CHM152 Quiz 2a 25 Pts Spring 2020 Name: Key

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)

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

1. (4 Pts) A proposed mechanism for the decomposition of N_2O_5 is as follows:

$N_2O_5 \xrightarrow{k_1} \cancel{NO_2} + \cancel{NO_3}$ slow step

$\cancel{NO_2} + \cancel{NO_3} \xrightarrow{k_2} NO_2 + O_2 + \cancel{NO}$ fast step

$\cancel{NO} + N_2O_5 \xrightarrow{k_3} 3NO_2$ fast step

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

$2N_2O_5 \rightarrow 4NO_2 + O_2$

b. (2 Pts) What is the rate law? rate = $k[N_2O_5]$

2. (2 Pts) The rate law for the chemical reaction
$5Br^-(aq) + BrO_3^-(aq) + 6H^+(aq) \rightarrow 3Br_2(aq) + 3H_2O(l)$
has been determined experimentally to be Rate = $k[Br^-][BrO_3^-][H^+]^2$. What is the overall order of the reaction?

$(1+1+2)=4$ 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?

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

$$\frac{1}{[0.200]} = 1.6\times10^{-2}(t) + \frac{1}{0.400}$$

$t = 156$ sec

4. (4 Pys) 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?

$$\ln\frac{1.5\times10^{-2}}{3.8\times10^{-2}} = \frac{E_a}{8.314}\left(\frac{1}{824} - \frac{1}{748}\right)$$

$$E_a = 62{,}674 \ \frac{J}{mol} = 63 \ \frac{kJ}{mol}$$

More questions on back.

Key

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?

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

or 1.0 s

6. (5 Pts) For the first-order reaction
$\frac{1}{2}N_2O_4(g) \rightarrow NO_2(g)$; $\Delta H = 28.6$ kJ
the rate constant is $k = 2.23 \times 10^5\ s^{-1}$ at –11°C, and the activation energy is 53.7 kJ/mol. What is the rate constant at 23°C?

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

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

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

7. (3 Pts) The decomposition of ozone may occur through the two-step mechanism shown below:
step 1 $O_3 \rightarrow O_2 + O$
step 2 $O_3 + O \rightarrow 2O_2$
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)

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

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

$$\ln\frac{1.8 \times 10^{-2}}{3.8 \times 10^{-2}} = \frac{E_a}{8.314\ \frac{J}{mol \cdot K}}\left(\frac{1}{820} - \frac{1}{698}\right)$$

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

2. (4 Pts) A proposed mechanism for the decomposition of N_2O_5 is as follows:

$N_2O_5 \xrightarrow{k_1} \cancel{NO_2} + \cancel{NO_3}$ slow step
$\cancel{NO_2} + \cancel{NO_3} \xrightarrow{k_2} NO_2 + O_2 + \cancel{NO}$ fast step
$\cancel{NO} + N_2O_5 \xrightarrow{k_3} 3NO_2$ fast step

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

$2N_2O_5 \rightarrow 4NO_2 + O_2$

b. (2 Pts) What is the rate law? rate = $k[N_2O_5]$

3. (5 Pts) For the first-order reaction
$\frac{1}{2}N_2O_4(g) \rightarrow NO_2(g)$; $\Delta H = 28.6$ kJ
the rate constant is $k = 1.14 \times 10^5\ s^{-1}$ at –18°C, and the activation energy is 53.7 kJ/mol. What is the rate constant at 1°C?

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

$= 1.756\ldots\ldots$

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

More questions on back.

Key (Pink)

4. (3 Pts) The decomposition of ozone may occur through the two-step mechanism shown below:
   step 1 $O_3 \rightarrow O_2 + O$
   step 2 $O_3 + O \rightarrow 2O_2$
   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 3.2×10^{-2} L/(mol · 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?

$$\frac{1}{[0.190]} = 3.2\times10^{-2}(t) + \frac{1}{[0.380]}$$

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

6. (2 Pts) The rate law for the chemical reaction
   $5Br^-(aq) + BrO_3^-(aq) + 6H^+(aq) \rightarrow 3Br_2(aq) + 3H_2O(l)$
   has been determined experimentally to be Rate = $k[Br^-][BrO_3^-][H^+]^2$. 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?

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

or 1.4 s