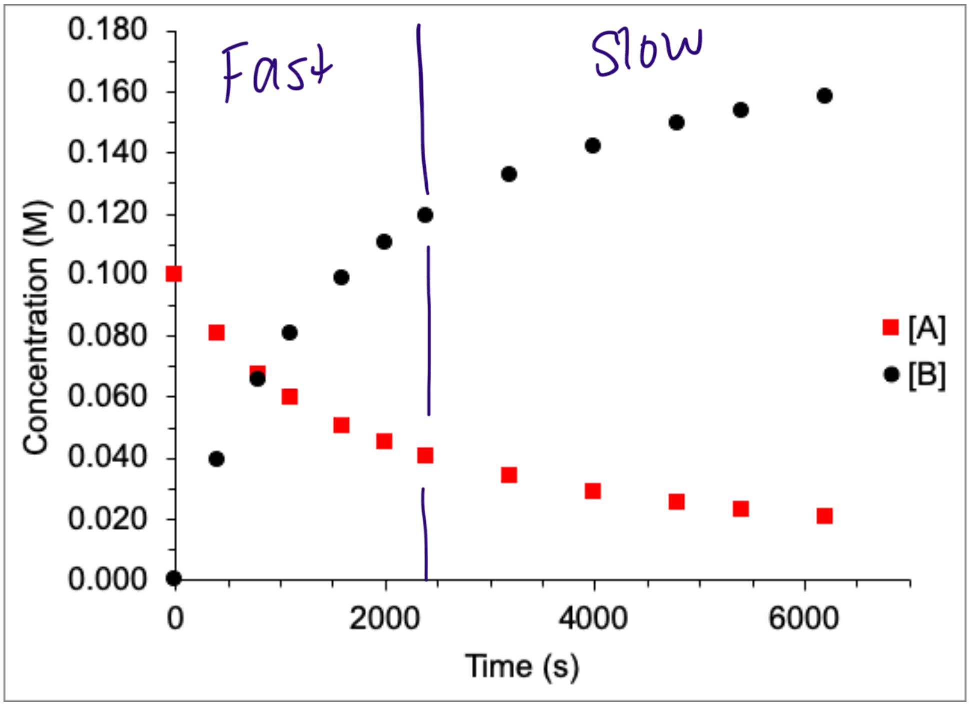
**CHEM 1032**

# Rates of Reaction

In order to better understand equilibrium, we need to understand the rate of reaction. Using the material we learned in lecture last week and your textbook, work with your group to determine how we can learn about speed of a reaction from one plot:




Look at time zero, what is the reactant? What is the product? How do you know?

[A] is high to start so it is a reactant

[B] is low/zero to start, so it is a product

Let's see if we can figure out the stoichiometry of the reaction.

Using A (squares) calculate the rate from 0 to 2000 s.

$$-\frac{\Delta[A]}{\Delta t} = -\frac{(0.045 M - 0.100 M)}{(2000 s - 0 s)} = -\frac{(-0.055 M)}{(2000 s)} = 2.75 \times 10^{-5} M/s$$

Using B (circles) calculate the rate from 0 to 2000 s.

$$\frac{\Delta[B]}{\Delta t} = \frac{(0.110 M - 0.00 M)}{(2000 s - 0 s)} = 5.5 \times 10^{-5} M/s$$

What is the stoichiometry between A and B? How do you know?

The rate of B is double the rate of A, so as A is used twice as much B is formed. So A : B
1 : 2

If A and B are the only compounds in the reaction, what is the chemical equation? $\left(\frac{5.5\times10^{-5}}{2.75\times10^{-5}}\right) = 2$

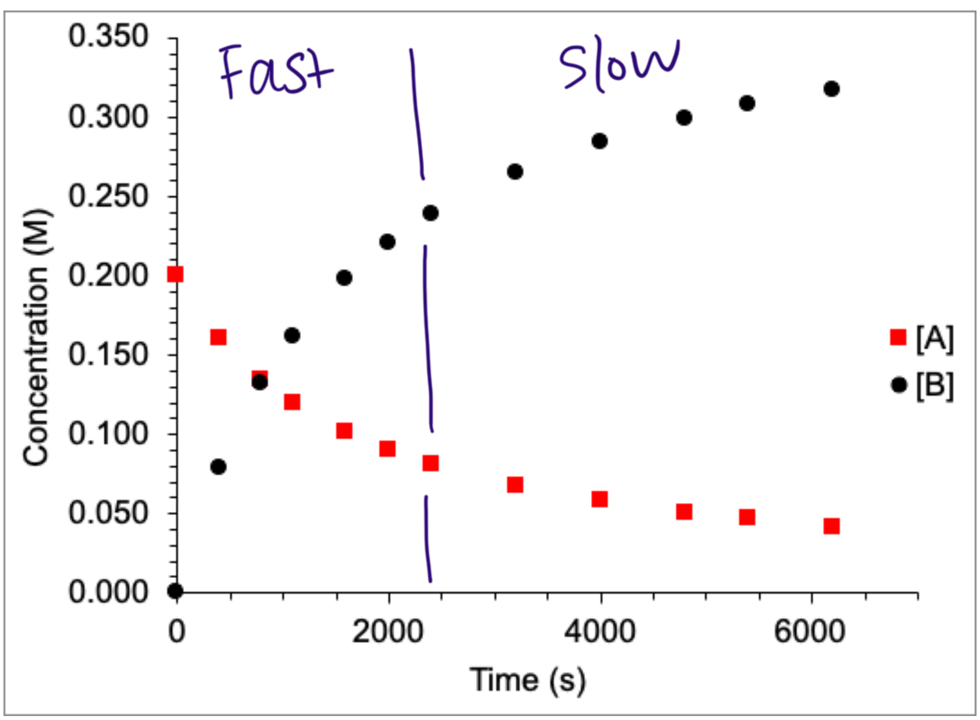
$$A \rightarrow 2B$$

From one plot we figured out a lot of information! But there is more! On the plot label the "fast" and "slow" region of the reaction.

How did you distinguish between the "fast" and "slow" region?

In the fast region, the slope of the line is steeper, more change is occuring. In the slower region the slope is shallower, so the rate of reaction has decreased.

Now let's look at another plot, the same reaction, but double the amount of reactant. On the plot label the "fast" and "slow" region of the reaction. Did they change compared to the first plot?




Regions are in the same place

So doubling the concentration of the reactant, did not change the overall look of the reaction, but did the speed of the reaction change? Using A (squares) calculate the rate from 0 to 2000 s.

$$-\frac{\Delta[A]}{\Delta t} = -\frac{(0.090M - 0.200M)}{(2000s - 0s)} = -\frac{(-0.11M)}{(2000s)} = 5.5\times10^{-5} M/s$$

Now compare the rate of A from plot 1 and from plot 2. How did the initial rate change when the initial concentration of reactant was doubled?

The rate of reaction doubled when the initial concentration doubled.

1st 0.1000 M $2.75\times10^{-5}$ M/s

2nd 0.2000 M $5.5\times10^{-5}$ M/s

$\times 2$

Is the rate of reaction dependent on the concentration of reactant? How do you know? What is the relationship between the two?

Yes, based on this data we see the concentration of the reactant affects the rate of reaction. We can tell this by the analysis above. The rate doubled, when concentration doubled, so they are directly proportional.

Based on these observations, what is the reaction order with respect to A? Why?

Since [reactant] x 2 yielded rate x 2, the reaction order must be 1.

Here, we used the method of initial rates to determine the reaction order. What's another method?

We could have also used the integrated rate laws.

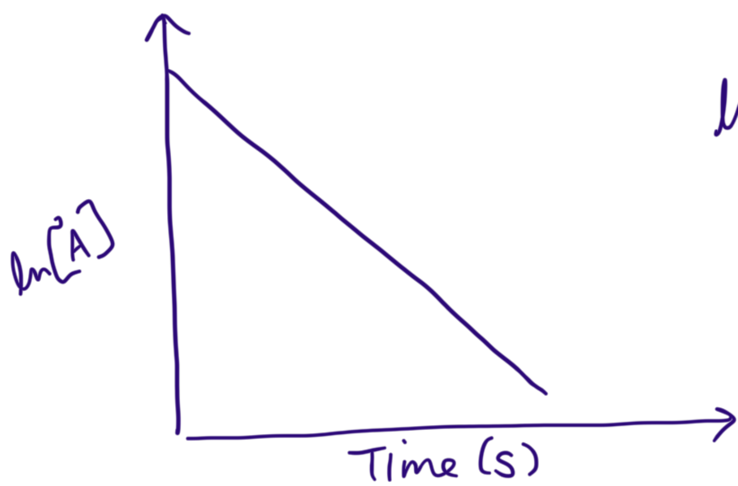
Describe how you would use this alternative method to determine the reaction order. What's the goal?

Here we would plot [A] vs time, ln[A] vs time, and 1/[A] vs time. Which ever plot yielded a straight line would tell us that the data fitted that reaction order.

What would the plot look like for this example? Write the integrated rate law, sketch a plot, and label the axes.

Since this reaction is 1st order ln[A] vs time

$$\ln[A]_t = -kt + \ln[A]_0$$




How can the rate constant be determined from this plot?

Since the line is linear and I know the equation. The slope is useful.

$$-m = k$$