Rate Laws Review #2 Key

1. The initial rate of the reaction of nitrogen monoxide and oxygen was measured at 25°C for various initial concentrations of NO and O_2. Determine the rate equation from these data. What is the value of the rate constant?

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

Experiment	[NO]	[O_2]	Initial Rate (mol/L*s)
1	0.02	0.01	0.028
2	0.02	0.02	0.057
3	0.02	0.04	0.114
4	0.04	0.02	0.227
5	0.01	0.02	0.014

$.057 = k(.02)^x(.02)^y$ / $.028 = k(.02)^x(.01)^y$

$2 = 2^y$

$y = 1$

$.227 = k(.04)^y(.02)$ / $.014 = k(.01)^x(.02)$

$16 = 4^x$

$x = 2$

$.114 = k(.02)^2(.04)$

$k = 7125$

$\text{Rate} = 7125[NO]^2[O_2]$

2. The decomposition of SO_2Cl_2 is a first order reaction:

$$SO_2Cl_2(g) \rightarrow SO_2(g) + Cl_2(g)$$

The rate constant for the reaction is 2.8×10^{-3} min^{-1} at 600K. If the initial concentration of SO_2Cl_2 is 1.24×10^{-3} mol/L, how long will it take for the concentration to drop to 0.31×10^{-3} mol/L?

$\ln(.31\times10^{-3}) = -2.8\times10^{-3}\,t + \ln(1.24\times10^{-3})$

$t = 495.1$ min

3. The reaction $2HOF(g) \rightarrow 2HF(g) + O2(g)$ occurs at 25°C.

a.)Using data in the following table, determine the rate law and then calculate the rate constant.

b.) What will be the concentration of HOF after 120 minutes?

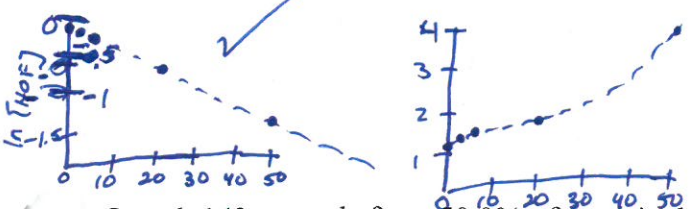
c.) What is the half-life of the reaction?

a.) $\ln(.243) = -k(50) + \ln(.85)$

$k = .025$

$\text{Rate} = .025[HOF]^1$

[HOF]	ln [HOF]	1/[HOF]	Time (min)
0.850	-.1625	1.18	0
0.810	-.211	1.23	2
0.754	-.282	1.33	5
0.526	-.642	1.9	20
0.243	-1.415	4.12	50



b.) $\ln[HOF] = -.025(120) + \ln(.85)$

$[HOF] = .042$ M

c.) $t_{1/2} = \frac{.693}{.025} = 27.72$ min

. It took 143 seconds for a 50.0% of a particular substance to decompose. If the initial concentration was 0.060 M and the decomposition reaction follows second-order kinetics, what is the value of the rate constant?

$t_{1/2} = 143$ sec. $= \frac{1}{k(.06)}$

$k = .117$