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 x 10^{-4} s^{-1};

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

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

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

$\ln[A] = -(6.9\times10^{-4}s^{-1})(504s) + \ln[0.25M]$

$\ln[A] = -0.34776 + -1.3863$

$\ln[A] = -1.734$

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

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

$\ln\frac{[0.15]}{[0.25]} = -6.9\times10^{-4}s^{-1}(t)$

$\frac{-0.5108}{-6.9\times10^{-4}s^{-1}} = \frac{-\cancel{6.9\times10^{-4}s^{-1}}(t)}{-\cancel{6.9\times10^{-4}s^{-1}}}$

$\boxed{t = 740s}$

c) 60% complete = 40% left over

$\ln\left(\frac{0.4}{1}\right) = -(6.9\times10^{-4}s^{-1})(t)$

$\frac{-0.91629}{-6.9\times10^{-4}s^{-1}} = \frac{-6.9\times10^{-4}s^{-1}(t)}{-6.9\times10^{-4}s^{-1}}$ $\quad \boxed{t = 1,328\ s}$

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

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

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

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

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

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

$$9.9208 = \frac{Ea}{8.314 J/mol \cancel{K}}\left(0.008 \cancel{K^{-1}}\right)$$

$$9.9208 = (Ea)(0.000962 22\ mol/J)$$

$$Ea = 10310.5\ J/mol = \boxed{1.03 \times 10^4\ J/mol}$$

$$\boxed{Ea = 10.3\ kJ/mol} \text{ or}$$

3. The reaction between carbon monoxide and nitrogen dioxide has the experimentally determined rate law; rate = $k[NO_2]^2$

$$CO + NO_2 \rightarrow CO_2 + NO$$

The following mechanisms have been proposed;

mechanism 1: $CO + NO_2 \rightarrow CO_2 + NO$

mechanism 2: $NO_2 + NO_2 \rightarrow NO_3 + NO$ slow
$NO_3 + CO \rightarrow NO_2 + CO_2$ fast

mechanism 3: $NO_2 \rightarrow NO + O$ slow
$CO + O \rightarrow CO_2$ fast

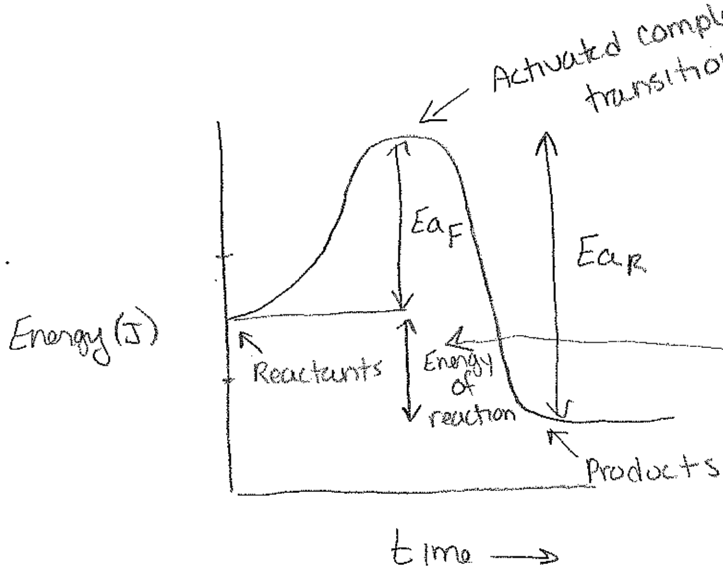
Which mechanism is most likely. Briefly explain your choice for each possibility

Mechanism 1: Not the correct mechanism.
Rate would be Rate = $k[CO][NO_2]$ for the one elementary step. Needs to be second order in $[NO_2]$.

* Mechanism 2: Correct mechanism *
Rate of slow (rate determining step) = $k[NO_2]^2$
Same as overall rate!

Mechanism 3: Not the correct mechanism
Rate of slow (rate determining step) = $k[NO_2]$
still not second order in $[NO_2]$

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, H for the reaction?



$\Delta H = \Delta H_{prod} - \Delta H_{react}$

can be obtained from

$Ea_R - Ea_f$

$217\ kJ/mol - 118\ kJ/mol$

$= \boxed{99\ kJ/mol}$

5. In a kinetic study of the reactiom;

$2\ NO(g) + H_2(g) \qquad N_2O(g) + H_2O(g)$

the data for the initial rates;

Initial concentrations (M)		Rate (M/s)
[NO]	[H_2]	
6.4×10^{-3}	2.2×10^{-3}	2.6×10^{-5}
12.8×10^{-3}	2.2×10^{-3}	1.0×10^{-4}
6.4×10^{-3}	4.4×10^{-3}	5.1×10^{-5}

Obtain the rate law

What is the value of the rate constant?

[NO] use exp 1 & 2

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

$(2)^m = 3.8$ (likely 4 w/ experimental error)

$m = 2$

$[H_2]$ use exp 1 & 3

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

need k

Rate Law = $k[NO]^2[H_2]$

$2^n = 1.96 \approx 2$

$n = 1$

k: exp#1 $2.6 \times 10^{-5} M/s = k[6.4 \times 10^{-3} M]^2 [2.2 \times 10^{-3} M]$ $\quad 2.6 \times 10^{-5} M/s = k(9.01 \times 10^{-8} M^3)$ $k = 2.9 \times 10^2\ M^{-2}s^{-1}$

exp#2 $1.0 \times 10^{-4} M/s = k[12.8 \times 10^{-3} M]^2 [2.2 \times 10^{-3} M]$ $\quad 1.0 \times 10^{-4} M/s = k(3.6045 \times 10^{-7} M^3)$ $k = 2.77 \times 10^2\ M^{-2}s^{-1}$

exp#3 $5.1 \times 10^{-5} M/s = k[6.4 \times 10^{-3} M]^2 [4.4 \times 10^{-3} M]$ $\quad 5.1 \times 10^{-5} M/s = k(1.802 \times 10^{-7} M^3)$ $k = 2.83 \times 10^2\ M^{-2}s^{-1}$

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

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

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

$1.461 \times 10^3 \div 3 = 4.87 \times 10^2\ M^{-2}s^{-1} \rightarrow 4.9 \times 10^2\ M^{-2}s^{-1} = k$

Rate $= 4.9 \times 10^2\ M^{-2}s^{-1}\ [NO]^2[H_2]$

should be minutes

6. The rate constant for a first-order reaction is 1.15/M 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?

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

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

$$\frac{-0.2006\ln}{-1.15} = \frac{-\cancel{1.15}\,(t)}{-\cancel{1.15}}$$

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

7. Consider this equilibrium;
$C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$

Which direction will this reaction go if;
a) CO is added to the reaction mixture toward reactants
b) H_2O 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.