Finding a Rate Law

# Introduction

A reaction is never instantaneous, and the time dependence of a reaction is described by a rate law. For example, if A + B 🡪 C, then the rate law will have the form

Rate = k [A]α [B]β

where k is the rate constant, the brackets represent to concentrations of the two reactants, and α and β are the respective rate orders.

The reaction examined in this experiment is between the iodate and bisulfite ions:

2 IO3− + 3 S2O52− + 3 H2O → 2 I− + 6 HSO4−

with a corresponding rate law

Rate = k [IO3−]α [S2O52−]β

Since the concentration of water is a constant 55.5 M, it is incorporated into the rate law constant.

The rate will be measured as the loss of bisulfite ion. The time when the concentration of the reactant becomes zero will be obvious after the first trial run (we want to surprise you). The method of initial rates, as discussed in class, will be used to find the rate orders and rate constant of this reaction.

The reaction is catalyzed by acid so H+ will appear in the rate law even though it is not a reactant or product. Each reaction mixture will have excess IO3−. When the S2O52− is used up, IO3− reacts with I− to give I3−. Each reaction mixture will also contain starch, which reacts with I3− to give a dark blue complex, whose sudden appearance marks the end of the reaction. Since S2O52− is used up in each experiment, it gives an easy way to calculate reaction rate.



# Materials and Methods

0.1 M sodium bisulfite, Na2S2O5

0.1 M potassium iodate, KIO3

5.0 M acetic acid buffer

Starch solution

# Procedure

1. Add the IO3- solution and buffer to one test tube and the other solutions in to a second one. The buffer is used to control the H+ concentration of the solutions, and distilled water is added so that the total volume is the same for all trials.
2. Start the reaction by pouring one test tube's contents into the other, then mix them quickly by pouring them back and forth two or three times. Record the time it takes to use up the bisulfite ion on the *Report Sheet*.
3. Perform each experiment at least twice. If the two times are not close to each other, repeat the experiment until the times are consistent.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Experiment | mL IO3− | mL buffer | mL S2O52− | mL starch | mL water | total volume |
| #1 | 2.0 | 1.0 | 2.0 | 1.0 | 0 | 6.0 mL |
| #2 | 1.0 | 1.0 | 2.0 | 1.0 | 1.0 | 6.0 mL |
| #3 | 2.0 | 1.0 | 1.0 | 1.0 | 1.0 | 6.0 mL |

The buffer is used to control the H+ concentration of the solutions. The distilled water is added so that the total volume will always be the same.

# Data and Results

1. Use the formula C1V1 = C2V2 to find the concentration of IO3− and HSO3− in each experiment. This will be the concentration *after mixing* but *before reaction*.
2. Calculate the average time taken for each experiment to turn blue.
3. Calculate the initial reaction rate for each experiment using the initial concentration of HSO3− and the time it took to turn blue.
4. Use the results of experiment 1 and 2 and the method of initial rates to find the rate order with respect to the iodate ion.
5. Use the results of experiment 1 and 3 and the method of initial rates to find the rate order with respect to the bisulfite ion.
6. Use the data from experiment 1 to calculate the rate constant. Remember to use the method of dilution to determine the actual concentrations of the two ions at the start of the reaction.