

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 $\frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$ ln $\frac{[A]_t}{[A]_0} = -kt$

3 1. (3 Pts) Nitrogen pentoxide decomposes by a first-order process yielding N₂O₄ and oxygen.
 2N₂O₅ → 2N₂O₄ + O₂
 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} = 0.693/k
 k = $\frac{0.693}{0.90 \text{ hr}} = 0.77 \text{ hr}^{-1}$

7 2. (2 Pts) Given the rate law for a reaction, 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 = $\frac{\text{rate}}{[A][B]^2} = \frac{\text{M} \cdot \text{s}^{-1}}{\text{M} \cdot \text{M}^2} = \text{M}^{-2} \cdot \text{s}^{-1}$

2 3. (4 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 $\frac{k_1}{2.5 \times 10^{-2}} = \frac{48 \times 10^3 \text{ J/mol}}{8.314 \text{ J/mol} \cdot \text{K}} \left(\frac{1}{288 \text{ K}} - \frac{1}{310 \text{ K}} \right)$
 ln $\frac{k_1}{2.5 \times 10^{-2}} = 1.423 \dots$
 k₁ = 0.104 = 1.0 × 10⁻¹ s⁻¹

5 4. (4 Pts) Nitric oxide reacts with chlorine to form nitrosyl chloride, NOCl. Use the following data to determine the rate equation for the reaction and the value of the rate constant.

NO + 1/2 Cl₂ → NOCl

Expt. #	[NO]	[Cl ₂]	Initial Rate
1	0.22	0.065	0.96 M/min
2	0.66	0.065	8.6 M/min
3	0.44	0.032	1.9 M/min

rate = k [NO]^x [Cl₂]^y

For NO use $\frac{\text{Expt \#2}}{\text{Expt \#1}}$:

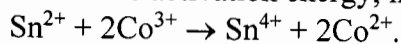
$\frac{8.6}{0.96} = \frac{k}{k} \left(\frac{0.66}{0.22} \right)^x \left(\frac{0.065}{0.065} \right)^y$
 9 = 3^x x = 2

For Cl₂ use any in which Cl₂ changes $\frac{3}{1}$:

$\frac{1.9}{0.96} = \frac{k}{k} \left(\frac{0.44}{0.22} \right)^2 \left(\frac{0.032}{0.065} \right)^y$
 2 = 4 (0.5)^y y = 1

rate = k [NO]² [Cl₂]
 k = $\frac{\text{rate}}{[\text{NO}]^2 [\text{Cl}_2]}$
 k = 300 M⁻² min⁻¹

5. (4 Pts) Calculate the activation energy, in kJ/mol, for the redox reaction



Temp (°C)	k (1/M·s)
27°C	3.12×10^3
27.300	27.0×10^3

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

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

$$E_a = 59207 \text{ kJ/mol} \quad \text{59 kJ/mol}$$

6. (4 Pts) The rate constant for the first-order decomposition of C_4H_8 at 500°C is $9.2 \times 10^{-3} \text{ s}^{-1}$. How long will it take for 10.0% of a 0.100 M sample of C_4H_8 to decompose at 500°C ?

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

$$t = 11.4 \text{ s}$$

7. (4 Pts) The reaction $2\text{A} + \text{B} \rightarrow \text{products}$ 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 $0.11 \text{ M}^{-1}\text{s}^{-1}$?

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

$$\text{SO: } \frac{1}{[\text{A}]_t} = kt + \frac{1}{[\text{A}]_0}$$

$$\frac{1}{[\text{A}]_t} = 0.11 \frac{\text{M}}{\text{s}} (2100 \text{ s}) + \frac{1}{[0.135]}$$

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