You must show all of your work for full credit.

1. In a manner similar to crystal violet experiment, the following kinetic data were obtained for the decomposition of nitrosyl bromide: (**12** pts.)

t(s)	[NOBr]	
0.0	0.0100	⟩ (0.0029)
2.00	0.0071	⟩ 0.0016
4.00	0.0055	⟩ 0.0010
6.00	0.0045	⟩ 0.0007
8.0	0.0038	⟩ 0.0005
10.0	0.0033	

$$2NOBr\ (g) \longrightarrow 2NO\ (g) + Br_2\ (g)$$

a) Use graphical methods to predict the order of the reaction with respect to NOBr.
b) Use graphical method to estimate the rate constant.
c) Determine the half life for the decomposition of NOBr.
d) How much time is required for the concentration of NOBr to decrease to 2.0 x 10^{-4} M?

a. ∴ the order is 2 w/ respect to [NOBr].

b. $pt_1 (10, 303.0)$ $pt_2 (1, 100.0)$

note: $\frac{1}{[NOBr]_t} = kt + \frac{1}{[NOBr]_0}$

$$\frac{\Delta Y}{\Delta X} = \frac{303.0 - 100}{9} = 22.56$$

$$y = mx + b$$
slope

c.) $t_{1/2} = \frac{1}{k[NOBr]_0} = \frac{1}{22.56\,(0.01)} \approx 4$ seconds.

d.) $\left(\frac{1}{2.0 \times 10^{-4} M}\right) = [22.56]\,t + \frac{1}{[0.0100]}$

$t \approx 217.2$ sec.

or 3.6 minutes.

t(min)			
0	ln(0.01) = -4.60	> 0.35	
2	ln(0.0071) = -4.95	> 0.25	
4	ln(0.0055) = -4.80		
6	ln(0.0045) = -5.40	> 0.17	
8	ln(0.0038) = -5.57	> 0.14	
10	ln(0.0033) = -5.7		

1st order rate law:

$$ln[NO_2Br]_t = -kt + ln[NO_2Br]_0$$

1st order test. ln[NO₂Br] vs t

t(time minutes)

$\frac{1}{[NO_2Br]}$ vs time.

$\frac{1}{0.01} = 100$ at t(min)

$\frac{1}{0.007} = 140.84$; 2

$\frac{1}{0.0055} = 181.8$; 4

$\frac{1}{0.0045} = 222.22$; 6

$\frac{1}{0.0038} = 263.15$; 8
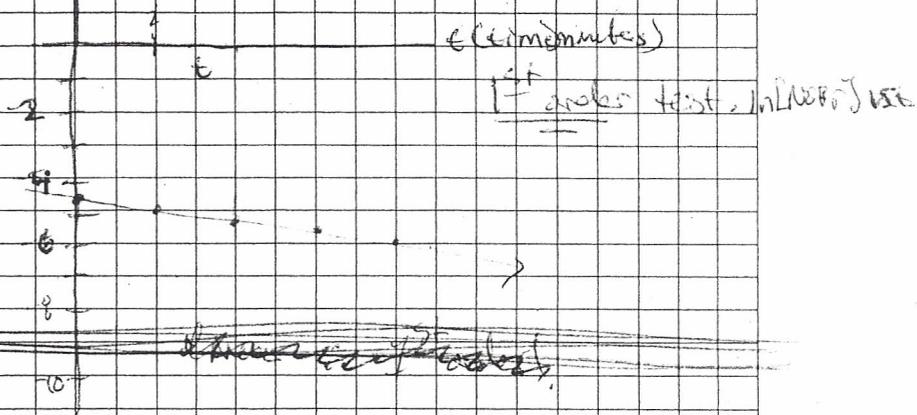
$\frac{1}{0.0033} = 303.03$ at ; 10.

∴ 2nd order!

2nd order test.

for [NO₂Br] vs t

2. At elevated temperatures, CH_3NC isomerizes to CH_3CN;

$$CH_3NC \text{ (g)} \longrightarrow CH_3CN \text{ (g)}$$

The dependence of the rate constant on temperature is studied and the graph below prepared from the results. Determine the activation energy. (8 pts.)

$$\ln(k) = \frac{-E_a}{R}\left(\frac{1}{T}\right) + \ln(A)$$

$$\underset{Y}{\uparrow} = \underset{M}{\uparrow}\ \underset{X}{\uparrow}\ + B$$

$B \approx -5.5$

$M = \dfrac{\Delta Y}{\Delta x} = \dfrac{-E_a}{R}$

$\Rightarrow \dfrac{-11-(-5.5)}{0.0022-0.0019} \approx \dfrac{-5.5}{0.0201} \approx -273.63$

$\therefore \dfrac{-E_a}{8.314 \text{ J/mol-k}} = -273.63$

$-E_a = -2274.98$

$\therefore E_a \approx 2,274.98 \text{ J/mol-k}$ or $\approx 2.27 \text{ kJ/mol-k}$
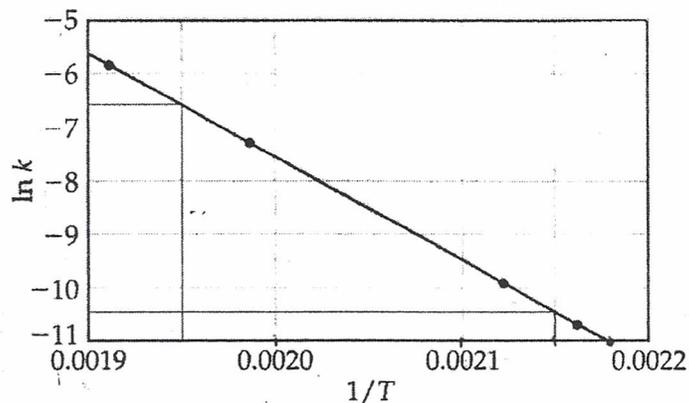
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3. For the reaction $2NO\,(g) + H_2\,(g) \longleftrightarrow N_2\,(g) + 2\,H_2O\,(g)$
These data were obtained at 1100 K: **(10 pts)**

[NO] (mol/L)	[H2] (mol/L)	Initial Rate (M/S)
1) 5.00×10^{-3}	2.50×10^{-3}	3.0×10^{-3}
2) 15.0×10^{-3}	2.50×10^{-3}	9.0×10^{-3}
3) 15.0×10^{-3}	10.0×10^{-3}	3.6×10^{-2}

What is the order with respect to NO?
What is the order with respect to H_2?
What is the overall order of the reaction?

rate law = $k[NO]^m[H_2]^n$

$\dfrac{rate\ 2}{rate\ 1} = \dfrac{9.0 \times 10^{-3}}{3.0 \times 10^{-3}} = \dfrac{k[15.0 \times 10^{-3}]^m[2.5 \times 10^{-3}]^n}{k[5.0 \times 10^{-3}]^m[2.5 \times 10^{-3}]^n}$

$3 = 3^m \quad \therefore \quad \underline{m = 1}$

$\dfrac{rate\ 3}{rate\ 2} = \dfrac{3.6 \times 10^{-2}}{9.0 \times 10^{-3}} = \dfrac{k[15.0 \times 10^{-3}]^m[10.0 \times 10^{-3}]^n}{k[15.0 \times 10^{-3}]^m[2.5 \times 10^{-3}]^n}$

$4 = 4^n \quad \therefore \quad \underline{n = 1}$

the overall order of the reaction is $\underline{\underline{2}}$.

rate $= k[NO]^1[H_2]^1$

4. Draw a rough sketch of the energy profile for each of the following cases? **(5 pts)**

    a) $\Delta E = 10$ KJ/mol, Ea = 25 KJ/mol
    b) $\Delta E = -10$ KJ/mol, Ea = 50 KJ/mol

a.) ; endothermic   b.)
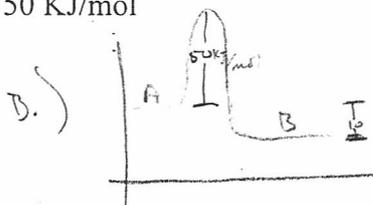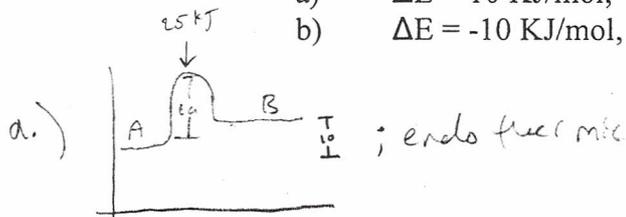
5. In the iodination of acetone experiment, what set of experimental conditions should you provide to ensure a constant reaction rate over time. (5 pts)

The rate of reaction must not totally depend on the concentration of Acetone and H^+. Concentrations of these substances must be in excess and thus the Iodine is made to be the limiting reagent in this experiment. If the initial concentrations of Acetone and H^+ are much larger than Iodine, then though the reaction is taking place, concentration [Ac] and [H^+] will not change dramatically and thus stay constant over time. In this way, one is able to remove the time dependence of the reaction rate in respect to Acetone and H^+. The reason for this is because it is experimentally difficult to know what the rate of reaction is fully when Ac and H^+ react.

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$\frac{35}{40}$