

### 1. Question 1

The differences in the two data sets are exactly what we would expect. According to the chemical kinetics, the rate of a reaction increases with an increase in the molarity of the reactant. We can evaluate the average rate of change for the first run as:

$$r_1 = \frac{\Delta p_{O_2}}{\Delta t} = \frac{510 \text{ torr} - 0 \text{ torr}}{180.00 \text{ s} - 0.00 \text{ s}} = 2.83 \frac{\text{torr}}{\text{s}}$$

For the second run:

$$r_2 = \frac{\Delta p_{O_2}}{\Delta t} = \frac{408 \text{ torr} - 0 \text{ torr}}{180.00 \text{ s} - 0.00 \text{ s}} = 2.27 \frac{\text{torr}}{\text{s}}$$

The average rate of change for the first run is greater than the second one, as expected, since a greater molarity of the hydrogen peroxide was used in the first trial.

### 2. Question 2

In each case, we're given a total volume of:

$$V_t = 150.0 \text{ mL}$$

The volume of the flask occupied by the hydrogen peroxide solution is:

$$V_{H_2O_2} = 50.0 \text{ mL}$$

The volume occupied by the platinum disc is:

$$V_{Pt} = 1.0 \text{ cm}^3$$

Hence, since ideal gasses occupy the whole free volume of any container, the volume of the oxygen gas produced in each case would be:

$$V_{O_2} = V_t - V_{H_2O_2} - V_{Pt} = 150.0 \text{ mL} - 50.0 \text{ mL} - 1.0 \text{ mL} = 99.0 \text{ mL} = 0.0990 \text{ L}$$

### 3. Question 3

Given the table:

$[\text{H}_2\text{O}_2]_0, \text{M}$	Time, s	$p_{\text{O}_2}, \text{torr}$	$n_{\text{O}_2}, \text{mol}$	$n_{\text{H}_2\text{O}_2}, \text{mol}$	$[\text{H}_2\text{O}_2], \text{M}$
0.882	0.00	0.0			0.882
	30.00	90.0			
	60.00	178.0			
	90.00	263.0			
	120.00	348.0			
	150.00	430.0			
	180.00	510.0			

First, we will demonstrate how the number of moles of oxygen is calculated using the ideal gas law. Since:

$$p_{\text{O}_2} V_{\text{O}_2} = n_{\text{O}_2} RT$$

Then:

$$n_{\text{O}_2} = \frac{p_{\text{O}_2} V_{\text{O}_2}}{RT}$$

Given the temperature:

$$T = 25.00^\circ\text{C} = 298.15 \text{ K}$$

The fixed volume of the oxygen gas:

$$V_{\text{O}_2} = 0.0990 \text{ L}$$

And the universal gas constant:

$$R = 0.08206 \frac{\text{L atm}}{\text{mol K}}$$

For the second data point, we get:

$$n_{\text{O}_2} = \frac{90.0 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} \cdot 0.0990 \text{ L}}{0.08206 \frac{\text{L atm}}{\text{mol K}} \cdot 298.15 \text{ K}} = 4.792 \times 10^{-4} \text{ mol}$$

We proceed with the same calculations for each moles using the Excel worksheet and obtain:

[H <sub>2</sub> O <sub>2</sub> ] <sub>o</sub> , M	Time, s	p <sub>O<sub>2</sub></sub> , torr	n <sub>O<sub>2</sub></sub> , mol	n <sub>H<sub>2</sub>O<sub>2</sub></sub> , mol	[H <sub>2</sub> O <sub>2</sub> ], M
0.882	0.00	0.0	0		0.882
	30.00	90.0	4.792 x 10 <sup>-4</sup>		
	60.00	178.0	9.477 x 10 <sup>-4</sup>		
	90.00	263.0	1.400 x 10 <sup>-3</sup>		
	120.00	348.0	1.853 x 10 <sup>-3</sup>		
	150.00	430.0	2.289 x 10 <sup>-3</sup>		
	180.00	510.0	2.715 x 10 <sup>-3</sup>		

According to the stoichiometry of this reaction, the number of moles of hydrogen peroxide reacted is twice as large:

$$n_{H_2O_2} = 2n_{O_2} = 2 \cdot 4.792 \times 10^{-4} \text{ mol} = 9.584 \times 10^{-4} \text{ mol}$$

We fill out the number of moles using the same logic and Excel:

[H <sub>2</sub> O <sub>2</sub> ] <sub>o</sub> , M	Time, s	p <sub>O<sub>2</sub></sub> , torr	n <sub>O<sub>2</sub></sub> , mol	n <sub>H<sub>2</sub>O<sub>2</sub></sub> , mol	[H <sub>2</sub> O <sub>2</sub> ], M
0.882	0.00	0.0	0	0	0.882
	30.00	90.0	4.792 x 10 <sup>-4</sup>	9.584 x 10 <sup>-4</sup>	
	60.00	178.0	9.477 x 10 <sup>-4</sup>	1.895 x 10 <sup>-3</sup>	
	90.00	263.0	1.400 x 10 <sup>-3</sup>	2.801 x 10 <sup>-3</sup>	
	120.00	348.0	1.853 x 10 <sup>-3</sup>	3.706 x 10 <sup>-3</sup>	
	150.00	430.0	2.289 x 10 <sup>-3</sup>	4.579 x 10 <sup>-3</sup>	
	180.00	510.0	2.715 x 10 <sup>-3</sup>	5.431 x 10 <sup>-3</sup>	

Finally, subtracting the molarity reacted, we obtain the remaining molarity. For the second data point, we get:

$$[H_2O_2]_t = 0.882 \text{ M} - \frac{9.584 \times 10^{-4} \text{ mol}}{0.0500 \text{ L}} = 0.8628 \text{ M}$$

Using Excel:

$[H_2O_2]_0, M$	Time, s	$p_{O_2}, \text{torr}$	$n_{O_2}, \text{mol}$	$n_{H_2O_2}, \text{mol}$	$[H_2O_2], M$
0.882	0.00	0.0	0	0	0.882
	30.00	90.0	$4.792 \times 10^{-4}$	$9.584 \times 10^{-4}$	$8.628 \times 10^{-1}$
	60.00	178.0	$9.477 \times 10^{-4}$	$1.895 \times 10^{-3}$	$8.441 \times 10^{-1}$
	90.00	263.0	$1.400 \times 10^{-3}$	$2.801 \times 10^{-3}$	$8.260 \times 10^{-1}$
	120.00	348.0	$1.853 \times 10^{-3}$	$3.706 \times 10^{-3}$	$8.079 \times 10^{-1}$
	150.00	430.0	$2.289 \times 10^{-3}$	$4.579 \times 10^{-3}$	$7.904 \times 10^{-1}$
	180.00	510.0	$2.715 \times 10^{-3}$	$5.431 \times 10^{-3}$	$7.734 \times 10^{-1}$

We fill out the second data table using the same approach:

$[H_2O_2]_0, M$	Time, s	$p_{O_2}, \text{torr}$	$n_{O_2}, \text{mol}$	$n_{H_2O_2}, \text{mol}$	$[H_2O_2], M$
0.706	0.00	0.0	0	0	0.706
	30.00	72.0	$3.833 \times 10^{-4}$	$7.667 \times 10^{-4}$	$6.907 \times 10^{-1}$
	60.00	142.0	$7.560 \times 10^{-4}$	$1.512 \times 10^{-3}$	$6.758 \times 10^{-1}$
	90.00	211.0	$1.123 \times 10^{-3}$	$2.247 \times 10^{-3}$	$6.611 \times 10^{-1}$
	120.00	278.0	$1.480 \times 10^{-3}$	$2.960 \times 10^{-3}$	$6.468 \times 10^{-1}$
	150.00	344.0	$1.832 \times 10^{-3}$	$3.663 \times 10^{-3}$	$6.327 \times 10^{-1}$
	180.00	408.0	$2.172 \times 10^{-3}$	$4.345 \times 10^{-3}$	$6.191 \times 10^{-1}$

#### 4. Question 4

We evaluate the average rates of decomposition for runs 1 and 2, respectively, as:

$$r_1 = \frac{\Delta[H_2O_2]}{\Delta t} = \frac{0.7734 M - 0.882 M}{180.00 s - 0.00 s} = -6.033 \times 10^{-4} \frac{M}{s}$$

$$r_2 = \frac{\Delta[H_2O_2]}{\Delta t} = \frac{0.6191 M - 0.706 M}{180.00 s - 0.00 s} = -4.828 \times 10^{-4} \frac{M}{s}$$

#### 5. Question 5

Assuming the rate law is:

$$r = k[H_2O_2]^x$$

We define the order as:

$$\frac{r_2}{r_1} = \frac{k[H_2O_2]_2^x}{k[H_2O_2]_1^x}$$

Using the law of initial rates with varying concentrations:

$$\frac{r_2}{r_1} = \left( \frac{[H_2O_2]_2}{[H_2O_2]_1} \right)^x$$

From here:

$$x = \frac{\ln\left(\frac{r_2}{r_1}\right)}{\ln\left(\frac{[H_2O_2]_2}{[H_2O_2]_1}\right)} = \frac{\ln\left(\frac{4.828}{6.033}\right)}{\ln\left(\frac{0.706}{0.882}\right)} = 1.001$$

Hence, the rate law has a form:

$$r = k[H_2O_2]$$

Identify the rate constant (calculated from the rate of consumption of hydrogen peroxide rather than the actual reaction rate) for each experiment:

$$k_1 = \frac{r_1}{[H_2O_2]_1} = \frac{6.033 \times 10^{-4} \frac{M}{s}}{0.882 M} = 6.840 \times 10^{-4} s^{-1}$$
$$k_2 = \frac{r_2}{[H_2O_2]_2} = \frac{4.828 \times 10^{-4} \frac{M}{s}}{0.706 M} = 6.839 \times 10^{-4} s^{-1}$$

Average value:

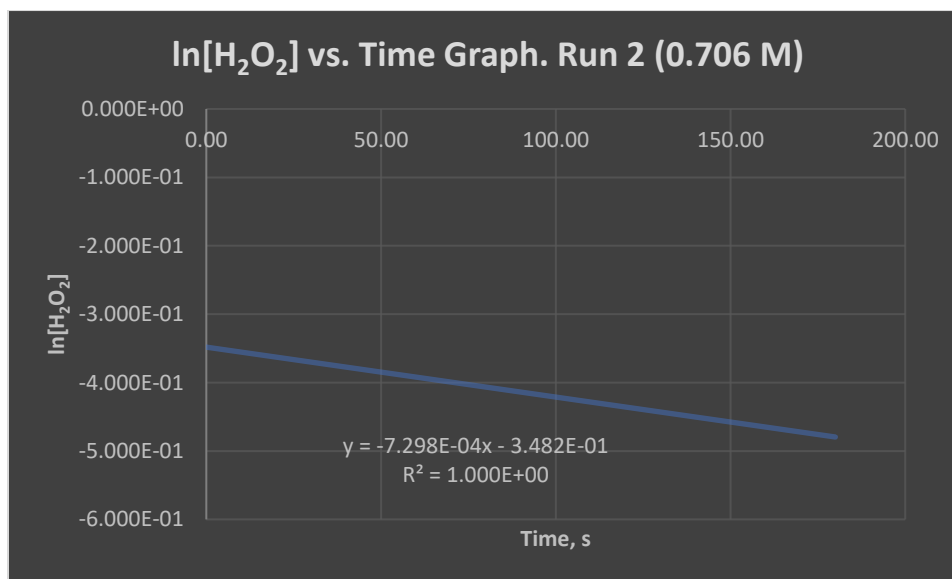
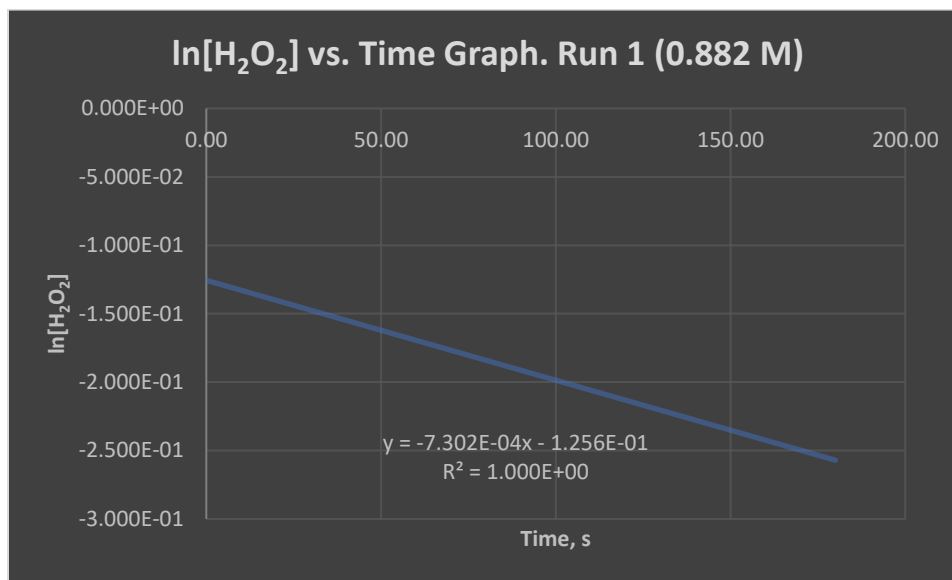
$$k' = \frac{1}{2}(k_1 + k_2) = 6.840 \times 10^{-4} s^{-1}$$

So that the rate law is:

$$r = 6.840 \times 10^{-4} s^{-1} [H_2O_2]$$

6. Question 6

The following graphs were obtained:



7. Question 7

Yes, the graphs are linear when the natural logarithm of the concentration is plotted against time. This indicates a first-order reaction, exactly the same order as we got ( $x = 1.00$ ). In addition, the rate constant is the negative of the slope value in each case. Our approximation provided a  $k$  value of:

$$k' = 6.840 \times 10^{-4} \text{ s}^{-1}$$

Which indicates an insignificant error compared to any of the two slopes:

$$k'' = -m = 7.298 \times 10^{-4} \text{ s}^{-1}$$

This error mainly comes from the approximation that the average rate could be considered as an instantaneous initial rate of the reaction.