During Class Invention

Integrated Rate Law Part I

Name___________________________

TA Name _______________________

Lab Section #______

1. The reaction:

$$A(g) \rightarrow \text{products}$$

follows simple first order kinetics. When the initial concentration of A is 0.500 M the initial rate of the reaction is determined to be 4.20×10^{-3} M s^{-1}. If the initial concentration of A is tripled, what would be the new initial rate of the reaction?

The rate law for the reaction that follows first order kinetics is

$$\textbf{Rate} = \mathbf{k[A]^1}$$

If [A] is equal to 3[A] then substitute into the rate law,

$$\textbf{Rate} = \mathbf{k[3[A]]^1}$$

The initial rate will triple.

$$\textbf{Rate} = \mathbf{3 \cdot 4.20 \times 10^{-3}\ M\ s^{-1} = 1.26 \times 10^{-2}\ M\ s^{-1}}$$

2. Write the integrated rate law for a reaction that follows simple first order kinetics.

$$\ln\left(\frac{[A]}{[A]_0}\right) = -kt$$

3. The decomposition of H_2O_2 to H_2O and O_2 follows first order kinetics with a rate constant of 0.0410 min^{-1} at a particular temperature.

$$H_2O_{2}(l) \rightarrow 2H_2O(l) + O_{2}(g)$$

Calculate the $[H_2O_2]$ after 10 mins if $[H_2O_2]_0$ is 0.200 M.

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

$$\ln\frac{[H_2O_2]}{0.200\ M} = -(0.0410\ \text{min}^{-1})(10\ \text{mins})$$

$$\ln\frac{[H_2O_2]}{0.200\ M} = -0.410$$

$$e^{\left(\ln\frac{[H_2O_2]}{0.200\ M}\right)} = e^{-0.410}$$

$$\frac{[H_2O_2]}{0.200\ M} = 0.6637$$

$$[H_2O_2] = 0.6637 \cdot (0.200\ M)$$

$$[H_2O_2] = 0.133\ M$$

4. The decomposition of N_2O_5 to O_2 and NO_2 follows first order kinetics. If a sample at 25 ˚C with the initial concentration of N_2O_5 of 1.25 x 10^{-3} M falls to 1.02 x 10^{-3} M in 100. minutes, calculate the rate constant for the reaction.

$$\ln \frac{[N_2O_5]}{[N_2O_5]_0} = -kt$$

$$\ln \frac{[1.02 \text{ x } 10^{-3}\text{ M}]}{[1.25 \text{ x } 10^{-3}\text{ M}]_0} = -k(100\text{ min})$$

$$\ln 0.816 = -k(100\text{ min})$$

$$-0.203 = -k(100\text{ min})$$

$$\frac{-0.203}{100\text{ min}} = -k$$

$$2.03 \text{ x } 10^{-3}\text{ min}^{-1} = k$$

5. Show how a plot of *ln*[concentration] versus time can provide the rate constant for a reaction which follows simple first order kinetics.

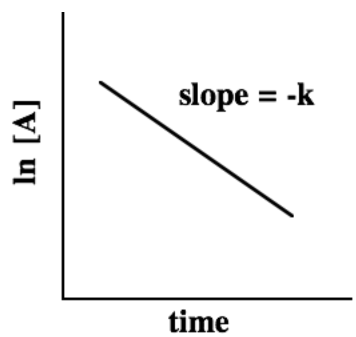
$$\ln\left(\frac{[A]}{[A]_0}\right) = -kt$$

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

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

this equation fits the general equation for a line, $y = mx + b$, where

$y = \ln[A]$: $m = -k$: $x = t$

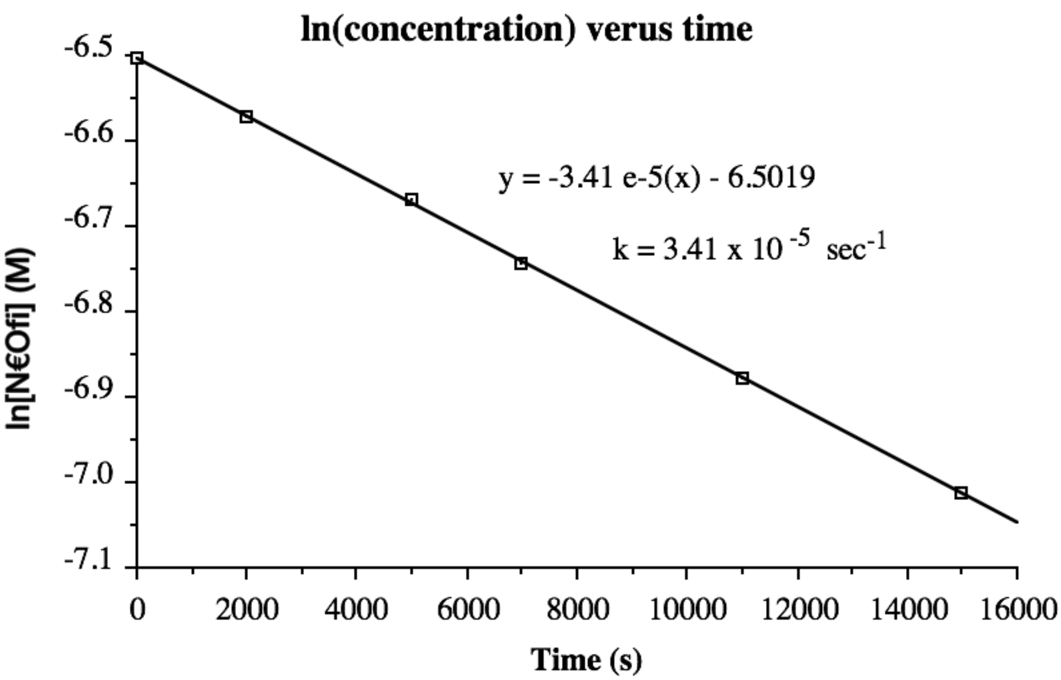


6. Using the following data, establish that the decomposition N_2O_5 according to the reaction,

$$2N_2O_{5(g)} \rightarrow 2NO_{2(g)} + O_{2(g)}$$

follows first order kinetics. Determine the rate constant for the reaction.

Time (sec)	$[N_2O_5]$ (M)	$\ln[NO_2]$
0	1.50 x 10^{-3}	**-6.50**
2000	1.40 x 10^{-3}	**-6.57**
5000	1.27 x 10^{-3}	**-6.67**
7000	1.18 x 10^{-3}	**-6.74**
11000	1.03 x 10^{-3}	**-6.88**
15000	9.00 x 10^{-4}	**-7.01**


ln(concentration) verus time
y = -3.41 e-5(x) - 6.5019
k = 3.41 x 10 -5 sec-1
ln[N€Ofi] (M)
-6.5
-6.6
-6.7
-6.8
-6.9
-7.0
-7.1
0
2000
4000
6000
8000
10000
12000
14000
16000
Time (s)