Kinetics I HW

Key and clues

#1

For N2O5, use the values in the table.

Remember rate = Δ[N2O5]/time

(1.7×10^{-4} M/sec)

For NO2, use the rate for N2O5 and compare using the mole:mole ratio.

3.4×10^{-4} M/sec

#2

a. If two lines mirror one another (a reactant and a product), they must have the same coeffecient.

Reactant concentrations should be decreasing while product concentrations should be increasing.

The amount which concentration increases or decreases is related to the coeffecient.

b. Equilibrium is reached when reactant and product concentrations are unchanging (not necessarily equal, but steady). Figure out what time this occurs on this graph.

c. **Instantaneous rate** is the rate at a particular moment, and determined by a tangent line. Be sure to draw on the graph as part of your explanation.

d. In order to determine reaction rate, you need one of two types of information

1. Run the experiment multiple times in which you change the concentration and monitor the rate.
2. Measure the concentration of one species over time.

In this case, the first method would be easiest. You could double (or triple or half or whatever) the concentration of one species (leave the other one constant) and see what happens to the rate. Repeat using the other reactant. You could then have a table of [ ] and rate that you could use to determine the order. (Look at the table on #4, you could run experiments to make something like this).

#3

a. Look at exponent
b. Look at exponent
c. Add the orders for the two reactants together.
d. Look at table in notes for order and rate constant units

#4

a. Use table to find order for each one.

b. Add orders together

c. 1.2×10^{-10} 1/M*min

If you are not getting c correct, you will want to check your math and then make sure you have a and b correct.

#5

a. Use table to find order for ClO2 (OH must remain constant)
b. Use table to find order for OH (ClO2 must remain constant)
c. Use orders as exponents
d. Add orders together
e. 230 $1/M^2$*sec

#6

Work similar to 4 and 5

e. 3.41 1/M*sec

#7

Notice we have different data here. Now the data is concentration and time NOT concentration and rate like in 4-6.

a. Remember the equation for rate that we used in #1.
b. Remember how to compare rates like we did in #1.
c. Figure out what order the reaction is (given in the introductory sentence). Use the integrated rate law for this order to find the rate constant.

K= – 3.59 x 10^{-3} 1/min

d. Now that you know k, use that same equation again, but plug in k and your two concentrations of interest in d.

433 minutes

e. Use the half-life equation that we developed for first order reactions.

#8

a. Note the order of the reaction given in the introductory sentence. Use this to determine the units on the y axis.
b. Use our chart of order and units on k.

#9

a. Which chart gives you a straight line? Look at the units on y on this chart. This allows you to determine the order for the reactant.
b. Use the order you got from a as you exponent in the rate law.
c. Use the integrated rate law for second order. Take numbers from the graphs and plug in.

k = 0.17 1/M*sec

d. Using the k value and the initial concentration from the graph and the time 250 sec to determine. M= 0.015 M

#10

Very similar to #9

c. 0.011 1/sec

d. 63 sec

e. 0.0078 M

#11

Graph the data 3 different ways and see which one yields a straight line. This will reveal the order

#12

13 mol

#13

60 minutes