During Class Invention

Integrated Rate Law Part II

Name___________________________

TA Name _______________________

Lab Section #______

1. The reaction:

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

follows simple second order kinetics. When the initial concentration of B is 0.500 M the initial rate of the reaction is determined to be 8.40×10^{-3} M s^{-1}. If the initial concentration of B is tripled, what would be the new initial rate of the reaction?

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

$$\textbf{Rate} = \mathbf{k[B]^2}$$

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

$$\textbf{Rate} = \mathbf{k[3[B]]^2}$$

The initial rate will increase by a factor of 9.

$$\textbf{Rate} = \mathbf{9 \cdot 8.40 \times 10^{-3}\ M\ s^{-1} = 7.56 \times 10^{-2}\ M\ s^{-1}}$$

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

$$\frac{1}{[A]} - \frac{1}{[A]_0} = kt$$

3a. The decomposition of $NOCl_{(g)}$

$$2NOCl_{(g)} \rightarrow 2NO_{(g)} + Cl_{2(g)}$$

is a second order reaction with a rate constant of 0.0480 $M^{-1}\cdot sec^{-1}$ at 200 °C. In an experiment at 200 °C, the initial concentration of NOCl was 0.400 M. What is the concentration of NOCl after 15.0 min have elapsed?

$$\frac{1}{[NOCl]} - \frac{1}{[NOCl]_0} = kt$$

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

$$\frac{1}{[NOCl]} = 0.0480\ M^{-1}\cdot sec^{-1}\ (15.0\ min)\left(\frac{60\ sec}{1\ min}\right) + \frac{1}{.400\ M}$$

$$\frac{1}{[NOCl]} = 45.7\ M^{-1}$$

$$[NOCl] = 0.0218\ M$$

b. How many minutes will it take for the concentration of $NOCl_{(g)}$ to drop to 0.150 M?

$$\mathbf{\frac{1}{[NOCl]} - \frac{1}{[NOCl]_0} = kt}$$

$$\mathbf{\frac{1}{0.150\ M} - \frac{1}{0.400\ M} = = 0.0480\ M^{-1}{\cdot}sec^{-1}\ (t)}$$

86.8 sec = t or 1.45 minutes

4. Derive a mathematical equation for the half-life for a reaction which follows simple second order kinetics.

We can also arrive at a half–life relationship for a second–order reaction of the type

$$\mathbf{A_{(g)} \rightarrow products}$$

At the half–life the $[A]_{t_{1/2}} = 0.5[A]_0$. Substituting

$$\mathbf{\frac{1}{0.5[A]_0} - \frac{1}{[A]_0} = kt_{1/2}}$$

$$\mathbf{\frac{2}{[A]_0} - \frac{1}{[A]_0} = kt_{1/2}}$$

$$\mathbf{\frac{1}{[A]_0} = kt_{1/2} \quad so \quad t_{1/2} = \frac{1}{k[A]_0}}$$

5. The initial concentration of NOCl, described in 13b above, is 0.400 M. Calculate the half-life for the decomposition reaction.

$$\mathbf{t_{1/2} = \frac{1}{k[A]_0}}$$

$$\mathbf{t_{1/2} = \frac{1}{0.0480\ M^{-1}{\cdot}sec^{-1}{\cdot}(0.400\ M)}}$$

$$\mathbf{t_{1/2} = 52.1\ sec}$$

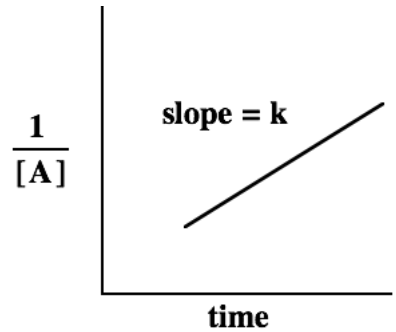
6. Show how a plot of ln [concentration] versus time can provide the rate constant for a reaction which follows simple second order kinetics.

$$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt_{1/2}$$

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

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

$$y = \frac{1}{[A]} : m = k : x = t$$

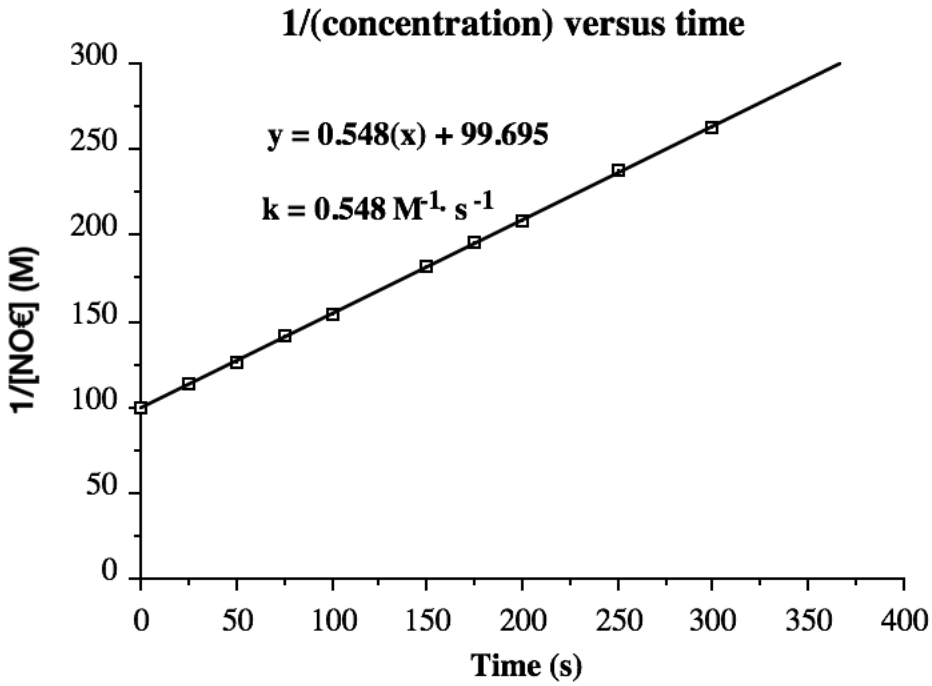


b. Using the following data establish that the decomposition NO_2 according to the reaction,

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

following second order kinetics. Determine the rate constant for the reaction.

Time (sec)	$[NO_2]$ (M)	$\frac{1}{[A]}$
0	0.0100	**100**
25	0.0088	**114**
50	0.0079	**127**
75	0.0071	**141**
100	0.0065	**154**
150	0.0055	**182**
175	0.0051	**196**
200	0.0048	**208**
250	0.0042	**238**
300	0.0038	**263**


1/(concentration) versus time
y = 0.548(x) + 99.695
k = 0.548 M-1. s -1
1/[NO€] (M)
Time (s)
0
50
100
150
200
250
300
350
400