Unit 5 Review for Midterm

① a) O_2 is disappearing at 6x the rate of $C_6H_{12}O_6$

CO_2 + H_2O are forming at the same rate

CO_2 + H_2O are forming at the same rate as O_2 disappearing

CO_2 + H_2O are forming 6x faster than $C_6H_{12}O_6$ is disappearing

b) $-\frac{\Delta[C_6H_{12}O_6]}{\Delta t} = -\frac{1}{6}\frac{\Delta[O_2]}{\Delta t} = \frac{1}{6}\frac{\Delta[CO_2]}{\Delta t} = \frac{1}{6}\frac{\Delta[H_2O]}{\Delta t}$

② a) Rate = $k[N_2O_5]^1$

b) Rate = $(4.82 \cdot 10^{-3}\ s^{-1})(.200\ M) = 9.64 \cdot 10^{-4}\ M/s$

c) Rate triples Rate = $k(3[N_2O_5])^1$

③ a) Rate = $k[S_2O_8^{2-}]^1[I^-]^1$

b) use expt 1 $2.2 \cdot 10^{-4}\ M/s = k(.080M)(.034M)$ $k = .081\ M^{-1}s^{-1}$

c) $5.7 \cdot 10^{-4}\ \frac{M}{s} = (.081\ M^{-1}s^{-1})(.280M)[I^-]$ $[I^-] = .025\ M$

d) 2nd order

e) Reaction mechanism is bimolecular

④ a)

$$2NO \rightarrow \cancel{N_2O_2}$$
$$\cancel{N_2O_2} + H_2 \rightarrow \cancel{N_2O} + H_2O$$
$$\cancel{N_2O} + H_2 \rightarrow N_2 + H_2O$$
$$2NO + 2H_2 \rightarrow N_2 + 2H_2O$$

b) Rate = $k_2[N_2O_2][H_2]$ ⟶ N_2O_2 is intermediate + can't be in rate law.

$Rate_f = k_f[NO]^2$

$Rate_r = k_r[N_2O_2]$ Treat step 1 as fast equilib.

$Rate_f = Rate_r$ $k_f[NO]^2 = k_r[N_2O_2]$

$\frac{k_f[NO]^2}{k_r} = [N_2O_2]$

plug in $K_1[NO]^2 = [N_2O_2]$

Rate = $k_2\left(\frac{k_f}{k_r}[NO]^2\right)[H_2]$

Rate = $K[NO]^2[H_2]$

c) Rate = $k[NO]^2$

2nd order w/ respect to NO; 0 order w/ respect to H_2

⑤ a) $A = \varepsilon b c$

$(.600) = (5.00 \cdot 10^3\ cm^{-1} M^{-1})(1.00\ cm)(c)$ $\quad c = 1.20 \cdot 10^{-4}\ M$

b) 1st order

$\ln[A]_t - \ln[A]_0 = -kt$

$\ln(4.0 \cdot 10^{-5}) - \ln(1.20 \cdot 10^{-4}) = -k(35.00\ min)$

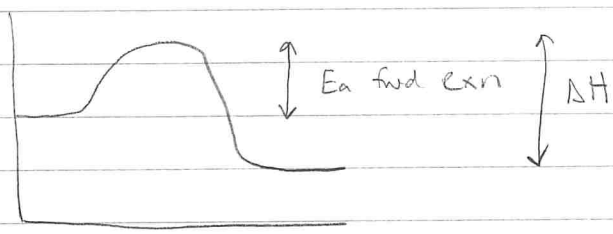
$k = .0314\ \frac{1}{min}$

c) $\ln(1.50 \cdot 10^{-5}) - \ln(1.20 \cdot 10^{-4}) = -(.0314)(t)$

$t = 66.2\ min$

d) $t_{1/2} = \frac{.693}{k} = \frac{.693}{.0314} = 22.1\ min$

⑥ a)



b) i) Rate = $k[A]^1$

ii) determine slope of $\ln[A]$ vs t graph $k = -$slope

⑦ Constant $t_{1/2}$ indicates 1st order rxn. 0 order + 2nd order half-lives are not constant

⑧ 2nd order

$k = .543\ M^{-1} s^{-1}$

$[NO_2]_0 = 1.260\ M$

$[NO_2]_t = .500\ M$

$t = ?$

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

$$\frac{1}{.500\ M} - \frac{1}{1.260\ M} = (.543)(t)$$

$t = 2.22\ s$

(9) 1st order

$[H_2O_2]_0 = .800M$

$[H_2O_2]_t = .175M$

$t = 54$ min

$t_{1/2} = ?$

$\ln[A]_t - \ln[A]_0 = -kt$

$\ln(.175) - \ln(.800) = -k(54)$

$k = .0281 \text{ min}^{-1}$

$t_{1/2} = \frac{.693}{k} = \frac{.693}{.0281} = 24.6$ min

(10)

$NO + Br_2 \rightleftharpoons NOBr_2$

$NOBr_2 + NO \rightarrow 2NOBr$

$2NO + Br_2 \rightarrow 2NOBr$

Rate $= k[NOBr_2][NO]$

$NOBr_2$ is intermediate

$Rate_f = Rate_r$

use 1st step.

$k_f[NO][Br_2] = k_r[NOBr_2]$

$\frac{k_f[NO][Br_2]}{k_r} = [NOBr_2]$

Plug into rate law

$Rate = k\left(\frac{k_f}{k_r}[NO][Br_2]\right)[NO]$

$= k[NO]^2[Br_2]$

(11)

$NO + O_3 \rightarrow NO_2 + O_2$

$NO_2 + O \rightarrow NO + O_2$

$O_3 + O \rightarrow 2O_2$

overall eqn

intermediate: NO_2

catalyst: NO