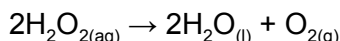


## PRE-LAB EXERCISE FOR EXPERIMENT 5

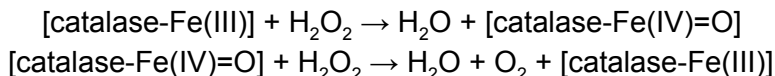
### A KINETIC STUDY: CATALASE

1. The partial order with respect to the peroxide (x) can be found using the equation  $rate = k[H_2O_2]^x [catalase]^y$ . To find it, the following steps must be followed:

-Write down the chemical equation of the reaction involved.



-Find the elementary steps of the chemical reaction involved. In this case, the Fe(III) undergoes oxidation. As a result, it acts as an electron exchanger. The two elementary reactions describe the behaviour of the individual molecules. They have to add up to make the overall reaction. Due to this, some species are produced in one step and consumed in another (reaction intermediate). As demonstrated below, the reaction intermediates are Fe(III) and Fe(IV)=O.



-Determine the rate-determining step. Both elementary reactions are bimolecular. That being said, the first step is the slowest step and, therefore, it is the rate-determining step. The rate-determining step defines the rate of the overall reaction.

-By obtaining the rate-determining step, find x. The partial order with respect to peroxide can be found by determining the order for the slowest step. This order is the order for the overall reaction.

2. Using the Arrhenius formula, the activation energy can be calculated.

$$\ln(k_2/k_1) = Ea/(R)(1/T_1 - 1/T_2)$$

Note: To convert from log<sub>10</sub> to natural logs, you multiply by 2.303. To convert in the other direction, you divide by 2.303. Therefore, the new Arrhenius formula is...

$$\log(k_2/k_1) = Ea/(R(2.303))(1/T_1 - 1/T_2)$$

$$k_2 = 0.63 \text{ mL s}^{-1}$$

$$k_1 = 0.46 \text{ mL s}^{-1}$$

$$R = 8.31 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$T_1 = 273 + 20 = 293 \text{ K}$$

$$T_2 = 273 + 30 = 303 \text{ K}$$

$$\log(0.63/0.46) = Ea/(8.314(2.303))(1/293 - 1/303)$$

$$\log(1.37) = Ea/(19.147)(0.00011264)$$

$$Ea = 23240.31161 \text{ J/mol}$$

3. To find the initial rate of the reaction, find the slope of the data points given. Therefore,  
initial rate = change in pressure/change in time.  
Initial rate =  $(105.86 - 105.30)/(1.00 - 0.80)$   
Initial rate =  $(0.56)/(0.20)$   
Initial rate = 2.8 kPa/s
4. I would expect the plot to pass through the origin. In other words, the volume of oxygen collected or the pressure of oxygen at time zero is zero. It is because, by looking at the chemical equation of the reaction, one can observe the oxygen is one of the products. At time zero, no oxygen is formed.