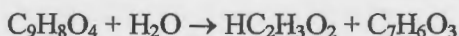


Show all Work to Receive Credit

$$\text{rate} = k \quad \text{rate} = k[A] \quad \text{rate} = k[A]^2 \quad [A]_t = -kt + [A]_0 \quad \ln[A]_t = -kt + \ln[A]_0 \quad R = 8.314 \text{ J}/(\text{mol}\cdot\text{K})$$

$$1/[A]_t = kt + 1/[A]_0 \quad t_{1/2} = [A]_0/2k \quad t_{1/2} = 0.693/k \quad t_{1/2} = 1/k[A]_0 \quad \ln \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right) \quad e=mc^2$$

1. (8 Pts) Aspirin, $\text{C}_9\text{H}_8\text{O}_4$, slowly decomposes at room temperature by reacting with water in the atmosphere to produce acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$, and 2-hydroxybenzoic acid, $\text{C}_7\text{H}_6\text{O}_3$ (this is why old bottles of aspirin often smell like vinegar):



Concentration and rate data for this reaction are given below.

	$[\text{C}_9\text{H}_8\text{O}_4] \text{ (M)}$	$[\text{H}_2\text{O}] \text{ (M)}$	Rate (M/s)
1	0.0100	0.0200	2.4×10^{-13}
2	0.0100	0.0800	9.6×10^{-13}
3	0.0200	0.0200	4.8×10^{-13}

Write the rate law for this reaction and calculate k (be sure to include the correct units).

$$\text{rate} = k [\text{C}_9\text{H}_8\text{O}_4]^x [\text{H}_2\text{O}]^y$$

For $\text{C}_9\text{H}_8\text{O}_4$ use $\frac{3}{1}$ $\frac{4.8 \times 10^{-13}}{2.4 \times 10^{-13}} = \frac{1}{k} \left(\frac{0.0200}{0.0100} \right)^x \left(\frac{0.0200}{0.0200} \right)^y$

$$2 = 2^x \quad x = 1 \quad (\text{first order})$$

For H_2O use $\frac{2}{1}$ $\frac{9.6 \times 10^{-13}}{2.4 \times 10^{-13}} = \frac{1}{k} \left(\frac{0.0800}{0.0200} \right)^y$

$$4 = (4)^y \quad y = 1 \quad (\text{first order})$$

$$\text{rate} = k [\text{C}_9\text{H}_8\text{O}_4] [\text{H}_2\text{O}]$$

$$k = \frac{2.4 \times 10^{-13}}{[0.0100][0.0200]} = 1.2 \times 10^{-9} \text{ M}^{-1} \cdot \text{s}^{-1}$$

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

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

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.

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

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

$$t = 315 \text{ hr} \quad (320 \text{ hr})$$

— more on back —

4. (3 Pts) For the first-order reaction $2\text{N}_2\text{O}_5 \rightarrow 2\text{N}_2\text{O}_4 + \text{O}_2$ at a particular temperature, the half-life of N_2O_5 is 0.90 hr. Determine the value of the rate constant?

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

$$k = \frac{0.693}{0.90}$$

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

5. (4 Pts) For the hypothetical reaction $\text{A} + 3\text{B} \rightarrow 2\text{C}$, the rate of appearance of C given by $(\Delta[\text{C}]/\Delta t)$ may also be expressed as

- A) $\Delta[\text{C}]/\Delta t = \Delta[\text{A}]/\Delta t$
 B) $\Delta[\text{C}]/\Delta t = -(3/2) \Delta[\text{B}]/\Delta t$
 C) $\Delta[\text{C}]/\Delta t = -(2/3) \Delta[\text{B}]/\Delta t$
 D) $\Delta[\text{C}]/\Delta t = -(1/2) \Delta[\text{A}]/\Delta t$

$$\text{rate} = -\frac{\Delta[\text{A}]}{\Delta t} = -\frac{\Delta[\text{B}]}{3\Delta t} = \frac{\Delta[\text{C}]}{2\Delta t} = k[\text{A}]^x[\text{B}]^y[\text{C}]^z$$

$$\frac{-\Delta[\text{B}]}{3\Delta t} = \frac{\Delta[\text{C}]}{2\Delta t}$$

$$\frac{-2\Delta[\text{B}]}{3\Delta t} = \frac{\Delta[\text{C}]}{\Delta t}$$

6. (3 Pts) The reaction $\text{A} + 2\text{B} \rightarrow \text{products}$ has been found to have the rate law, $\text{rate} = k[\text{A}][\text{B}]^2$. If the concentration of A is tripled and the concentration of B is doubled. Predict by what factor the rate of reaction increases.

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

$$[3][2]^2 = 12 \text{ Fold}$$