

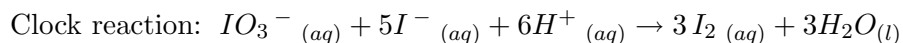
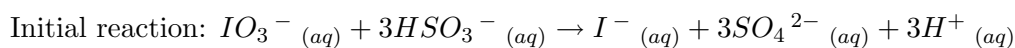
# Chemistry 11 IB      Determining the Order of a Reaction

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## 1 Planning B

### 1.1 Reactions occurring



### 1.2 Materials Required

1. Beakers
2. Graduated Cylinders
3. Stopwatches
4. Water bottle
5. Droppers
6. Chemicals of known concentrations

### 1.3 Procedure

1. Decide on the volume of chemical to be taken for each given chemical and record it.
2. Measure out the decided volume of  $NaHSO_3$  and  $H_2SO_4$  into a beaker, and stir well. Then mix the  $KIO_3$  into the beaker, and start the stop-clock as soon as the  $KIO_3$  is added.
3. As soon as the solution turns bluish-black, stop the timer. The average rate for the reaction can be calculated with the formula

$$Rate = \frac{\Delta [HSO_4^-]}{\Delta t}$$

4. Repeat the above 3 steps several times but vary the volume of *one* of the chemicals and replace the lost volume with water each time. Record the new volumes of the chemicals and the resulting rates for each run.

5. By observing the multiplication factor of a reactant and the corresponding multiplication factor of the rate, estimate the order for that reactant. Repeat the same procedure for the other two reactants, and then determine the order for all three rate-determining reactants.
6. Find the sum of the individual orders to obtain the overall order for the reaction.

## 2 Data Collection

*Attached*

## 3 Data Analysis

*Attached*

## 4 Evaluation

From analyzing our data, the following exponents were obtained:

Reactant	Exponent
$IO_3^-$	0.894
$HSO_3^-$	1.25
$H^+$	0.831

The overall reaction order for this given reaction is the sum of the exponents for each reactant ( $x + y + z$ ):

$$\text{Overall order} = 0.894 + 1.25 + 0.831 = 2.98$$

It would thus be possible to say that the overall reaction is close to being a third-order reaction.

Although the accepted values have to be determined by means of experimentation, it was noted by the teacher that the exponent of each reactant should be 1. Most of our exponents are close to being 1, but our overall order is almost 3.

Also, we were only calculating and comparing the *average rate* for the reaction and not the *initial rate*. We would thus neither be able to write the rate law for the reaction nor will be able to determine the value of the rate constant  $k$ .

## 4.1 Errors

There are a number of sources for error in our experiment:

1. Measurement errors in apparatus, such as the balance and the graduated cylinder.
2. Failure to start and stop the stop-clock at the precise moment.
3. Errors in perception — weak solutions barely turned blue, and this change in colour was difficult to spot.

## 5 Conclusion

Although it wasn't possible to write out the actual rate law for this particular reaction (initial rates weren't found), we were able to estimate the overall order of the reaction to a close approximation. Since the rate law couldn't be written, the correct value for the rate constant  $k$  also couldn't be determined. However, the method was an easy one to determine just the order.

A problem in our procedures was that we were having only a few sets of accurate data to base our results on. If the stop-clock was stopped a little earlier or a little later, the results got skewed. Also, we depended on a clock reaction to tell us whether or not the original reaction was over. We were thus assuming that the concentration of the reactant being observed is zero when the stop-clock is stopped.

It would have been better to obtain the concentrations of our reactants at different intervals of time, as opposed to just one interval. Having several data points for the concentration will help us determine the rate (initial or average) more accurately by means of a graph, as is done when determining the integrated rate law for a reaction.

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