Chapter 13 reading guide. Tro 3rd ed

Be active while reading the text. Take notes; think about what you've read, and ask yourself questions while reading. Use this document as a guide for making your reading more meaningful.

Section 13.2, Rates of Chemical Reactions

How are rates of chemical reactions similar to rates measured for cars on a highway?

What kind of data would be needed to calculate a rate of a reaction based on eq 13.1?

Calculate the average rate of the slow reaction in figure 13.1, in terms of the formation of Z. (1 Z/16 s, or 3 Z per 46 s, or about 0.063 Z/s)

Can you use stoichiometry to determine the relative rate of HI to H₂ in the example equations in section 13.2?

View the graph in figure 13.2. Why are the two curved lines different in their direction and in their overall gain/loss? At what time of the reaction does the figure show the instantaneous rate calculation? How is this different from the average rate?

Briefly look at equation 13.4 and 13.5, and then erase all memory of 13.5 from your mind. Seriously. Ignore how the author works out example 13.1. Instead, use chem 201A-style stoichiometry via the balanced chemical equation to work from the rate of change of one substance [ in mol/(L*s) ] to rate of another substance. Just like we did in class.

Solve For Practice 13.1 and check your answers.

Skim over the portion on measuring reaction rates. We will do something just like what is shown in figure 13.3 during our first lab to measure the absorbance (and thus concentration) of a reactant changing with time.

Section 13.3, The Rate Law

The general form of a “rate law” is shown in 13.6. What are the most common values for ‘n’ and what is the order of a reaction if n = 3 in eq 13.6?

What relationship does the rate law describe between the rate of a reaction and the concentration of a reactant, assuming that n is larger than 1.0?

In a zero order reaction, does the reaction rate change as the concentration of a reactant changes?

In a first or second order reaction, does the reaction rate change as the concentration of a reactant changes?

View figure 13.5. How do the plots of [A] vs time vary for zero, first, and second order?

View figure 13.6. How do the plots of reaction rate vs [A] vary for zero, first, and second order? Why does one of the rates not change with [A], while the other rates change as [A] decreases?

There is a subsection titled “Determining the Order of a Reaction,” and the sentence below this heading should be big, bold, blue, and big. It's important. After reading this sentence, do you think that reaction orders and a rate law can be determined merely by looking at the balanced chemical equation for a reaction?

The very next sentence is important too. There are two experimental ways to determine the order of a reaction (and a 3rd which is a cheap shortcut).
What might MOIR refer too, and what two measurements are needed for MOIR?

The text does not do a great job of explaining how to solve problems with MOIR. Check your class notes for the full solution. There are a few nice little tips in the margin of page 605, and you practiced these two math skills on your Chem 201B math review worksheet.

Example 13.2 is not very good for an intro to this topic, but if you like mental math then it might work for you.

Section 13.4, The Integrated Rate Law

The rate law is great, but it does not tell you anything about concentrations at specific times. Through the magic of calculus, we can take two expressions of rate (the one from section 13.2 and the one from the rate law), set them equal to each other, and crank out one new equation. If you’ve had integral calculus, you should be able to derive these equations. If you haven’t had calculus, don’t panic! We just need the result, as shown in eqs 13.12 and 13.13.

What form of the 1st order integrated rate law has the form of an equation of a straight line? What are the values of the slope and y intercept? What would a person graph to get a straight line out of data that followed 1st order kinetics?

Follow Examples 13.3 and 13.4, then solve For Practice 13.3 and 13.4, and check your answers.

The 2nd order integrated rate law is shown in eq 13.15. What are the values of the slope and y intercept? What would a person graph to get a straight line out of data that followed 2nd order kinetics?

Follow Example 13.5, then solve For Practice 13.5, and check your answer.

Repeat the above for zero order reactions.

The half-life of a reactant is a useful concept, especially for 1st order reactions. How does the half-life of a 1st order reaction change as the concentration of the reactant changes? What equation can you use to calculate half-life?

Look at the graph on p 613 (Figure 13.11). If half of the reactant disappears after the first 100 s, why isn’t the reaction over, with no reactant left, after 200 s?

How does the half-life of a 2nd order reaction change as the concentration of the reactant changes? What equation can you use to calculate half-life?

There is a great bulleted summary of basic kinetic relationships on p 614, so check it out. Even better is Table 13.2, the rate law summary. You should know what graph gives a straight line for a reactant following certain kinetics, the equation for the integrated rate law, the units for k, and the generic rate law for 0th, 1st, and 2nd order reactions.

That’s the first week of kinetics in less than two pages! This is what reading a textbook in a science class entails, doing this kind of work to get something out of the book. Often, students will say that they don’t learn
anything by reading the book, but you need to put effort into reading. Learning won't just pop out at you at this level. I have had so many students say that once they start reading, everything makes more sense, they have more fun in class, and learn more.

Try to read a bit before class, even if you don't think you're getting much out of that time. Just learning the words and direction of the material will help once you get to lecture.

**Section 13.5: How does temperature effect reaction rates?**

We discussed in class which part of a rate law changes with temperature, and this is noted in the text also.

The Arrhenius eq (13.24) describes this relationship. In class, we worked through what happens to k if T is increased, or if Ea is increased. You should do that on your own again to refresh your memory. It is discussed further in the “The Exponential Factor” portion of 13.5. Can you determine how the exponential factor changes if T is increased, or if Ea is lowered?

Note that there is a brand new R (a different form of your old friend the ideal gas constant). Memorize the value and units. We will use the new R, and the old one, quite a bit this term. I call this new R the ‘energy’ R and the old one the ‘gas law’ R. It’s good to know which version to use in a particular problem.

Examine the figures 13.12 and 13.13, and be able to reproduce or interpret similar figures, including the terms shown and the units of the axis.

After the figures, there is a key phrase regarding rates of reactions. What is the only energetic factor that controls the rate of a chemical reaction?

Figure 13.4 shows an interesting relationship between temperature (there are two shown) and Eₐ. How does Eₐ change with temperature?

Equation 13.24 isn’t all that useful, but 13.26 and 13.27 are. Note how 13.26 is a linearized form of 13.24, and how one can calculate Eₐ from the slope of a graph with LN k vs.1/T. This calculation is worked out in Example 13.7.

Equation 13.27 is derived from 13.26, shown on p 619. Work through Example 13.8 and For Practice 13.8; these are similar to ones we did in class.

We discussed the collision model in class. Note how the frequency factor (A) in 13.24 is split into two values.

**Section 13.6: Which path do reactions take?**

What are elementary steps?

Is it possible to deduce a mechanism from the overall balanced chemical equation?

What is a rate-determining step?

Why can’t the reaction of NO₂ + CO be a process where one NO₂ combines with a CO in a single step? Be able to draw or decipher a graph like figure 13.16.

Reactions mechanisms with a first slow step are pretty straightforward. Why is this?
The bottom line with reactions with an initial fast step is that every reactant prior to the slow step ends up in the rate law.

Students should be able to determine a mechanism from a reaction coordinate diagram, or choose the correct mechanism from a few choices given the observed kinetics.

**Section 13.7: What is a catalyst?**

Examine figure 13.17. In terms of thermodynamics, how does a catalyst increase a reaction rate?

What kind of impact have catalytic converters had on air pollutants?

In class, we discussed several ways that catalysts work to reduce activation energy. How can $E_a$ be reduced?

Can you draw a reaction coordinate diagram that shows the progression of reactants to products for a catalyzed and an uncatalyzed reaction?