

*Procedure:*

1) Place approximately 2 mL of 3% $H_2O_2$ in a small test tube.
2) Observe.
3) Add 1-2 crystals of $MnO_2$ (or powder equivalent to about 2 grains of sand to the test tube).
4) Observe.

*Analysis & Report:*

1) Comment on the reaction rate with and without the catalyst.
2) Write the complete, balanced equation for the reaction, including state symbols and the proper notation of the catalyst.

**System B: Nature of Reactants – Acids and Metals**

*Background:*

Some acids react more rapidly that others with various metals due in part to the level of dissociation of the hydrogen ions from the acid molecule. Some metals are also more easily oxidized by acids, and some are not at all. In this system, you will study a few metals and acids qualitatively.

*Procedure:*

1) Variation of the acid.
   A) Fill a set of four small test tubes with the following acids: 3 *M* $H_2SO_4$, 6 *M* HCl, 6 *M* $CH_3COOH$ and 6 *M* $H_3PO_4$ to a height of approximately 1.5 cm.
   B) Submerge a 1-cm strip of magnesium into each test tube.
   C) Observe the relative reaction rates qualitatively and record.

2) Variation of the metal.
   A) Fill a set of three small test tubes with 6 M HCl to a height of approximately 1.5 cm.
   B) Submerge a 1-cm strip of copper, magnesium, and zinc into separate test tubes.
   C) Observe the relative reaction rates qualitatively and record.
   D) Recover, clean, and replace any unreacted metal.

*Analysis & Report:*

1) List the acids in order of decreasing reaction rate with magnesium.
2) List the metals in order of decreasing reaction rate with HCl.
3) Write the balanced chemical equation, including state symbols for the reaction of magnesium metal with HCl, $H_2SO_4$, and $H_3PO_4$.

**System C: Concentration and Reaction Rates: Magnesium in HCl solution**

*Background:*

In part A, you observed the reaction of HCl with Mg. Now, you will observe the effect of the HCl concentration on the rate of the reaction.

*Procedure:*

1) Pipet 1ml of 6 *M* HCl, 4 *M* HCl, 3 *M* HCl, and 2 *M* HCl into four separate labeled small test tubes.
2) Cut four 1-cm Mg strips, polish with steel wool, wipe clean, and mass. Keep track of the mass of each strip.
3) Place one of the strips into the 6 *M* HCl solution and start time.
4) When all traces of the magnesium strip have reacted, record the end time.
5) Repeat the reaction for the other three solutions.

*Analysis & Report:*

1) Calculate the moles of HCl, moles of Mg and the $\frac{\text{mol HCl}}{\text{mol Mg}}$ ratio for each trial.
2) Plot $\frac{\text{mol HCl}}{\text{mol Mg}}$ (y-axis) vs. time in seconds (x-axis).
3) Comment on the effect of the HCl concentration on the reaction time.

## System D: Temperature and the Formation of Sulfur

*Background:*

In acidic solution, sodium thiosulfate undergoes a disproportionation reaction to generate elemental sulfur and sulfur dioxide gas by the following reaction:

$$2\ HCl_{(aq)} + Na_2S_2O_{3(aq)} \rightarrow S_{(s)} + SO_{2(g)} + 2\ NaCl_{(aq)} + H_2O_{(l)}$$

The visible production of yellow, elemental sulfur can serve as a consistent indicator of the reaction rate.

*Procedure:*

1) Prepare the solutions.
   A) Pipet 2 mL of 0.1 *M* $Na_2S_2O_3$ into each of three 150-mm (medium) test tubes.
   B) Pipet 2 mL of 0.1 *M* HCl into a second set of three 150-mm (medium) test tubes.

2) Reaction at low temperature, room temperature, and high temperature.
   A) Prepare a small ice water bath in an Erlenmeyer flask.
   B) Place a pair of thiosulfate/HCl test tubes into the bath and allow to stand for at least five minutes to reach thermal equilibrium.
   C) Record the temperature of the bath.
   D) Pour the HCl solution into the thiosulfate solution, start time, agitate for several seconds, and return to the bath.
   E) Stop time when the cloudiness from the sulfur appears, and record the time elapsed.
   F) Prepare a water bath in the range of 50-60ºC.
   G) Repeat steps B-E with this water bath.
   H) Prepare a room temperature water bath.
   I) Repeat steps B-E with this water bath.

*Analysis & Report:*

1) Plot the temperature (y-axis) vs. the time (x-axis) for the three trials.
2) Based on the graph, briefly discuss the effect of temperature on reaction rate.
3) From the graph, estimate the temperature at with the reaction should take 15s for sulfur to appear.

## System E: Alternate Temperature Experiment

This experiment may be assigned as an alternate to system B. Please consult with the instructor.