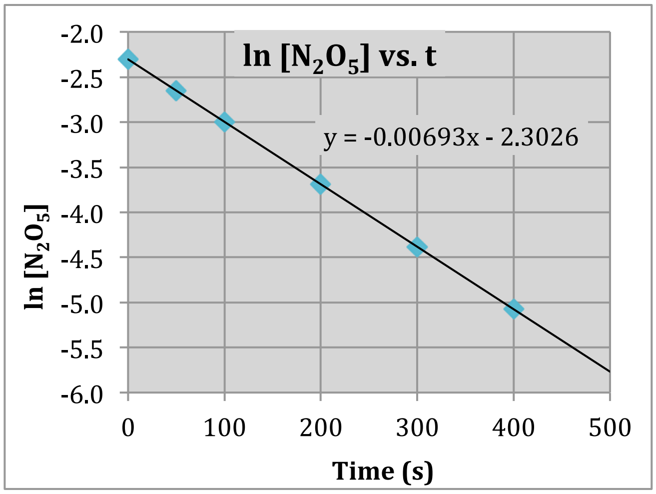
Integrated Rate Law Example Data Sets (Reactant Concentration as a function of time)

1) The decomposition reaction, $2\ N_2O_5\ (g) \rightarrow 4\ NO_2\ (g) + O_2\ (g)$ was studied at constant temperature and the following kinetic data were collected:

Time (s)	$[N_2O_5]$ (M)
0	0.1000
50	0.0707
100	0.0500
200	0.0250
300	0.0125
400	0.00625



Then the data was graphically analyzed. The results were as follows:

- Plot of $[N_2O_5]$ vs. time : not linear
- Plot of $\ln[N_2O_5]$ vs. time: linear as shown in graph→
- Plot of $1/[N_2O_5]$ vs. time: not linear

a. What is the order with respect to N_2O_5? (The answer is also the order of the overall rxn b/c N_2O_5 is the only reactant.)

Since $\ln[N_2O_5]$ vs. time is a straight line, the reaction is 1st order

b. Write the general integrated rate law equation for this order of reaction. Then rearrange the equation so that it is in the form of "y= mx+b."

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

c. The specific equation of the line for the graph is shown on the graph. Use that equation of the line to determine the value of the rate constant for this reaction.

slope $= -0.0069\ s^{-1} = -k$

Thus, $k = 0.0069\ s^{-1}$

d. Determine the half-life of the reaction.

$$t_{1/2} = \frac{0.693}{k} = \frac{0.693}{0.00693\ s^{-1}} = 100\ s$$

e. At what time will $[N_2O_5]$=0.0350 M? (Note: the initial N_2O_5 concentration is given in data chart.)

$$\ln[A] - \ln[A_0] = -kt, \text{ so } t = -\frac{\ln[0.0350] - \ln[0.1000]}{(0.00693\ s^{-1})} = 151\ s$$

f. Determine the value of $[N_2O_5]$ at t=1000 s.

$\ln[A] - \ln[A_0] = -kt$ can be written as $[A] = [A_0]e^{-kt}$;

$$[A] = (0.1000\ M)e^{-(0.0069\ s^{-1})(1000\ s)} = 9.8 \times 10^{-5}\ M$$

g. Suppose the same reaction is done at the same temperature, but one started with a different initial concentration of N_2O_5 . How long would it take for 90% of the N_2O_5 to react?

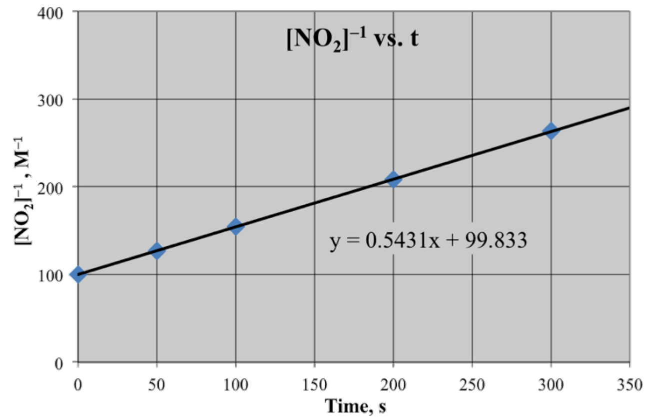
$\ln[A]_t - \ln[A]_0 = -kt$

If 90% of N_2O_5 reacts, than 10% is left unreacted. Thus, starting with 100%, 10 % remains. Since concentrations are used as a ratio, %'s can be used instead.

$$\ln\left(\frac{[A]_t}{[A]_0}\right) = -kt \quad \text{Thus, } \ln\left(\frac{10}{100}\right) = -(0.0069\ s^{-1})t; \quad \boxed{t = 330\ s}$$

2) The following data were obtained for the gas-phase decomposition of nitrogen dioxide, $NO_2(g) \rightarrow NO(g) + \frac{1}{2} O_2(g)$, at 300°C

Time (s)	$[NO_2]$, M	$\ln[NO_2]$	$[NO_2]^{-1}$, M^{-1}
0.0	0.01000	-4.6052	100.0
50.0	0.00787	-4.8447	127.1
100.0	0.00649	-5.0375	154.1
200.0	0.00481	-5.3371	207.9
300.0	0.00380	-5.5728	263.2



Then the data was graphically analyzed. The results were as follows:

- Plot of $[NO_2]$ vs. time : not linear
- Plot of $\ln[NO_2]$ vs. time: not linear
- Plot of $1/[NO_2]$ vs. time: linear as shown in graph→

a. What is the order with respect to NO_2? (The answer is also the order of the overall rxn b/c NO_2 is the only reactant.)
$[NO_2]^{-1}$ vs. time is linear, so the reaction is 2nd order in NO_2

b. Write the general integrated rate law equation for this order of reaction. Then, rearrange the equation so that it is in the form of "y= mx+b."

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

c. The specific equation of the line for the graph is shown on the graph. Use that equation of the line to determine the value of the rate constant for this reaction.
k = slope = 0.543 M^{-1} s^{-1}

d. If the initial concentration of NO_2 is 0.0500 M, what is the remaining concentration after 0.500 hours?

To determine [A], plug the initial concentration, rate constant and time into the equation:

$$\frac{1}{[A]} = \frac{1}{[A_0]} + kt \text{ so } \frac{1}{[NO_2]} = \frac{1}{0.0500\text{ M}} + (0.543\text{ M}^{-1}\text{s}^{-1})(0.500\text{ hr})\left(\frac{3600\text{ s}}{1\text{ hr}}\right)$$

$$\frac{1}{[NO_2]} = 20\text{ M}^{-1} + 977\text{ M}^{-1} = 997\text{ M}^{-1} \text{ and } \boxed{[NO_2] = 1.00\times10^{-3}\text{ M}}$$