Mastery Questions (Due June 11th by the END of class):

For the following questions, you may need physical constants that aren't listed. They can be found either in your book or on the internet. Or, ask and we'll find them together.

Mastery #1: The following data was collected for the reaction of nitrogen gas and chlorine gas to form nitrogen trichloride gas.

Experiment	Initial Nitrogen (atm)	Initial Chlorine (atm)	Initial Rate (atm/min)	Temp (K)
1	0.250	0.250	0.012	300 K
2	0.500	0.500	0.047	300 K
3	0.500	0.250	0.011	300 K
4	0.250	0.500	0.072	500 K

0.666 atm of nitrogen gas and 0.333 atm of chlorine gas are put in a 2 L flask at 400 K. What is the initial rate of the reaction?

Solution:

1st I need the rate law.

Doubling chlorine caused the rate to quadruple. 2nd order.

Doubling nitrogen caused essentially no change in the rate. 0th order.

$$Rate = k[Cl_2]^2$$

This means I need the 2nd order integrated rate law, but that equation

$$\frac{1}{[Cl_2]} = kt + \frac{1}{[Cl_2]_0}$$

Requires that I know the rate constant.

Now, I can get the rate constants – at both 300 K and 501 K

$$Rate = k[Cl_2]^2$$
$$0.012\ atm/min = k_1[0.250\ atm]^2$$
$$k_1 = 0.192\frac{1}{atm\ min}$$
$$0.047\ atm/min = k_2[0.50\ atm]^2$$
$$k_2 = 0.188\frac{1}{atm\ min}$$
$$0.011\ atm/min = k_3[0.250\ atm]^2$$
$$k_3 = 0.176\frac{1}{atm\ min}$$

$$k_{avg} = 0.185 \frac{1}{atm\ min}\ at\ 300\ K$$

$$.072\ atm/min = k_4[0.50\ atm]^2$$
$$k_4 = 0.288 \frac{1}{atm\ min}\ at\ 500\ K$$

The question asks about 400 K not 300 or 501. So, Arrhenius:

The 300 and 501 data gets me the activation energy. Then I can get k at any old temperature that I want.

$$\ln\left(\frac{k_2}{k_1}\right) = -\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

$$\ln\left(\frac{0.288}{0.185}\right) = -\frac{E_a}{8.314 \frac{J}{mol\ K}}\left(\frac{1}{500\ K} - \frac{1}{300\ K}\right)$$

$$0.4426 = -\frac{E_a}{8.314 \frac{J}{mol\ K}}(-1.33 \times 10^{-3} K^{-1})$$
$$E_a = 2760 \frac{J}{mol}$$

$$\ln\left(\frac{k_2}{k_1}\right) = -\frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

$$\ln\left(\frac{0.288}{k_{400}}\right) = -\frac{2760 \frac{J}{mol}}{8.314 \frac{J}{mol\ K}}\left(\frac{1}{501\ K} - \frac{1}{400\ K}\right)$$

$$\ln\left(\frac{0.288}{k_{400}}\right) = -\frac{2760 \frac{J}{mol}}{8.314 \frac{J}{mol\ K}}(-5 \times 10^{-4} K^{-1}) = 0.16598$$

$$e^{\ln\left(\frac{0.288}{k_{400}}\right)} = e^{0.16598}$$

$$\left(\frac{0.288}{k_{400}}\right) = 1.18$$

$$k_{400} = 0.244\ min^{-1}$$

$$Rate = k[Cl_2]^2$$

$$Rate = (0.244\ min^{-1})[0.333\ atm]^2$$

$$Rate = 0.0271\frac{atm}{min}$$

Mastery #2: The following data was collected for the reaction of nitrogen gas and chlorine gas to form nitrogen trichloride gas.

Experiment	Initial Nitrogen (atm)	Initial Chlorine (atm)	Initial Rate (atm/min)	Temp (K)
1	0.250	0.250	0.012	300 K
2	0.500	0.500	0.047	300 K
3	0.500	0.250	0.011	300 K
4	0.250	0.500	0.072	501 K

0.666 atm of nitrogen gas and 0.333 atm of chlorine gas are put in a 2 L flask at 400 K. How much nitrogen trichloride would be produced after 10 minutes?

SOLUTION:

It says “10 minutes”. That’s a time reference. All time references imply “integrated rate law”.

1^{st} I need the rate law.

Doubling chlorine caused the rate to quadruple. 2^{nd} order.

Doubling nitrogen caused essentially no change in the rate. 0^{th} order.

$$Rate = k[Cl_2]^2$$

This means I need the 2^{nd} order integrated rate law, but that equation

$$\frac{1}{[Cl_2]} = kt + \frac{1}{[Cl_2]_0}$$

Requires that I know the rate constant.

Now, I can get the rate constants – at both 300 K and 501 K

$$Rate = k[Cl_2]^2$$
$$0.012\ atm/min = k_1[0.250\ atm]^2$$
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$$k_3 = 0.176\frac{1}{atm\ min}$$

$$k_{avg} = 0.185\frac{1}{atm\ min}\ at\ 300\ K$$

$$.072\ atm/min = k_4[0.50\ atm]^2$$
$$k_4 = 0.288\frac{1}{atm\ min}\ at\ 501\ K$$

The question asks about 400 K not 300 or 501. So, Arrhenius:

The 300 and 501 data gets me the activation energy. Then I can get k at any old temperature that I want.

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$$\ln\left(\frac{0.288}{0.185}\right) = -\frac{E_a}{8.314\frac{J}{mol\ K}}(\frac{1}{501\ K} - \frac{1}{300\ K})$$

$$0.4426 = -\frac{E_a}{8.314\frac{J}{mol\ K}}(-1.33 \times 10^{-3}K^{-1})$$
$$E_a = 2760\frac{J}{mol}$$

$$\ln\left(\frac{k_2}{k_1}\right) = -\frac{E_a}{R}(\frac{1}{T_2} - \frac{1}{T_1})$$

$$\ln\left(\frac{0.288}{k_{400}}\right) = -\frac{2760\frac{J}{mol}}{8.314\frac{J}{mol\,K}}\left(\frac{1}{501\,K} - \frac{1}{400\,K}\right)$$

$$\ln\left(\frac{0.288}{k_{400}}\right) = -\frac{2760\frac{J}{mol}}{8.314\frac{J}{mol\,K}}(-5 \times 10^{-4}K^{-1}) = 0.16598$$

$$e^{\ln\left(\frac{0.288}{k_{400}}\right)} = e^{0.16598}$$

$$\left(\frac{0.288}{k_{400}}\right) = 1.18$$

$$k_{400} = 0.244\ min^{-1}$$

Now, I can use my integrated rate law:

$$\frac{1}{[Cl_2]} = kt + \frac{1}{[Cl_2]_0}$$

10 minutes later, this will tell me how much chlorine is left:

$$\frac{1}{[Cl_2]} = (0.244\ \mathrm{min}^{-1})(10\ \mathrm{min}) + \frac{1}{0.333}$$

$$[Cl_2] = 0.134\ atm\ LEFT$$

The question asks about how much NCl_3 I made. That is stoichiometrically related to the amount of Cl_2 REACTED.

The balanced equation is:

$$N_2(g) + 3Cl_2(g) = 2NCl_3(g)$$

$$0.333\ atm\ initial - 0.134\ atm\ leftover = 0.199\ atm\ reacted$$

I can interpret the stoichiometry as mole ratios – as usual – or as atm rations (P is directly proportional to moles)

$$0.199\ atm\ Cl_2 reacted \frac{2\ atm\ NCl_3}{3\ atm\ Cl_2} = 0.133\ atm\ NCl_3 made$$