

Key

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)

1/[A]<sub>t</sub> = kt + 1/[A]<sub>0</sub>

t<sub>1/2</sub> = [A]<sub>0</sub>/2k

t<sub>1/2</sub> = 0.693/k

t<sub>1/2</sub> = 1/k[A]<sub>0</sub>

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

e=mc<sup>2</sup>

**SHOW ALL WORK TO RECEIVE CREDIT.**

1. (5 Pts) A nuclear stress test utilizes a gamma-emitting radioisotope such as thallium-201 to follow the flow of blood through the heart – first at rest, and then under stress. The first-order rate constant for the decay of thallium-201 is  $9.5 \times 10^{-3} \text{ hr}^{-1}$ . Calculate how long it takes for the amount of thallium-201 to fall to 5.0% of its original value.

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

$$\ln \frac{5}{100} = -9.5 \times 10^{-3} (t)$$

$$t = 315 \text{ hrs}$$

2. (4 Pts) The reaction  $2A + B \rightarrow \text{products}$  is second order with respect to A and zero-order with respect to B. The concentration of A is tripled and the concentration of B is doubled. What is the effect on the reaction rate? (How many times faster?)

$$\text{rate} = k [A]^2 [B]^0 = k [A]^2$$

$$\text{so } [3]^2 = 9 \text{ fold increase}$$

3. (2 Pts) Given the rate law for a reaction,  $\text{rate} = k[A]^2$ , where rate is measured in units of  $\text{M min}^{-1}$ , what are the units for the rate constant k?

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

4. At a certain temperature, the data below were collected for the reaction below.



Initial concentrations (M)		Initial Rate of Formation of I <sub>2</sub>
[ICl]	[H <sub>2</sub> ]	Mol/L·s
0.10	0.10	0.0015
0.20	0.10	0.0030
0.10	0.050	0.00075

- a. (6 Pts) Determine the rate law.

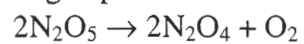
see 2a on 'yellow key'

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b. (5 Pts) Determine the value and the units for the rate constant.

see "2 b" one yellow key.

5. (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 (be sure to include its units)?

see "1a" yellow key

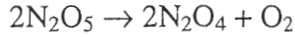
yellow

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)

1/[A]<sub>t</sub> = kt + 1/[A]<sub>0</sub>      t<sub>1/2</sub> = [A]<sub>0</sub>/2k      t<sub>1/2</sub> = 0.693/k      t<sub>1/2</sub> = 1/k[A]<sub>0</sub>      ln(k<sub>1</sub>/k<sub>2</sub>) = (E<sub>a</sub>/R)(1/T<sub>2</sub> - 1/T<sub>1</sub>)      E = mc<sup>2</sup>

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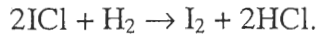
1. (3 Pts) Nitrogen pentoxide decomposes by a first-order process yielding N<sub>2</sub>O<sub>4</sub> and oxygen.



At a given temperature, the half-life of N<sub>2</sub>O<sub>5</sub> is 0.90 hr. What is the first-order rate constant for N<sub>2</sub>O<sub>5</sub> decomposition (be sure to include its units)?

t<sub>1/2</sub> = ln 2 / k      k = ln 2 / 0.90 = 0.77 hr<sup>-1</sup>

2. At a certain temperature, the data below were collected for the reaction below.



	Initial concentrations (M)		Initial Rate of Formation of I <sub>2</sub>
	[ICl]	[H <sub>2</sub> ]	Mol/L·s
1	0.10	0.10	0.0015
2	0.20	0.10	0.0030
3	0.10	0.050	0.00075

a) (6 Pts) Determine the rate law. rate = k [ICl]<sup>x</sup> [H<sub>2</sub>]<sup>y</sup>

For [ICl] use 1 & 2: 0.0030 / 0.0015 = k [0.20]<sup>x</sup> [0.10]<sup>y</sup> / k [0.10]<sup>x</sup> [0.10]<sup>y</sup>

2 = (0.20/0.10)<sup>x</sup>      2 = 2<sup>x</sup>      x = 1 (1st order)

For [H<sub>2</sub>] use 1 & 3: 0.0015 / 0.00075 = k [0.10]<sup>x</sup> [0.10]<sup>y</sup> / k [0.10]<sup>x</sup> [0.050]<sup>y</sup>

2 = (0.10/0.050)<sup>y</sup>      y = 1

rate law = rate = k [ICl]<sup>1</sup> [H<sub>2</sub>]<sup>1</sup>

b) (5 Pts) Determine the value and the units for the rate constant.

k = rate / ([ICl][H<sub>2</sub>]) = M/s / (M \* M) = M<sup>-1</sup>·s<sup>-1</sup> or L·mol<sup>-1</sup>·s<sup>-1</sup>

Value: k = 0.0015 / [0.10][0.10] = 0.15 M<sup>-1</sup>·s<sup>-1</sup>

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# Yellow Key

3. (5 Pts) A nuclear stress test utilizes a gamma-emitting radioisotope, such as thallium-201 to follow the flow of blood through the heart – first at rest, and then under stress. The first-order rate constant for the decay of thallium-201 is  $9.5 \times 10^{-3} \text{ hr}^{-1}$ . Calculate how long it takes for the amount of thallium-201 to fall to 5.0% of its original value.

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

$$\ln \left( \frac{5}{100} \right) = -kt$$

$$\ln \left( \frac{5}{100} \right) = -9.5 \times 10^{-3} t$$

$$t = 3.15 \text{ hrs}$$

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

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

5. (4 Pts) The reaction  $2A + B \rightarrow \text{products}$  is second order with respect to A and zero-order with respect to B. The concentration of A is tripled and the concentration of B is doubled. What is the effect on the reaction rate? (How many times faster?)

$$\text{rate} = k[A]^2[B]^0 = k[A]^2$$

$$\text{(if [A] is tripled)} \text{ rate} = k[3]^2 = 9 \text{ fold increase}$$