

Old Exam 2

Name: Key

CHM 112 Exam 1 Spring

Short Answer

1. For a first-order reaction that has a rate constant of $6.9 \times 10^{-4} \text{ s}^{-1}$;

- if the initial concentration of the only reactant is 0.25 M, what is the concentration after 8.4 min?
- How long will it take for the concentration to decrease to 0.15 M?
- How long will it take for the reaction to be 60% complete?

a) $\ln[A] = -kt + \ln[A]_0$

$$k = 6.9 \times 10^{-4} \frac{\text{s}^{-1}}{\text{min}} \quad 8.4 \text{ min} \left(\frac{60 \text{ s}}{\text{min}} \right) = 504 \text{ s}$$

$$\ln[A] = -(6.9 \times 10^{-4} \text{ s}^{-1})(504 \text{ s}) + \ln[0.25 \text{ M}]$$

$$\ln[A] = -0.34976 + -1.3863$$

$$\ln[A] = -1.734$$

$$[A] = e^{-1.734} = 0.1766 \text{ M} \rightarrow \boxed{0.18 \text{ M}}$$

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

$$\ln \frac{[0.15]}{[0.25]} = -6.9 \times 10^{-4} \text{ s}^{-1} (t)$$

$$\frac{-0.5108}{-6.9 \times 10^{-4} \text{ s}^{-1}} = \frac{-6.9 \times 10^{-4} \text{ s}^{-1} (t)}{-6.9 \times 10^{-4} \text{ s}^{-1}}$$

$$\boxed{t = 740 \text{ s}}$$

c) 60% complete = 40% left over

$$\ln \left(\frac{0.4}{1} \right) = -(6.9 \times 10^{-4} \text{ s}^{-1})(t)$$

$$\frac{-0.91629}{-6.9 \times 10^{-4} \text{ s}^{-1}} = \frac{-6.9 \times 10^{-4} \text{ s}^{-1} (t)}{-6.9 \times 10^{-4} \text{ s}^{-1}}$$

$$\boxed{t = 1,328 \text{ s}}$$

2. The rate constants for a reaction were determined at two temperatures. At 100.0 K the rate constant is $2.0 \times 10^3 \text{ s}^{-1}$, and at 500.0 K the rate constant is $4.07 \times 10^7 \text{ s}^{-1}$. Calculate the activation energy for the reaction.

$$k_1 = 100.0 \text{ K } 2.0 \times 10^3 \text{ s}^{-1}$$

$$k_2 = 500.0 \text{ K } 4.07 \times 10^7 \text{ s}^{-1}$$

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

$$\ln\left(\frac{4.07 \times 10^7 \text{ s}^{-1}}{2.0 \times 10^3 \text{ s}^{-1}}\right) = \frac{E_a}{8.314 \text{ J/mol K}} \left(\frac{1}{100.0 \text{ K}} - \frac{1}{500.0 \text{ K}} \right)$$

$$\ln(20350) = \frac{E_a}{8.314 \text{ J/mol K}} (0.01000 \text{ K}^{-1} - 0.002000 \text{ K}^{-1})$$

$$9.9208 = \frac{E_a}{8.314 \text{ J/mol K}} (0.008 \text{ K}^{-1})$$

$$9.9208 = (E_a)(0.0009622 \text{ mol/J})$$

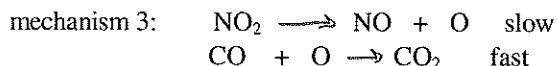
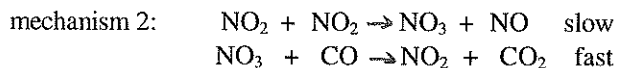
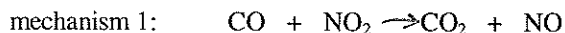
$$E_a = 10310.5 \text{ J/mol} = 1.03 \times 10^4 \text{ J/mol}$$

$$E_a = 10.3 \text{ kJ/mol}$$

3. The reaction between carbon monoxide and nitrogen dioxide has the experimentally determined rate law; rate = $k[\text{NO}_2]^2$



The following mechanisms have been proposed;



Which mechanism is most likely. Briefly explain your choice for each possibility

Mechanism 1: Not the correct mechanism.

Rate would be $\text{Rate} = k[\text{CO}][\text{NO}_2]$ for the one elementary step. Needs to be second order in $[\text{NO}_2]$.

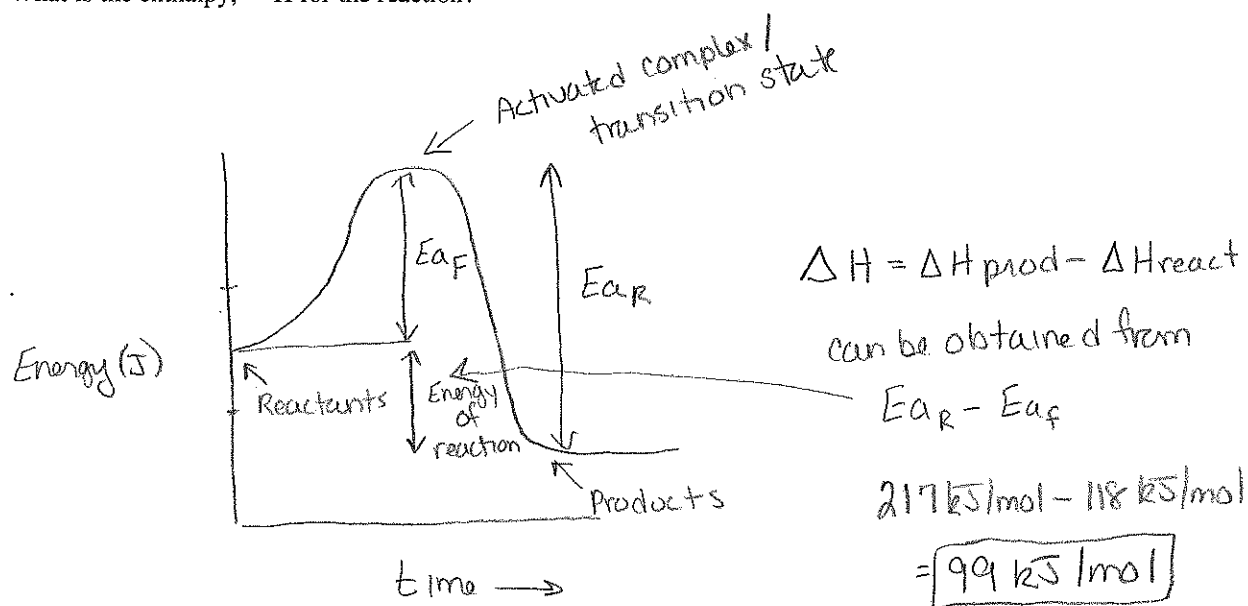
* Mechanism 2: Correct mechanism *

Rate of slow (rate determining step) = $k[\text{NO}_2]^2$
same as overall rate!

Mechanism 3: Not the correct mechanism

Rate of slow (rate determining step) = $k[\text{NO}_2]$
still not second order in $[\text{NO}_2]$

4. In a reversible reaction, the energy of activation for the forward reaction is 118 kJ/mol, and the energy of activation for the reverse direction is 217 kJ/mol. Sketch a reaction coordinate diagram. Label completely. What is the enthalpy, ΔH for the reaction?



5. In a kinetic study of the reaction;



the data for the initial rates;

Initial concentrations (M)		Rate (M/s)
[NO]	[H ₂]	
6.4 x 10 ⁻³	2.2 x 10 ⁻³	2.6 x 10 ⁻⁵
12.8 x 10 ⁻³	2.2 x 10 ⁻³	1.0 x 10 ⁻⁴
6.4 x 10 ⁻³	4.4 x 10 ⁻³	5.1 x 10 ⁻⁵

Obtain the rate law

What is the value of the rate constant?

[NO] use exp 1 & 2

$$\left(\frac{12.8 \times 10^{-3} \text{ M}}{6.4 \times 10^{-3} \text{ M}} \right)^m = \frac{1.0 \times 10^{-4} \text{ M/s}}{2.6 \times 10^{-5} \text{ M/s}}$$

$$(2)^m = 3.8 \quad (\text{likely 4 w/ experimental error})$$

$$m = 2$$

[H₂] use exp 1 & 3

$$\left(\frac{4.4 \times 10^{-3} \text{ M}}{2.2 \times 10^{-3} \text{ M}} \right)^n = \frac{5.1 \times 10^{-5}}{2.6 \times 10^{-5}}$$

$$2^n = 1.96 \approx 2$$

$$n = 1$$

need k
 ↓
 Rate Law = $k [\text{NO}]^2 [\text{H}_2]$

k : exp #1 $2.6 \times 10^{-5} \text{ M/s} = k [6.4 \times 10^{-3} \text{ M}]^2 [2.2 \times 10^{-3} \text{ M}]$ $2.6 \times 10^{-5} \text{ M/s} = k (9.01 \times 10^{-8} \text{ M}^3)$ $k = 2.9 \times 10^2 \text{ M}^{-2} \text{ s}^{-1}$
 exp #2 $1.0 \times 10^{-4} \text{ M/s} = k [12.8 \times 10^{-3} \text{ M}]^2 [2.2 \times 10^{-3} \text{ M}]$ $1.0 \times 10^{-4} \text{ M/s} = k (3.6045 \times 10^{-7} \text{ M}^3)$ $k = 2.77 \times 10^2 \text{ M}^{-2} \text{ s}^{-1}$
 exp #3 $5.1 \times 10^{-5} \text{ M/s} = k [6.4 \times 10^{-3} \text{ M}]^2 [4.4 \times 10^{-3} \text{ M}]$ $5.1 \times 10^{-5} \text{ M/s} = k (1.802 \times 10^{-7} \text{ M}^3)$ $k = 2.83 \times 10^2 \text{ M}^{-2} \text{ s}^{-1}$

Aug: $2.9 \times 10^2 \text{ M}^{-2} \text{ s}^{-1}$

$2.77 \times 10^2 \text{ M}^{-2} \text{ s}^{-1}$

$2.83 \times 10^2 \text{ M}^{-2} \text{ s}^{-1}$

$\frac{2.83 \times 10^2 \text{ M}^{-2} \text{ s}^{-1}}{1.461 \times 10^3 \div 3} = 4.87 \times 10^2 \text{ M}^{-2} \text{ s}^{-1}$

$\rightarrow 4.9 \times 10^2 \text{ M}^{-2} \text{ s}^{-1} = k$
 Rate = $4.9 \times 10^2 \text{ M}^{-2} \text{ s}^{-1} [\text{NO}]^2 [\text{H}_2]$

6. The rate constant for a second-order reaction is 1.15/M ^{← should be minutes} at 25 degrees C. How long (seconds) will it take for the concentration of the single reactant to decrease from 0.55 M to 0.45 M?

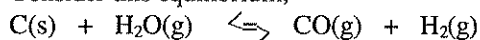
$$\ln\left(\frac{[A]}{[A_0]}\right) = -kt$$

$$\ln\left(\frac{0.45\text{M}}{0.55\text{M}}\right) = -1.15(t)$$

$$\frac{-0.20067}{-1.15} = \frac{-1.15(t)}{-1.15}$$

$$t = 0.1745\text{min} \left(\frac{60\text{s}}{\text{min}}\right) = \boxed{10.5\text{s}}$$

7. Consider this equilibrium;



Which direction will this reaction go if;

- a) CO is added to the reaction mixture *toward reactants*
- b) H₂O is condensed and removed from the reaction mixture *toward reactants*
- c) C is added to the reaction mixture *toward products*

Equilibrium is not on your exam 1 so you do not need to do this type of problem on this exam.