Purpose: This is a guide for your as you work through the chapter. The major topics are provided so that you can write notes on each topic and work the corresponding problems.

This should serve a s a study guide as you go on to do the problems in Sapling and take the quizzes and exams.

The Problems are embedded in the Topics and Space for Notes

12.1 Define rate.

- Define chemical reaction rate

- Derive rate expressions from the balanced equation for a given chemical reaction

- Calculate reaction rates from experimental data

Remember: for the reaction: aA + bB → cC + dD.. the relative rates are:

$$\text{rate} = -\frac{1}{a}\frac{\Delta[A]}{\Delta t} = -\frac{1}{b}\frac{\Delta[B]}{\Delta t} = \frac{1}{c}\frac{\Delta[C]}{\Delta t} = \frac{1}{d}\frac{\Delta[D]}{\Delta t}$$

Rate = $\frac{\Delta Concentration}{\Delta time}$ **(always +)**

1. Write a balanced reaction for which the following rate relationships are true.

$$\text{Rate} = -\frac{1}{2}\frac{\Delta[N_2O_5]}{\Delta t} = \frac{1}{4}\frac{\Delta[NO_2]}{\Delta t} = \frac{\Delta[O_2]}{\Delta t}$$

A) $2\ N_2O_5 \rightarrow 4\ NO_2 + O_2$
B) $4\ NO_2 + O_2 \rightarrow 2\ N_2O_5$
C) $2\ N_2O_5 \rightarrow NO_2 + 4\ O_2$
D) $\frac{1}{4}\ NO_2 + O_2 \rightarrow \frac{1}{2}\ N_2O_5$
E) $\frac{1}{2}\ N_2O_5 \rightarrow \frac{1}{4}\ NO_2 + O_2$

2. Given the following balanced equation, determine the rate of reaction with respect to $[SO_3]$. If the rate of O_2 loss is 3.56×10^{-3} M/s, what is the rate of formation of SO_3?

$2\ SO_2(g) + O_2(g) \rightarrow 2\ SO_3(g)$

12.2 and 12.3 Rate Law

Describe the effects of chemical nature, physical state, temperature, concentration, and catalysis on reaction rates

Define Rate Law

What effects the rate of the reaction?

Explain the form and function of a rate law

Remember: Concentration dependence.

For the reaction $aA + bB \rightarrow cC$, Rate = $k[A]^x[B]^y$ (Note: $x \neq a$, $y \neq b$) Units of k are $1/(M^{\text{overall order-1}}\ \text{time})$

3. A reaction is first order in A and second order in B.

(A) Write the rate law.

(B) What will happen to the rate if the concentration of A is doubled (B is held constant)?

(C) What will happen to the rate if the concentration of B is doubled (A is held constant)?

How do you use rate laws to calculate reaction rates?

How do you use rate and concentration data to identify reaction orders and derive rate laws?

4. (Rate Law) The following data were collected for the reaction of: A + B $\rightarrow$ products(g)

Run	Initial [A]	Initial [B]	Initial Rate M/s
1	0.100	0.050	.0400
2	0.200	0.050	.0800
3	0.100	0.100	.160

(A) Find the Rate Law.

Rate = ______________________________

(B) Calculate k (don't forget the units.)

12.4 Integrated Rate Law:

- **Identify the order of a reaction from concentration/time data**
- **Integrated Rate Law: (time dependence) 0,1st and 2nd order.**
- **If given $t_{1/2}$ calculate k – make sure units of k and t match.**

Order	Rate Law	Concentration-Time Equation	Half-Life
0	rate = k	$[A] = [A]_0 - kt$	$t_{½} = \frac{[A]_0}{2k}$
1	rate = k [A]	$\ln[A] = \ln[A]_0 - kt$	$t_{½} = \frac{\ln 2}{k}$
2	rate = k $[A]^2$	$\frac{1}{[A]} = \frac{1}{[A]_0} + kt$	$t_{½} = \frac{1}{k[A]_0}$

5. What data should be plotted to show that experimental concentration data fits a second-order reaction?
For example: a straight line plot of concentration of the reaction versus time gives you a zero order reaction with the slope of the line equal to -k.

x axis	y axis	slope	reaction
time	[reactant] or [A]	-k	zero order
			Second order
			First order

Be able to use the integraded Rate Law: (time dependence) 0,1st and 2nd order.

Write the integrated rate laws for first, second and zero order.

6. A sample of wood from an ancient fire has been dated using ^{14}C dating. Given that the $t_{1/2}$ of carbon-14 is 5730 years, a fresh sample of ^{14}C has a decay rate of 15.3 $\frac{disintegrations}{gram\ C \cdot min}$ and that the decay rate of the wood from the fire is 2.05$\frac{disintegrations}{gram\ C \cdot min}$, how old is the wood?

7. The second-order decomposition of NO_2 has a rate constant of 0.255 $M^{-1}s^{-1}$. How much NO_2 **remains after** 4.00 s if the initial concentration of NO_2 (1.00 L volume) is 1.33 M?

12.5 Collision Theory

- Use the postulates of collision theory to explain the effects of physical state, temperature, and concentration on reaction rates

  Explain collision theory. What happens if you increase the temperature?

- Define the concepts of activation energy and transition state

  Draw an energy diagram. Label reactants, products, transition state, activation energy for the forward reaction, activation energy for the reverse reaction. Also include if the reaction is endothermic or exothermic.

- Use the Arrhenius equation in calculations relating rate constants to temperature

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

Define k and E_a

Arrhenius Equation (Temperature dependence)

8. (A first order reaction has an Activation Energy of 35.5 kJ/mole. If the rate constant is 0.0021 s^{-1} at 35 °C, what is the value for the rate constant at 55 °C?

12.6 and 12.7 Mechanism and Catalysis.

- Distinguish net reactions from elementary reactions (steps)
- Identify the molecularity of elementary reactions

- Write a balanced chemical equation for a process given its reaction mechanism
- Derive the rate law consistent with a given reaction mechanism
- Explain the function of a catalyst in terms of reaction mechanisms and potential energy diagrams

Mechanism: What is occurring on the microscopic level!

Rules!!

Write a rate law for each elementary step.
For fast first step rate forward = rate reverse.
Overall rate comes from the slow step.
If there is an equilibrium, set rate forward = rate reverse
Catalysts – speed up reactions by providing an alternate mechanism with a lower activation energy.

Define:
Elementary step

Intermediate

Equilibrium

Overall rate

Rate determining step

Catalyst (both homogeneous and heterogeneous)

9. The reaction of $2O_3(g) \rightarrow 3O_2(g)$ has the experimental rate: rate = $k[O_3]^2 [O_2]^{-1}$.

The following mechanism has been proposed:

1. $O_3 \rightleftharpoons O_2 + O$ fast, equilibrium
2. $O + O_3 \rightarrow 2\,O_2$ slow

(A) Identify any intermediates.

(B) What is the rate law predicted by this mechanism?

(C) Is the rate law predicted by the mechanism consistent with the experimental rate law?

(D) How would adding a catalyst effect this reaction?

10. Read up on the Ozone hole. Also consider watching this video. What is the mechanism for the destruction of the ozone layer?