

SHOW ALL WORK TO RECEIVE CREDIT.

rate = k    rate = k[A]    rate = k[A]<sup>2</sup>    [A]<sub>t</sub> = -kt + [A]<sub>0</sub>    ln[A]<sub>t</sub> = -kt + ln[A]<sub>0</sub>    R = 8.314 J/(mol·K)

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

$$1/[A]_t = kt + 1/[A]_0$$

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

$$t_{1/2} = 0.693/k$$

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

$$\ln \frac{k_1}{k_2} = \frac{E_a}{R} \left( \frac{1}{T_2} - \frac{1}{T_1} \right)$$

1. (4 Pts) The reaction  $2A \rightarrow \text{products}$  is second order with respect to A. If the concentration of A drops from 1.05 M to 0.815 M in a time of 15.0 min, what is the rate constant for this reaction (the same time units may be used)?

2nd order %  $\frac{1}{[0.815]} = k(15.0 \text{ min}) + \frac{1}{[1.05]}$   
 $k = 0.0183 \text{ M}^{-1} \cdot \text{min}^{-1}$

2. (4 Pts) a. Given the rate law for a reaction,  $\text{rate} = k[A][B]^2$ , where rate is measured in units of  $\text{M s}^{-1}$ , what are the units for the rate constant k?

$$k = \frac{\text{rate}}{[A][B]^2} = \frac{\text{M}}{\text{s}} \frac{1}{\text{M}} \frac{1}{\text{M}^2} = \text{s}^{-1} \cdot \text{M}^{-2}$$

$\text{M}^{-2} \cdot \text{s}^{-1}$

- b. What would be the effect on the rate if the concentration of A and B were both doubled?

$$[2][2]^2 = 8 \text{ Fold rate increase}$$

3. (4 Pts) Determine the rate law that corresponds to the data shown for the reaction  $2A + B \rightarrow C$

Exp.	Initial [A]	Initial [B]	Initial rate
1	0.015	0.022	0.125
2	0.030	0.044	0.500
3	0.060	0.044	0.500
4	0.060	0.066	1.125

General rate law  $\text{rate} = k[A]^x[B]^y$

experiments 2 & 3 show [A] to be 0 order

experiment 1 & 2 show [B] to be 2nd order

$$\frac{\text{rate}_2}{\text{rate}_1} = \frac{k[A]^x[B]^y}{k[A]^x[B]^y} = \frac{(0.030)^x}{(0.015)^x} = \frac{0.500}{0.125}$$

$$2^x = 4$$

$$x = 2$$

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

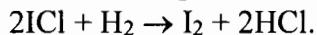
# Quiz 1 Key

4. (4 Pts) The rate constant for the first-order decomposition of  $C_4H_8$  at  $500^\circ C$  is  $9.2 \times 10^{-3} s^{-1}$ . How long will it take for 10.0% of a 0.100 M sample of  $C_4H_8$  to decompose at  $500^\circ C$ ?

90% left  $\ln\left[\frac{90}{100}\right] = (-9.2 \times 10^{-3} s^{-1})(t)$

$t = 11 \text{ sec}$

5. (6 Pts) At a certain temperature, the data below were collected for the reaction below.



- a. Determine the rate law for the reaction.

Initial concentrations (M)		Initial Rate of Formation of $\text{I}_2$
[ICl]	[ $\text{H}_2$ ]	Mol/L·s
0.10	0.10	0.0015
0.20	0.10	0.0030
0.10	0.050	0.00075

rate =  $k[\text{ICl}]^x [\text{H}_2]^y$

$\left(\frac{2}{1}\right) [\text{ICl}] : \left(\frac{0.20}{0.10}\right)^x = \frac{0.0030}{0.0015}$   
 $2^x = 2 \quad x = 1$

$\left(\frac{1}{3}\right) [\text{H}_2] : \left(\frac{0.10}{0.050}\right)^y = \frac{0.0015}{0.00075}$   
 $2^y = 2 \quad y = 1$

rate =  $k[\text{ICl}]^1 [\text{H}_2]^1$

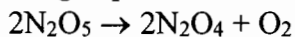
- b. Determine the value and the units of the rate constant.

$k = \frac{\text{rate}}{[\text{ICl}][\text{H}_2]} = \frac{0.0015 \text{ M/s}}{0.10 \text{ M} \cdot 0.10 \text{ M}}$

$= 0.15 \text{ M}^{-1} \cdot \text{s}^{-1}$

$0.15 \frac{\text{L}}{\text{mol} \cdot \text{s}}$

6. (3 Pts) Nitrogen pentoxide decomposes by a first-order process yielding  $\text{N}_2\text{O}_4$  and oxygen.



At a given temperature, the half-life of  $\text{N}_2\text{O}_5$  is 0.90 hr. What is the first-order rate constant for  $\text{N}_2\text{O}_5$  decomposition?

$t_{1/2} = \frac{\ln 2}{k}$

$k = \frac{\ln 2}{t_{1/2}} = \frac{\ln 2}{0.90 \text{ hr}}$

$0.77 \text{ hr}^{-1}$