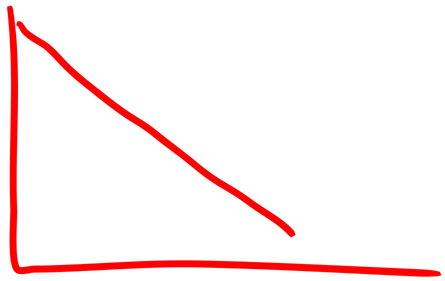
33) to make the first assumption that the reaction is first order



differential rate law: rate $= k[H_2O_2]$

integrated rate law: $\ln[H_2O_2] = -kt + \ln[H_2O_2]_0$

slope: $-k = \frac{\Delta Y}{\Delta X} = \frac{0-(3.00)}{0-(3600)} = -8.3\times10^{-4}\ s^{-1}$

$k = 8.3\times10^{-4}\ s^{-1}$

$t = 4000s$ $\quad [H_2O_2]_0 = 0.500M$

$\ln[H_2O_2] = -kt + \ln[H_2O_2]_0$

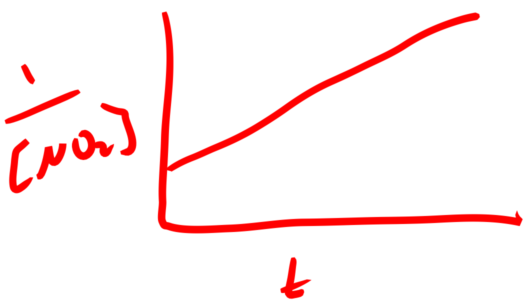
$\ln\frac{[H_2O_2]}{[H_2O_2]_0} = -kt$

$\ln\frac{x}{1.00} = (-8.3\times10^{-4}\ s^{-1})(4000s)$

$\ln x = -3.3$ $\quad x = e^{-3.3}$ $\quad x = 0.037M$

35) assume reaction is first order

not first order



is second order

rate law

$$rate = k[NO_2]^2$$

integrated rate law

$$\frac{1}{[NO_2]} = kt + \frac{1}{[NO_2]_0}$$

$$\text{slope} = k = \frac{\Delta y}{\Delta x} = \frac{(5.75 - 2.00)}{(1.8\times10^4 - 0)} = 2.08\times10^{-4}\ \frac{L}{mol \cdot s}$$

$$\frac{1}{[NO_2]} = kt + \frac{1}{[NO_2]_0}$$

$$\frac{1}{x} = \left(2.08\times10^{-4}\ \frac{L}{mol\cdot s}\right)(2.70\times10^4\ s) + \frac{1}{0.500\ M}$$

$$\frac{1}{x} = 7.62 \qquad x = [NO_2] = 0.131\ M$$

37) a) $[C_2H_5OH]$ vs t linear

zero order in $[C_2H_5OH]$

$$\text{slope} = -k = \frac{\Delta y}{\Delta x} = 4.00 \times 10^{-5} \frac{mol}{L \cdot s}$$

rate law

rate = k

integrated rate law

$$[C_2H_5OH] = -kt + [C_2H_5OH]_0$$

b) $$t_{\frac{1}{2}} = \frac{[A]_0}{2k} = \frac{1.25 \times 10^{-2} \frac{mol}{L}}{2\left(4.00 \times 10^{-5} \frac{mol}{L \cdot s}\right)} = 156\ s$$

c) $$[C_2H_5OH] = -kt + [C_2H_5OH]_0$$

$$0 \frac{mol}{L} = -\left(4.00 \times 10^{-5} \frac{mol}{L \cdot s}\right) t + 1.25 \times 10^{-2} \frac{mol}{L}$$

$$t = \frac{1.25 \times 10^{-2} \frac{mol}{L}}{4.00 \times 10^{-5} \frac{mol}{L \cdot s}} = 313\ s$$