

1. Please identify the following statements as either true or false:

- a) T Adding solute to a pure solvent widens the temperature range at which the solution is liquid. $\downarrow FP, \uparrow BP$
- b) F The units of the rate constant, k , are the same for all order reactions. $\frac{0}{M/s} \quad \frac{1}{s} \quad \frac{2}{M \cdot s}$
- c) F The graph of reactant concentration vs. time is linear for all order reactions.
- d) T Radioactive decay is always a first-order process.
- e) T For first-order reactions, the reaction half-life is always $t_{1/2} = \frac{\ln 2}{k}$. \rightarrow only for 0 order rxns
- f) F The order of a reaction with respect to a particular reactant is always that reactant's stoichiometric coefficient as written in the reaction.
- g) F The concentration of a catalyst can never appear in the rate equation.
- h) T A catalyst can increase the rate at which reactants are converted to products.
- i) F A catalyst can make a reaction more product favored. \rightarrow catalysts don't influence rxn equilibrium

2. For the following reaction at 856 °C:



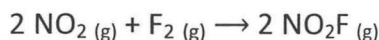
the average rate of disappearance of NH_3 over the time period from $t = 0 \text{ s}$ to $t = 4186 \text{ s}$ is found to be $1.50 \times 10^{-6} \text{ M s}^{-1}$. The average rate of formation of H_2 over the same time period is:

$$\text{avg rxn rate} = -\frac{1}{2} \frac{\Delta[\text{NH}_3]}{\Delta t} = \frac{\Delta[\text{H}_2]}{\Delta t} = \frac{1}{3} \frac{\Delta[\text{H}_2]}{\Delta t}$$

$$\frac{\Delta[\text{NH}_3]}{\Delta t} = -1.50 \times 10^{-6} \frac{M}{s}$$

$$-\frac{1}{2} (-1.50 \times 10^{-6} M/s) = \frac{1}{3} \frac{\Delta[\text{H}_2]}{\Delta t}$$

3. Write an expression for the reaction rate law and calculate the value of the rate constant, k based on the following data. What is the overall order of the reaction?



[NO ₂] (M)	[F ₂] (M)	Initial Rate (M/s)
0.100	0.100	0.026
0.200	0.100	0.051
0.200	0.200	0.103
0.400	0.400	0.411

first order for NO₂ $\times 2$

$$\frac{\Delta[\text{H}_2]}{\Delta t} = 2.25 \times 10^{-6} \frac{M}{s}$$

$\times 2$ first order for F₂

$$\text{rxn rate} = k [\text{NO}_2] [\text{F}_2]$$

$$k = \frac{\text{rxn rate}}{[\text{NO}_2] [\text{F}_2]} = \frac{0.026 \text{ M/s}}{[0.100 \text{ M}] [0.100 \text{ M}]} = 2.6 \frac{1}{M \cdot s}$$

Second order rxn!

have to ask yourself for which values would the graph be linear? In this case $1/[NOBr]$ vs.

4. Please use the table below to determine the order of NOBr decomposition and the value of k: t is linear,

Time (s)	[NOBr] (M)	$\ln[NOBr]$	$1/[NOBr]$
10	0.50	-0.693	2.0
20	0.33	-1.11	3.0
30	0.25	-1.39	4.0
40	0.20	-1.60	5.0

so this is a second order rxn

For second order rxn, $k = \text{slope}$, b/c

$$\frac{1}{[NOBr]_t} = kt + \frac{1}{[NOBr]_0} \quad \text{slope} = \frac{\text{rise}}{\text{run}} = \frac{5.0 - 4.0}{40 - 30} = \frac{1}{10}, \text{ so } k = 0.10 \text{ M}^{-1}\text{s}^{-1}$$

5. The decomposition of XY is second order in XY and has a rate constant of $7.02 \times 10^{-3} \text{ l/M}\cdot\text{s}$ at a certain temperature:

a) How long will it take for the concentration of XY to decrease to 12.5% of its initial concentration when the initial concentration is 0.100 M?

$$\frac{1}{[XY]_t} = kt + \frac{1}{[XY]_0}$$

$$[XY]_0 = 0.100 \text{ M}$$

$$[XY]_t = 0.100 \text{ M} \times 0.125 = 0.0125 \text{ M}$$

$$k = 0.00702 \text{ M}^{-1}\text{s}^{-1}$$

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

$$\frac{1}{0.0125} - \frac{1}{0.100} = 0.00702 t \quad 70 = 0.00702 t \quad t = 9.97 \times 10^3 \text{ seconds}$$

b) How long will it take for the concentration of XY to decrease to 12.5% of its initial concentration when the initial concentration is 0.200 M?

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

$$[XY]_0 = 0.200 \text{ M}$$

$$[XY]_t = 0.200 \times 0.125 = 0.0250 \text{ M}$$

$$k = 0.00702 \text{ M}^{-1}\text{s}^{-1}$$

$$\frac{1}{0.0250} - \frac{1}{0.200} = 0.00702 t$$

$$35 = 0.00702 t \quad t = 4.99 \times 10^3 \text{ sec}$$

c) If this were a first order reaction, how would your calculations differ? Please explain briefly.

If it were first order, the initial concentrations would not have mattered - it would have taken the same amount of time to reach 12.5% of the initial amount.

d) If the initial concentration of XY is 0.052 M, what is the concentration of XY after 64 s?

$$\frac{1}{[XY]_t} = kt + \frac{1}{[XY]_0}$$

$$\frac{1}{[XY]_t} = (0.00702 \text{ M}^{-1}\text{s}^{-1})(64 \text{ s}) + \frac{1}{0.052 \text{ M}}$$

$$\frac{1}{[XY]_t} = 19.7$$

$$[XY]_t = 0.051 \text{ M}$$