Chapter #14 – Chemical Kinetics

1. Use the following table to calculate the rate of disappearance for X_2 at different times.

$$X_2 + 3Y \rightarrow Y_3X$$

Do NOT get this confused with rate = K[stuff]...

This is specifically disappearance

Time (s)	$[X_2]$
0.00 s	1.30 M
15.0 s	1.10 M
30.0 s	0.80 M
60.0 s	0.20 M
120.0 s	~~0.50 M~~ 0.050 M

b

a

$$\text{Rate} = \frac{-\Delta[\text{Reactants}]}{\Delta \text{time}}$$

, where Δ means final - initial of that quantity

2 SF(s)

a. 0 s → 120 s

$$\text{Rate} = \frac{-[(0.050\text{M}) - (1.30\text{M})]}{(120\text{s}) - (0.00\text{s})} = 0.012 \frac{\text{M}}{\text{s}}$$ Molarity per second

b. 0 s → 60 s

$$\text{Rate} = \frac{-[(0.20\text{M}) - (1.30\text{M})]}{(60.0\text{s}) - (0.00\text{s})} = 0.018 \frac{\text{M}}{\text{s}}$$

c. 15 s → 60 s

$$\text{Rate} = \frac{-[(0.20\text{M}) - (1.10\text{M})]}{(60.0\text{s}) - (15.0\text{s})} = 0.020 \frac{\text{M}}{\text{s}}$$

$X_2 + 3Y$

$2X + 3Y$

$\frac{2X}{3} = \frac{3Y}{3}$

$Y = \frac{2}{3}X$

d. What would the rate of disappearance be for Y at these time intervals?

The rate of disappearance for Y would be $\frac{2}{3}X$ the rate of disappearance for X_2.

2. Give the $t_{1/2}$ equation. Half Life, scary.

$$t_{1/2} = \frac{\ln(2)}{K} \quad \text{or} \quad t_{1/2} = \frac{0.693}{K}$$

, where K is the rate constant.

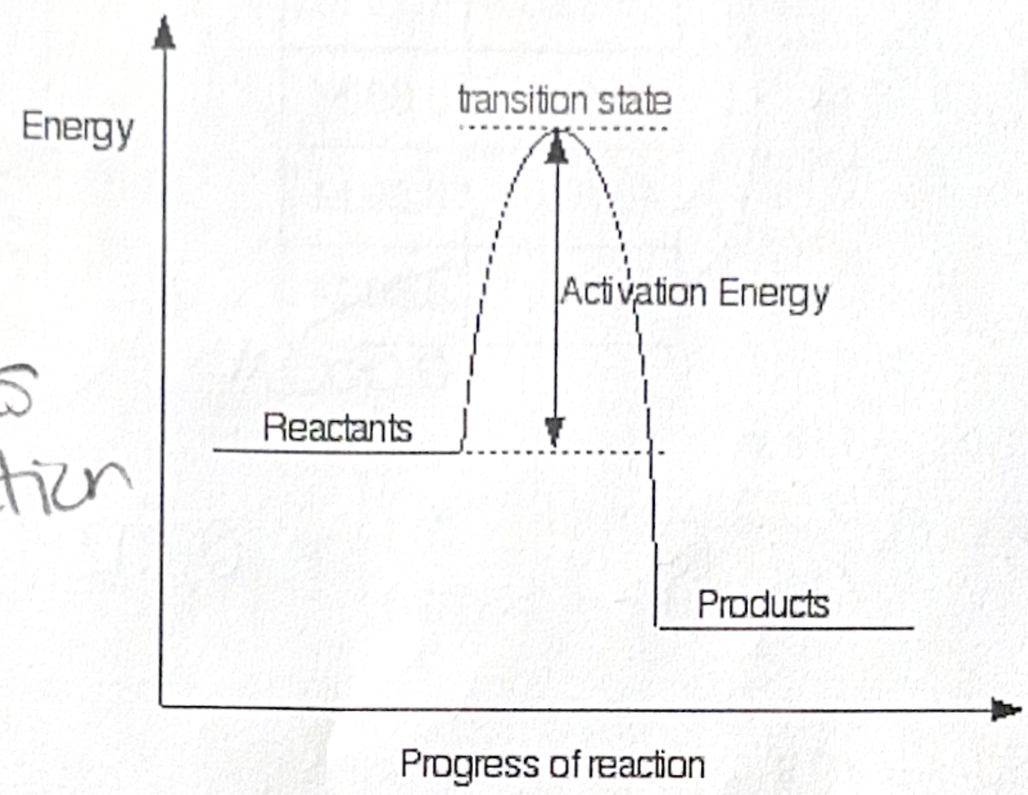
3. What effect does temperature have on the rate constant (k)?

Temperature and the magnitude of the rate constant, K, are DIRECTLY proportional.

4. How does the activation energy affect the rate constant (k)?

A high activation energy results in a low rate constant; it takes MORE energy for the rxn to proceed

5. Label the following graph:



The calculus is Integration

$\int$

1st order reaction: natural log order

rate = $k[A]$ → calculus → $\ln[A]_t = \ln[A]_0 - kt$

from this, a plot of $\ln[A]$ vs. time makes a straight line if the rxn is 1st order.

2nd order reaction: inverse order

rate = $k[A]^2$ → calculus → $\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$

from this, a plot of $\frac{1}{[A]}$ vs. time makes a straight line if the rxn is 2nd order.

1. Consider the data for hydrogen concentration $[H_2]$, iodine concentration $[I_2]$ and the rate of reaction (moles per liter per second or $mol \cdot L^{-1}/s$) for this reaction:

$$H_2\ (g) + I_2\ (g) \rightarrow 2\ HI\ (g)$$

Trial	$[H_2]$ (mol/L)	$[I_2]$ (mol/L)	Rate ($mol \cdot L^{-1}/s$)
1	0.01	0.05	0.04
×2	×1	×2	
2	0.02	0.05	0.08
3	0.05	0.01	0.02
×1	×3	×27	
4	0.05	0.03	0.54

$2^n = 2$

$n = 1$

So, $[H_2]^1$

$3^n = 27$

$n = 3$

So, $[I_2]^3$

What is the overall reaction order?

rate = k[Stuff]...

rate = $k[H_2]^1[I_2]^3$, add 1 and 3;

The overall order is 4.

2. Consider a hypothetical reaction:

$$A + B \rightarrow C + D$$

1 Doubling the concentration of A causes the reaction rate to increase by a factor of four. This is done while the concentration of B is held constant. Tripling the concentration of B, while the concentration of A is held constant, causes the reaction rate to increase by a factor of nine. What is the rate law expression for this reaction?

2

2A = 4 rate ; 3B = 9 rate

$2n = 4$; $3n = 9$

$n = 2$; $n = 2$

rate = $k[A]^2[B]^2$

3. Consider the reaction:

$$NO_2^- (aq) + NH_4^+ (aq) \rightarrow N_2 (g) + 2\ H_2O (l)$$

Trial	$[NO_2^-]$ (mol/L)	$[NH_4^+]$ (mol/L)	Rate (mol·L-1/s x 10^{-7})
1	0.0100	0.200	5.4
2	0.0200	0.200	10.8
3	0.0400	0.200	21.6
4	0.200	0.200	10.8
5	0.200	0.400	21.6
6	0.200	0.600	32.4

×2 ×1 ×2

×2 ×2

$2^n = 2$
$n = 1$

$2^n = 2$
$n = 1$

Determine the rate law expression for this reaction:

So, rate $= k[NO_2^- (aq)][NH_4^+ (aq)]$

Use trial #1 :

$$\frac{(5.4\ \frac{M}{s})}{(0.0100M)(0.200M)} = \frac{k(0.0100\ M)(0.200\ M)}{(0.0100M)(0.200\ M)}$$

4. In a reaction involving only one reactant, A, the rate of the reaction increases by a factor of 27 when the concentration of A is tripled. What is the rate law expression for this reaction?