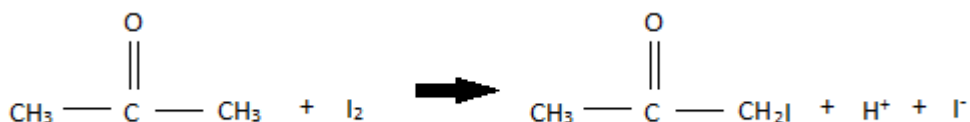


## Rates of Reaction: Iodination of Acetone

### Introduction:

The rate at which a chemical reaction occurs depends on several factors: the nature of the reaction, the concentrations of the reactants, the temperature, and the presence of possible catalysts. In this experiment you will study the kinetics of the reaction between iodine and acetone in acid solution:



For this reaction, you will determine the order of the reaction with respect to acetone and HCl and find a value for the rate constant,  $k$ . Since the concentrations of acetone and HCl are much higher than that of  $\text{I}_2$ , the concentrations of acetone and HCl will change very little. Thus the rate will be determined by the time needed for iodine to be used up. Iodine has color so you can easily follow changes in iodine concentration visually. The equation,  $\text{rate} = k(\text{A})^m(\text{H}^+)^n(\text{I}_2)^p$ , can be simplified to  $\text{rate} = k[\text{I}_2]/t$  since the values for acetone and HCl essentially remain constant during the course of any run.

### Purpose:

The purpose of this reaction is to determine the orders for the reactants, the rate expression, and the rate constant for the reaction between iodine and acetone.

### Equipment/Materials:

4.0 M acetone solution	125 mL Erlenmeyer flasks
1.0 M HCl solution	10 mL graduated cylinders
0.0050 M iodine solution	watch or other timing device
100 mL beakers	watch glass covers for beakers

### Safety:

- Always wear an apron and goggles in the lab
- Acetone is flammable. There should be no open flames in the room.

**Procedure:**

1. Fill in the volumes in Data Table 1 for Trials 2 -- 4. Double the volume of only one reagent at a time, and use the water to maintain a constant total volume. (Volumes for Trial 5 will be determined once data on the preceding trials has been collected.)
2. For Trial 1, pipet the appropriate amount of acetone, HCl, and water; the iodine must be added last. (Be sure to use the correct pipet tip for each liquid.)
3. Simultaneously add the iodine and start the stopwatch; swirl the flask at a consistent rate. Stop the stopwatch when the color disappears and record the time for Run 1.
4. Rinse and dry the flask. Repeat the procedure for the remaining Trials and Runs.
5. For Trial 5, choose volumes for each liquid; the total volume must remain constant. If desired, choose volumes that should yield a short reaction time. Collect data for Trial 5.

**Data and Calculations:****I. Reaction Rate Data**

Trial	Volume Acetone	Volume HCl	Volume Iodine	Volume H <sub>2</sub> O	Time 1 <sup>st</sup> Run	Time 2 <sup>nd</sup> Run	Average Time
1	5 mL	5 mL	5 mL	10 mL			
2							
3							
4							
5							

## II. Determination of Orders

$$\text{Rate} = k[\text{acetone}]^m[\text{I}_2]^n[\text{H}^+]^p$$

Trial	[acetone]	[H <sup>+</sup> ]	[I <sub>2</sub> ]	Rate = [I <sub>2</sub> ]/ave. Time
1				
2				
3				
4				
5				

Order of Acetone “m”

Order of Iodine “n”

Order of Hydrogen Ion “p”

The Rate Law for the reaction is: \_\_\_\_\_

### III. Determination of the Rate Constant k

Trial 1

Trial 2

Trial 3

Trial 4

Average Value for k \_\_\_\_\_

### IV. Prediction of Reaction Rate

Use the data from Trial 5 to compare actual and predicted rates of reaction.

$$\text{Rate} = k[\text{acetone}]^m[\text{I}_2]^n[\text{H}^+]^p$$

$$[\text{acetone}] = \text{_____} \quad [\text{I}_2] = \text{_____}$$

$$[\text{H}^+] = \text{_____} \quad k(\text{average}) = \text{_____}$$

$$\text{Predicted Rate} = \text{_____} \quad \text{Experimental Rate} = [\text{I}_2]/t = \text{_____}$$

**Questions:**

1. Why is the concentration of iodine so much lower than the other reactants?
2. How are time and rate related? How are  $1/\text{time}$  and rate related?
3. What does it mean when someone says a reaction is “first order”?
4. In a reaction,  $A + B \rightarrow C$ , it is found that the reaction is first order in terms of A and B. What happens to the rate if the concentrations of A and B are doubled?