

Show work where possible

1. (6 Pts) What is the rate law that corresponds to the data shown for the reaction $2A + B \rightarrow C$? (Determine the order with respect to each reactant)

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

*A - [

*A: Since there is no rate change for 2 & 3, the order of A is 0.

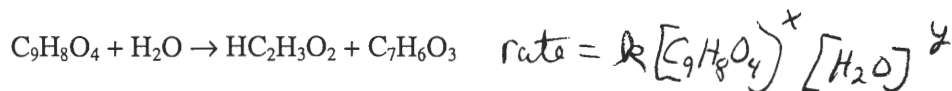
B: picking any pair where B changes, i.e. 2 over 1.

$$\frac{0.500}{0.125} = \left(\frac{0.044}{0.022}\right)^x$$

$$4 = 2^x \quad x = 2$$

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

2. (11 Pts) Aspirin, $C_9H_8O_4$, slowly decomposes at room temperature by reacting with water in the atmosphere to produce acetic acid, $HC_2H_3O_2$, and 2-hydroxybenzoic acid, $C_7H_6O_3$ (this is why old bottles of aspirin often smell like vinegar):



Concentration and rate data for this reaction are given below.

Exp	$[C_9H_8O_4]$ (M)	$[H_2O]$ (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}

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

for $C_9H_8O_4$ use $\frac{\text{Exp 3}}{\text{Exp 2}}$: $\frac{4.8 \times 10^{-13}}{9.6 \times 10^{-13}} = \left(\frac{0.0200}{0.0100}\right)^x \quad x = 1$

for H_2O use $\frac{\text{Exp 2}}{\text{Exp 1}}$: $\frac{9.6 \times 10^{-13}}{2.4 \times 10^{-13}} = \left(\frac{0.0800}{0.0200}\right)^y \quad y = 1$

$$\text{rate} = k [C_9H_8O_4][H_2O]$$

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

- 3 (4 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?

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

4. (4 Pts) Which one of the following is not a valid expression for the of the reaction below? Explain your answer.



a. $-\frac{\Delta[O_2]}{7\Delta t}$

b. $-\frac{\Delta[NO_2]}{4\Delta t}$

c. $\frac{\Delta[H_2O]}{6\Delta t}$

d. $-\frac{\Delta[NH_3]}{4\Delta t}$

e. $\text{rate} = k[NH_3]^x[O_2]^y$

NO_2 is a product, should be positive.

SHOW ALL WORK TO RECEIVE CREDIT

rate = k rate = k[A] rate = k[A]² [A]_t = -kt + [A]₀ ln[A]_t = -kt + ln[A]₀ R = 8.314 J/(mol•K)

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 $\frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$ e=mc²

1. (5 Pts) The reaction 2A + B → products is second order with respect to A and zero-order with respect to B. Starting with 0.135 M of A, what is the concentration of A after 35 min if the rate constant is 0.11 M⁻¹s⁻¹?

rate = k [A]² [B]⁰ ⇒ rate = k [A]²
 2nd order uses $\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$

$$\frac{1}{[A]_t} = 0.11(35 \times 60) + \frac{1}{0.135}$$

$$[A]_t = 4.2 \times 10^{-3} \text{ M}$$

2. (5 Pts) Nitrogen pentoxide decomposes by a first-order process yielding N₂O₄ and oxygen.



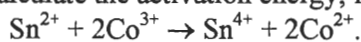
At a given temperature, the half-life of N₂O₅ is 0.90 hr. What is the first-order rate constant for N₂O₅ decomposition?

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

$$k = \frac{\ln 2}{t_{1/2}} = \frac{\ln 2}{0.90} = 0.77 \text{ hr}^{-1}$$

$$2.1 \times 10^{-4} \text{ s}^{-1}$$

3. (5 Pts) Calculate the activation energy, in kJ/mol, for the redox reaction



Temp (°C)	k (1/M•s)
2	3.12 × 10 ³
27	27.0 × 10 ³

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

$$\ln \frac{3.12 \times 10^3}{27.0 \times 10^3} = \frac{E_a}{8.314} \left(\frac{1}{300} - \frac{1}{275} \right)$$

$$E_a = 59200 \text{ J/mol}$$

$$E_a = 59.2 \text{ kJ/mol}$$

4. (5 Pts) The rate constant for the first-order decomposition of C₄H₈ at 500°C is 9.2 × 10⁻³ s⁻¹. How long will it take for 10.0% of a 0.100 M sample of C₄H₈ to decompose at 500°C?

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

$$\ln \frac{90}{100} = -9.2 \times 10^{-3} (t)$$

$$t = 11.4 = 11 \text{ seconds}$$

5. (5 Pts) Given that E_a for a certain biological reaction is 48 kJ/mol and that the rate constant is 2.5 × 10⁻² s⁻¹ at 15°C, what is the rate constant at 37°C?

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

$$\ln \frac{k}{2.5 \times 10^{-2}} = \frac{48 \times 10^3}{8.314} \left(\frac{1}{288} - \frac{1}{310} \right)$$

$$\ln \frac{k}{2.5 \times 10^{-2}} = 1.4226 \dots$$

$$k = 0.104 \text{ s}^{-1}$$