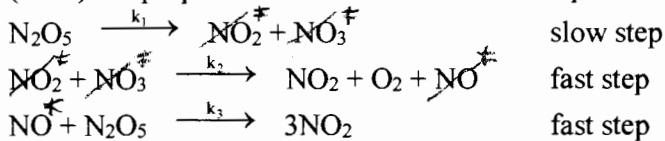


Show all work to receive credit.

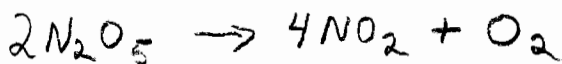
rate = k rate = k[A] rate = k[A]² [A]_t = -kt + [A]₀ ln[A]_t = -kt + ln[A]₀ R = 8.314 J/(mol•K)

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. (4 Pts) A proposed mechanism for the decomposition of N₂O₅ is as follows:



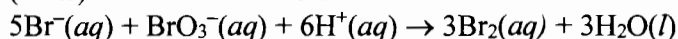
a. (2 Pts) What is the overall chemical equation predicted by this mechanism?



b. (2 Pts) What is the rate law?

$$\text{rate} = k[\text{N}_2\text{O}_5]$$

2. (2 Pts) The rate law for the chemical reaction



has been determined experimentally to be Rate = k[Br⁻][BrO₃⁻][H⁺]². What is the overall order of the reaction?

$$(1+1+2) = 4 \text{ fourth}$$

3. (4 Pts) A reaction that is second-order in one reactant has a rate constant of 1.6 × 10⁻² L/(mol · s). If the initial concentration of the reactant is 0.400 mol/L, how long will it take for the concentration to become 0.200 mol/L?

$$\begin{aligned} \frac{1}{[A]_t} &= kt + \frac{1}{[A]_0} \\ \frac{1}{[0.200]} &= 1.6 \times 10^{-2}(t) + \frac{1}{0.400} \\ t &= 156 \text{ sec} \end{aligned}$$

4. (4 Pts) The rate constant for a first-order reaction is 1.5 × 10⁻² s⁻¹ at 748 K and 3.8 × 10⁻² s⁻¹ at 824 K. What is the activation energy?

$$\begin{aligned} \ln \frac{1.5 \times 10^{-2}}{3.8 \times 10^{-2}} &= \frac{E_a}{8.314} \left(\frac{1}{824} - \frac{1}{748} \right) \\ E_a &= 62,674 \frac{\text{J}}{\text{mol}} = 63 \frac{\text{kJ}}{\text{mol}} \end{aligned}$$

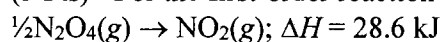
More questions on back.

5. (3 Pts) A first-order chemical reaction is observed to have a rate constant of 43 min^{-1} . What is the corresponding half-life for the reaction?

$$t_{1/2} = \frac{\ln 2}{43} = 1.6 \times 10^{-2} \text{ min}$$

or 1.0 s

6. (5 Pts) For the first-order reaction



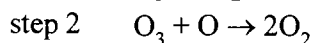
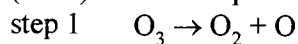
the rate constant is $k = 2.23 \times 10^5 \text{ s}^{-1}$ at -11°C , and the activation energy is 53.7 kJ/mol . What is the rate constant at 23°C ?

$$\ln \frac{k_2}{2.23 \times 10^5} = \frac{53.7 \times 10^3}{8.314} \left(\frac{1}{262} - \frac{1}{296} \right)$$

$$\ln \frac{k_2}{2.23 \times 10^5} = 2.8317 \dots$$

? $k_2 = \underline{3.78 \times 10^6 \text{ s}^{-1}}$

7. (3 Pts) The decomposition of ozone may occur through the two-step mechanism shown below:



The oxygen atom is considered to be a(n)

A) reactant.

B) product.

C) catalyst.

D) reaction intermediate.

E) activated complex.

CHM152 Quiz (2B) 25 Pts Spring 2020 Name: Key Pink
 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)

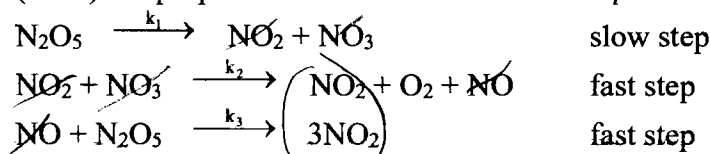
1/[A]_t = kt + 1/[A]₀ t_{1/2} = [A]₀/2k t_{1/2} = 0.693/k t_{1/2} = 1/k[A]₀ ln $\frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$ E=mc²

1. (4 Pys) The rate constant for a first-order reaction is $1.8 \times 10^{-2} \text{ s}^{-1}$ at 698 K and $3.8 \times 10^{-2} \text{ s}^{-1}$ at 820 K. What is the activation energy?

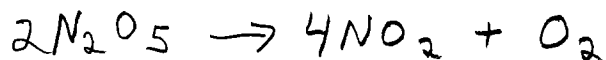
$$\ln \frac{1.8 \times 10^{-2}}{3.8 \times 10^{-2}} = \frac{E_a}{8.3145 \frac{\text{J}}{\text{mol} \cdot \text{K}}} \left(\frac{1}{820} - \frac{1}{698} \right)$$

$$E_a = 29,145 \text{ J/mol or } 29 \text{ kJ/mol}$$

2. (4 Pts) A proposed mechanism for the decomposition of N₂O₅ is as follows:



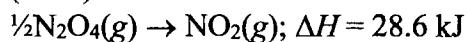
a. (2 Pts) What is the overall chemical equation predicted by this mechanism?



b. (2 Pts) What is the rate law?

$$\text{rate} = k [\text{N}_2\text{O}_5]$$

3. (5 Pts) For the first-order reaction



the rate constant is $k = 1.14 \times 10^5 \text{ s}^{-1}$ at -18°C , and the activation energy is 53.7 kJ/mol. What is the rate constant at 1°C ?

$$\ln \frac{k_2}{1.14 \times 10^5} = \frac{53.7 \times 10^3 \text{ J/mol}}{8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}}} \left(\frac{1}{255} - \frac{1}{274} \right)$$

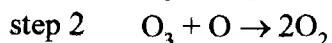
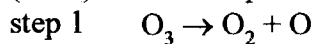
$$= 1.756 \dots$$

$$k_2 = 660246 \text{ s}^{-1} \text{ or } 6.6 \times 10^5 \text{ s}^{-1}$$

More questions on back.

Key (pink)

4. (3 Pts) The decomposition of ozone may occur through the two-step mechanism shown below:



The oxygen atom is considered to be a(n)

A) reactant.

B) product.

C) catalyst

D) reaction intermediate.

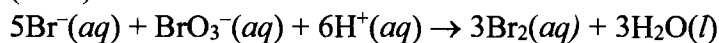
E) activated complex.

5. (4 Pts) A reaction that is second-order in one reactant has a rate constant of $3.2 \times 10^{-2} \text{ L}/(\text{mol} \cdot \text{s})$. If the initial concentration of the reactant is 0.380 mol/L , how long will it take for the concentration to become 0.190 mol/L ?

$$\frac{1}{[0.190]} = 3.2 \times 10^{-2} (t) + \frac{1}{[0.380]}$$

$$t = 82.2 \text{ s}$$

6. (2 Pts) The rate law for the chemical reaction



has been determined experimentally to be $\text{Rate} = k[\text{Br}^-][\text{BrO}_3^-][\text{H}^+]^2$. What is the overall order of the reaction?

$$1 + 1 + 2 = 4$$

fourth

7. (3 Pts) A first-order chemical reaction is observed to have a rate constant of 29 min^{-1} . What is the corresponding half-life for the reaction?

$$t_{1/2} = \frac{\ln 2}{k} = \frac{\ln 2}{29 \text{ min}^{-1}} = 0.0239 \text{ min}$$

$$\text{or } 1.4 \text{ s}$$