

CHEM 1212
Principles of Chemistry II
Course Study Guide

Introduction

This study guide has been provided to assist you in learning the material that you need to master in CHEM 1212. It should not be used solely without consultation with your textbook and your notes. In addition, we strongly advise you to avoid looking at the answers until you have diligently and thoroughly worked through a problem to the best of your ability, using both your notes and the textbook as a resource. This will allow you to work through the learning process as you figure out the answer or the method of problem solving on your own. If you choose to instead try to learn how to solve the problem by looking the answer, you are seldom learning much and of course, when you get to the test and don't have the answer key as a crutch, you likely won't be able to come up with a way to solve the problem. We have found that successful students work through these problems consistently throughout the semester, not waiting to do them a few days (or worse, the night before) the test. They will not help you if those are the conditions in which you practice these problems.

We hope this study guide helps you in your attempt to master CHEM 1212. Please report any typos or errors that you find, we will be most appreciative.

Dr. Spencer and Dr. Myers

Study Skills and Suggestions

There are lots of ideas out there and some are organized here. Most important:

- 1) You must study and you must practice to be successful.
- 2) This is college and now you are responsible to do this regardless of whether or not specifics are assigned by the professor. If the professor does assign homework and such, this is very unlikely to be sufficient.
- 3) More is not always better. **Quality** counts as much as quantity—sometimes even more.

My most recent recommendations for how to study come from a book called *A Mind For Numbers* by Barbara Oakely.

Here is her summary:

Ten Rules of Good Studying:

1. Use recall. After you read a page, look away and recall the main ideas. Highlight very little, and never highlight anything you haven't put in your mind first by recalling. Try recalling main ideas when you are walking to class or in a different room from where you originally learned it. An ability to recall—to generate ideas from inside yourself—is one of the key indicators of good learning.

2. Test yourself. On everything. All the time. Flash cards are your friend.

3. Chunk your problems. Chunking is understanding and practice with a problem solution so that it can all come to mind in a flash. After you solve a problem, rehearse it. Make sure you can solve it cold—every step. Pretend it's a song and learn to ply it over and over again in your mind, so the information combines into one smooth chunk you can pull up whenever you want.

4. Space repetition. Spread out your learning in any subject a little every day, just like an athlete. Your brain is like a muscle—it can handle only a limited amount of exercise on one subject at a time.

5. Alternate different problem-solving techniques during your practice. Never practice too long at any one session using only one problem-solving technique—after a while; you are just mimicking what you did on the previous problem. Mix it up and work on different types of problems. This teaches you both how and when to use a technique. (Books are generally not set up this way, so you'll need to do this on your own.) After every assignment and test, go over your errors, make sure you understand why you made them, and then rework your solutions. To study most effectively, handwrite (DON'T type) a problem on one side of a flash card and the solution on the other. (Handwriting builds stronger neural structures in memory than typing.) You might also want to photograph the card if you want to load it into a study app on your smartphone. Quiz yourself randomly on different types of problems. Another way to do this is to randomly flip through your book, pick out a problem, and see whether you can solve it cold.

6. Take breaks. It is common to be unable to solve problems or figure out concepts in math or science the first time you encounter them. This is why a little study every day is much better than a lot of studying all at once. When you get frustrated with a math or science problem, take a break so that another part of your mind can take over and work in the background.

7. Use explanatory questioning and simple analogies. Whenever you are struggling with a concept, this to yourself, *How can I explain this so that a ten-year-old could understand it?* Using an analogy really helps, like saying that the flow of electricity is like the flow of water. Don't just think your explanation—say it out loud or put it in writing. The additional effort of speaking and writing allows you to more deeply encode (that is, convert into neural memory structures) what you are learning.

8. Focus. Turn off all interrupting beeps and alarms on your phone and computer, and then turn on a timer for twenty-five minutes. Focus intently for those twenty-five minutes and try to work as diligently as you can. After the timer goes off, give yourself a small, fun reward. A few of these sessions in a day can really move your studies forward. Try to set up times and places where studying—not glancing at your computer or phone—is just something you naturally do.

9. Eat your frogs first. Do the hardest thing earliest in the day, when you are fresh.

10. Make a mental contrast. Imagine where you've come from and contrast that with the dream of where your studies will take you. Post a picture or words in your workspace to remind you of your dream. Look at that when you find your motivation lagging. This work will pay off both for you and those you love!

Ten Rules of Bad Studying.

1. Passive rereading—sitting passively and running your eyes back over a page. Unless you can *prove* that the material is moving into your brain by recalling the main ideas without looking at the page, rereading is a waste of time.

2. Letting the highlights overwhelm you. Highlighting your text can fool your mind into thinking you are putting something in your brain when all you're really doing is moving your hand. A little highlighting here and there is okay—sometimes it can be helpful in flagging important points. But if you are using highlighting as a memory tool, make sure that what you mark is also going into your brain.

3. Merely glancing at a problem's solution and thinking you know how to do it. This is one of the worst errors students make while studying. You need to be able to solve a problem step-by-step, without looking at the solution.

4. Waiting until the last minute to study. Would you cram at the last minute if you were practicing for a track meet? Your brain is like a muscle—it can handle only a limited amount of exercise on one subject at a time.

5. Repeatedly solving problems of the same type that you already know how to solve. If you just sit around solving similar problems during your practice, you're not actually preparing for a test—it's like preparing for a big basketball game by just practicing your dribbling.

6. Letting study session with friends turn into chat sessions. Checking your problems solving with friend, and quizzing one another on what you know, can make learning more enjoyable, expose flaws in your thinking, and deepen your learning. But if your joint study sessions turn to fun before the work is done, you're wasting your time and should find another study group.

7. Neglecting to read the textbook before you start working problems. Would you dive into a pool before you knew how to swim? The textbook is your swimming instructor—it guides you toward the answer. You will flounder and waste your time if you don't bother to read it. Before you begin to read, however, take a quick glance over the chapter or section to get a sense of what it's about.

8. Not checking with your instructors or classmates to clear up points of confusion. Professors are used to lost students coming in for guidance—it's our job to help you. The students we worry about are the ones who don't come in. Don't be one of those students.

9. Thinking you can learn deeply when you are being constantly distracted. Every tiny pull toward an instant message or conversation means you have less brain power to devote to learning. Every tug of interrupted attention pulls out tiny neural roots before they can grow.

10. Not getting enough sleep. Your brain pieces together problem-solving techniques when you sleep, and it also practices and repeats whatever you put in mind before you go to sleep. Prolonged fatigue allows toxins to build up in the brain that disrupt the neural connections you need to think quickly and well. If you don't get a good sleep before a test, **NOTHING ELSE YOU HAVE DONE WILL MATTER.**

Practical aspects to consider putting into your life this semester:

#1—Practice Retrieval

“Knowledge” of basic facts is the basis of all further learning. Students often think that knowledge means having something in your head. Actually, it's being able to get something out of your head on demand!!! If you think of your brain as a big file room, the easiest things to find will be where you have worn a path to by going to get it over and over.

Retrieval is not:

- Reading your text or notes (however many times you do it!)
- Looking up the answer in your text or notes
- Reading someone else's solution of the problem
- Once and done; once learned you do not have to review

Retrieval is:

- Flashcards—if you actually try to answer the question without looking
- Attempting practice problems without reference to notes or resources
- Closing your eyes when you answer a question
- Use spaced repetition*
- Taking quizzes....even mock quizzes your generate for each other in a study group
- Explaining to someone else how to get an answer in your own words

Retrieval is what you will be expected to do on tests. Why wouldn't you practice it beforehand?

**Spaced repetition* means taking breaks between practice, increasing the break every time, so that you spend days or even a week before you try it again. Also do your work in different spaces, it's amazing how much of our memory is tied to locational clues. *Spaced repetition is critical to actually learning something.*

You have a list of things to memorize. That is a perfect thing to practice your retrieval skills on!

#2—Use your textbook wisely

A textbook is not a novel. Reading it like one is not an effective study technique. Think of it more like a treasure hunt for information.

Reading the textbook before class:

Don't expect to understand much, but the guided questions provided in this study guide (called previews) should help locate key words/topics that you will need to know. Note key words (you know, the ones in bold and/or italics!) and watch for them in class. Look at the pictures, so when your instructor attempts a drawing you know what it is supposed to look like. Some students don't like reading it before class...that's ok, try afterwards for review and clarification of concepts that were fuzzy in class.

Using your textbook during class:

In this class...don't. We will not use the textbook in class. Save your back.

Soon after class:

Use your text to proofread your notes. Did you get things written correctly? Did you miss something important? Is there a conflict between the notes and the text? Use their pictures and graphs to better understand your own. (The text is right unless the instructor specifically said otherwise, but if you are

paranoid, you can double-check with the instructor.) Note what sections of the text you covered, you will need that for the next step.

Before the next class:

Work on practice problems. The best thing about any textbook is all the practice problems it provides. You noted the sections you covered so: 1) In the text of that section there are example problems. Try to work them without looking at the provided solution. There is generally a second one where the solution is in the back of the book. 2) Look at the end of chapter problems. They are divided into sections that correspond with sections in the book. Do the problems associated with section you are studying. Check your answers with the back of the book for the bold problems. If you need more explanation, there are solution manuals on reserve at the library. ONLY use these to check problems you have already attempted. Reading the solutions will not help (See study tip #1). 3) Do not forget the “visual problems.” These are “conceptual” like your final and many test questions. 4) If you have the access code, you can also do the electronic homework and get immediate feedback.

#3—Problem-solving is not about the answer

The phrase “show all work” indicates that you are explaining how to get the answer, the answer itself is a minor part of the procedure. If a question requires significant problem-solving, it will likely be worth a lot of points (8 – 12 on a test) but only one of those points will be for the actual answer. The good news: this means you can get partial credit! To earn all these points, you need to show your work clearly. Remember, it’s not practice if you don’t do it the same way as you would do it on the test. So get in the habit of writing all this stuff out (you are also explaining it to yourself and/or your study buddies) and use the same calculator you will be using on the exam!

If you are solving the problem algebraically: Start with the equation you will be solving/using. Use standard variables whenever possible. For the variables for which there is a numeric value, relate that number (with units and appropriate significant figures!) to the variable. You can do this with an equality or by putting the number in *exactly the same spot* it is on the equation. On a separate line, show the algebraic manipulation, use multiple lines if needed. You should finish with the variable equal to your final answer, double-check for significant figures and units and circle it so that the instructor knows that is what you mean to report.

For example: Density = mass/volume = m/V

$$13.6 \text{ g/mL} = 7.42 \text{ g/V}$$

$$(13.6 \text{ g/mL})V = 7.42 \text{ g}$$

$$V = 7.42\text{g}/13.6 \text{ g/mL}$$

$$V = 0.545 \text{ mL}$$

If you are solving the problem with dimensional analysis (this is the preferred method): Start, on the left, at a logical point, which is NOT one of the conversion factors. Make sure that all conversions have units and the units cancel top to bottom. Do not cross through the units (this impedes the ability of the professor to read your work). Circle your final answer making sure the units and significant figures are correct. You may want to write conversions off to the side until you know if you need to use them. Values in the problem that have two units divided by a slash are likely meant to be a conversion factor rather than your starting point. Rewrite them as conversion factors. For example: 13.6 g/mL can be $13.6 \text{ g} = 1 \text{ mL}$; 2.456 m/s can be $2.456 \text{ m} = 1 \text{ s}$; \$4.99/lb can be $\$4.99 = 1 \text{ lb}$.

For example:

$$0.729 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mL}}{13.6 \text{ g}} \times \frac{1 \text{ cm}^3}{1 \text{ mL}} = 54.5 \text{ cm}^3$$

#4—Test Preparation

Most of your learning should already be done. The day before the test is too late to memorize or understand ideas that will be tested. Remember, tests are not about knowing information, they are about retrieving and USING information. Chemistry tests in particular want you to take information and use it in a new situation and maybe combine it with other ideas you should know. Still there are things that you *can* do:

A) Organize the information

It is jumbled around in your head right now. What is it you need to know? How are those ideas related to each other? Some ideas for organizing:

- Make lists
 - Of topics that will appear
 - Of equations and definitions
 - Of the big ideas and even some of the smaller ones

- Make a concept map

It looks like a spider web. You put a main or central concept in the middle and link to related concepts and/or definitions. Next to the links, use single words or short phrases to describe the relationships. See this link for an example. <http://defiant.corban.edu/jjohnson/pages/bioi/02/chemconmapframe.html>

- Make a “cheat sheet”

What if you were allowed to look at one page of stuff, what would you put on it? How would you organize it to find the information you need? Do it! Just because you can't use it on the test does not mean that it is not useful. These are great tools for a last minute review. **NONE OF THESE IDEAS WORK UNLESS YOU MAKE YOUR OWN BY HAND. COMPUTER CUT AND PASTE IS NOT THE SAME.** Using someone else's is totally useless.

B) Practice

You know that everything is easier with practice. Practice as close to the actual circumstances as you can.

- Find practice tests (I do not provide them, but Goggle is your friend here)
- Make your own practice tests: Remember to mix up concepts. Questions do not have the nice labels that the homework at the end of the chapter does. Try “problems from hat”: Copy homework problems (mine and book and internet), cut away all identifying information and pull them at random to work.
- Set time limits

You only have 50 minutes to complete the exam, practice with these limits. Normally no more than 1 minute for a multiple choice question, 1-2 minutes for fill in the blank (depending on the type of question) and 5-10 minutes for a problem.

- Do it like the test

Not only the time limits, but use the same calculator, pencil, periodic table, etc. Even do a practice in the same room if possible.

C) Come prepared

You will need a fully functional brain to be successful.

- Get a good night's sleep.

It is dangerous to drive when sleep deprived. It is also dangerous to take a test.

- Eat breakfast

Your brain needs fuel. Some caffeine is good. Too much is distracting...jitters and a potty break eat into your test-taking time.

- Bring the right materials.

You will need a pencil with a good eraser (NOT A PEN!) and the appropriate calculator. Maybe bring a back-up or two. You will not be allowed any other items.

#5—Interweave topics of study/spaced repetition

Chemistry is naturally cumulative. You need to know one topic to understand the next. That means you cannot learn material then mentally dump it until the final exam. If it doesn't come back in a future chapter, it will return in CHEM 1212 or your next science class. For example, if the current study of naming, you need to know from previous material: element names and symbols, parts of the periodic table whether referred to by section name or group name, how ions are made and predict the charges...just as a start. Naming is one of the more difficult topics because of all the skills it requires. This means that you want to keep reviewing previous topics. Sneak them in as you do the new work. Part of this goes back to the file cabinet analogy. If you want to use information from your brain's filing cabinets, it's useful not only to have a well-worn path there, but many paths to that filing cabinet. You should be able to find it from any direction.

The good news: This interweaving of new and old topics has been shown to be a more effective study tool. Students who do mass studying (studying the same thing for a long time) and students who interweave (break up new studies with review of old) or do the same amount of studying but in several shorter sessions perform the same on an immediate test. HOWEVER, if they are tested a month later, the students who have done interweaved studying do substantially (at least a letter grade) better.

So as you study, go back and test yourself on previous material. If you are making a practice test for your study buddies, add a question from the first test. If you are doing "problems from a hat" make sure that there are problems from all chapters, not just the one you are currently learning.

Do your studying in many short sessions rather than one long one. Remember, your tests are only 50 minutes long, if you can't get a lot of things done in that amount of time, you will not have time to be successful on the test either. HOWEVER, you still should average a minimum of 2 hours of real study for every hour you spend in class. Those hours should occur between each class so that you have learned the previous material and are ready to learn new things. You will have to schedule some extra time before a test, but not as much if you stay caught up and current on ALL the topics that will be on the test. Test prep will be for review and organizing, not learning. This is as is should be.

#6—Set Goals

There is a lot of research pointing to the "study with purpose" or "mindful studying". Essentially, it means the studier must set goals for each study session. It's the act of setting the goals that puts the brain (and the material) into an organized state of reference and context. Plus, there is more personal satisfaction when the studier knows when success has happened... the goal has been reached. Example goals (not a complete list):

- Complete 5-10 problems in 25 minutes, without looking at the answers.
- Review flashcards, getting over 80% correct
- Create a study sheet
- Make a practice quiz or test for a friend...pass each other's tests
- Make a list of all the important topics since the last exam
- Rank them in difficulty
- Do 10 problems of the most difficult type, 9 of the next most difficult, etc.
- Make flashcards for all the important vocabulary terms
- Use an online resource to find example problems...do the problems
- Explain a periodic trend or big idea
- List ideas from the first exam needed for the second exam, add some of those to your practice problems

#7 – Explain yourself

Every teacher knows that the best way to learn something is to have to teach it. You can take advantage of this by teaching yourself or your study group. YES, you must say things out loud and/or write it for this to work!

As you prepare your “lesson” consider the things you want in a good teacher. You want to know: 1) What is it you are trying to do? (in other words: What is your purpose and goal?) 2) How are you going to accomplish your goal? (What tools? What steps?) 3) Present your problem-solving method in an orderly fashion, explaining how each step leads to the next. Orderly means start at the upper left and go left to right and/or top to bottom; not skipping even “obvious” steps. 4) Evaluate your answer. What does it mean? Does it make sense? 5) What is the significance of your lesson? What does solving this tell you about the world? Consider both the atomic perspective (what are those atoms/molecules/electrons doing?) and macroscopic perspective (what can we see with our own eyes). 6) How does it connect to other topics in this class and to other classes (Biology? Math? Psychology?). It is not to torture you that most homework problems ask you to “explain.” It is intended to make you think about your answer.

#8—Don’t Brain Dump...constantly review

Chemistry (like many subjects) is cumulative knowledge and often taught in a “spiral.” Thus, after covering a topic, you circle back to it and look at it again in a deeper way. Just because you have been tested over a topic does not mean that you can forget it. Now is the time of year when this becomes very obvious.

For example, our first test will include intermolecular forces. To successfully evaluate which substance would have a higher melting point, you need to know from CHEM 1211: 1) Is the substance ionic or covalent? 2) What are the ions that make up an ionic substance? 3) What is the Lewis structure of a covalent substance? 4) As Lewis structure predicts electronic geometry you need that to determine if the substance is polar or nonpolar. After answering these OLD questions, you have enough information to move on to the new ones. If you can’t answer these questions, you will not get the new one correct.

Not only will the new material continue to draw on old knowledge, but also recognize that: 1) Your final exam is cumulative for the entire term. 2) CHEM 1212 is a continuation of this course, so it is expected that you will use and build on your CHEM 1211 content. 3) Lab will use class information and class may use lab information. It should not matter where it is first presented. 4) Advanced Chemistry courses constantly refer back to CHEM 1211 and 1212 knowledge. 5) OTHER science courses consistently use ideas from chemistry courses! Why else do you think it is a required part of the major and often a prerequisite course?

What should you do?

- Review old tests and homework.
 - Make sure that you know the correct answer to every question.
 - Identify whatever misconception caused you to miss it the first time.
- Randomly add old questions to new practice problems.
- Note which “old” skills are required for new problems.
 - If you do not have these old skills, go back and acquire them.
- Look for chemistry skills/knowledge appearing in other classes.
 - This will give you more practice in a different context.
- Notice which ideas and skills keep recurring.
 - That is a hint that these are important.

#9—Keep your purpose in mind

What is your ultimate goal? Picture yourself as a success. (maybe even literally, stage a photoshoot then put the picture of the successful you somewhere you will see it when you study.) There will always be obstacles to your goal. Your only choices are to overcome or turn around.

But why chemistry? Why are the powers-that-be making you take this course? Some reasons:

1) Chemistry is the central science. Its principles show up in all the other science courses. Notice (particularly in biology) how many times ideas that we have learned in chemistry are used to explain a biological process. Most of the interesting research is on the cutting edge of chemistry and something else.

2) Chemistry develops useful habits of mind.

a. *Paying attention to detail.* I know you find getting all the details right frustrating. (Really NA, is not good enough? Answer: NA IS TOTALLY WRONG!) However, details are important. Where you put the decimal on a drug prescription is the difference between medicine and poison. Should I amputate the right or left leg? (Doesn't matter, they look the same, right?) A or B is only one slip of the mouse when entering grades.

b. *Show your work.* Communication skills are highly valued by all employers. Knowing how to communicate your thoughts in a context-appropriate way is invaluable. Someday you will have to convince your patient/customer/boss that your conclusion is correct; can you lead them through your thought process convincingly? Besides, any teacher will tell you that teaching is the best way to learn.

c. *Problem solving and critical thinking.* These are higher-order (more difficult) thinking skills that are also valued by employers. It's not enough to memorize the answer to a familiar problem, you need to be able to apply your learned skills to new problems. You need to be able to look at other people's ideas and decide if they are useful or harmful. You need to question information and not assume every Nigerian is a prince with a secret bank account ☺.

Important information from CHEM 1211 that you still need to know (and there's more, just didn't make it to this list!)

1) Names and symbols of the following elements

Al	Aluminum	Mg	Magnesium
Ar	Argon	Mn	Manganese
Ba	Barium	Hg	Mercury
Be	Beryllium	Ne	Neon
B	Boron	Ni	Nickel
Br	Bromine	N	Nitrogen
Ca	Calcium	O	Oxygen
C	Carbon	P	Phosphorus
Cl	Chlorine	Pt	Platinum
Cr	Chromium	K	Potassium
Co	Cobalt	Sc	Scandium
Cu	Copper	Si	Silicon
F	Fluorine	Ag	Silver
Au	Gold	Na	Sodium
He	Helium	S	Sulfur
H	Hydrogen	Ti	Titanium
I	Iodine	U	Uranium
Fe	Iron	V	Vanadium
Pb	Lead	Zn	Zinc
Li	Lithium		

2) Table SI prefix multipliers and symbols

(common Mega (M) to nano (n))

Table 1.1

Mass, length, volume, time, temperature, quantity of substance, energy

3) Formulas for

- 1) density
- 2) percent
- 3) temperature conversion between C and K

4) Atomic/Subatomic Particles

- Alpha α or ${}^4\alpha$ or ${}^4\text{He}$ or ${}^4_2\alpha$
- Beta β or β^- or ${}^0_{-1}\beta$
- Gamma γ or ${}^0_0\gamma$
- Positron β^+ or ${}^0_1\beta$
- Proton 1_1p
- Neutron 1_0n
- Electron ${}^0_{-1}e$
- Deuterium 2_1H or 2_1D
- Tritium 3_1H or 3_1T

5) Groups on the periodic table

Metals

Transition metals

Alkaline earth metals

Nonmetals

Noble gases

Halogens

Metalloids

Alkali metals

6) Polyatomic Ions:

$C_2H_3O_2^-$	acetate	SO_4^{2-}	sulfate
NH_4^+	ammonium	Hg_2^{2+}	mercury(I)
CO_3^{2-}	carbonate	SO_3^{2-}	sulfite
HCO_3^-	bicarbonate	HSO_4^-	bisulfate
ClO_4^-	perchlorate	HSO_3^-	bisulfite
ClO_3^-	chlorate	NO_3^-	nitrate
ClO_2^-	chlorite	NO_2^-	nitrite
ClO^-	hypochlorite	CrO_4^{2-}	chromate
PO_4^{3-}	phosphate	$Cr_2O_7^{2-}$	dichromate
CN^-	cyanide	SCN^-	thiocyanate
OH^-	hydroxide	MnO_4^-	permanganate

7) Basic organic compounds (simple alkanes)

CH_4 = methane

C_2H_6 = ethane

C_3H_8 = propane

C_4H_{10} = butane

C_5H_{12} = pentane

8) Covalent prefixes

Mono (1) to Deca (10)

(table 4.4)

9) Strong acids

HCl = hydrochloric acid

HBr = hydrobromic acid

HI = hydroiodic acid

$HClO_4$ = perchloric acid

HNO_3 = nitric acid

H_2SO_4 = sulfuric acid

$HClO_3$ = chloric acid

10) solubility rules (table 8.3 with some modifications)

a) all compounds containing the following ions are soluble in water:

group 1 ions, ammonium ion, nitrate ion, acetate ion and perchlorate ion

b) compounds containing the following ions are usually soluble, common exceptions are noted

group 17 ions (halides)

unless the cation is: Ag^+ , Cu^+ , Hg_2^{2+} or Pb^{2+} , in which case it is NOT soluble

sulfate ion

unless the cation is: Ba^{2+} , Ca^{2+} , Hg^{2+} , Pb^{2+} or Sr^{2+} , in which case it is NOT soluble

c) $Ca(OH)_2$, $Sr(OH)_2$ and $Ba(OH)_2$ are soluble.

Formula Sheet for General Chemistry:

Formulas with an asterisk* are provided on tests and the final exam. All other formulas must be memorized.

Description	Equation
Density	$D = \frac{m}{V}$
Converting °C to K	$K = ^\circ C + 273.15$
Converting °F to °C	$^\circ C = (^\circ F - 32)/1.8$
Converting °C to °F	$^\circ F = 1.8 (^\circ C) + 32$
Molarity	$M = \frac{\text{moles of solute}}{L \text{ of solution}}$
Dilution of solution	$M_1V_1 = M_2V_2$
Ideal gas equation	$PV = nRT$
Charles' Law	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Boyle's Law	$P_1V_1 = P_2V_2$
Combined gas law	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
Density of ideal gas	$D = \frac{PM}{RT}$
Mole fraction of A	$X_A = \frac{\text{moles of A}}{\text{total moles}}$
Partial pressure of a gas	$P_A = P_{\text{total}}X_A$
Total gas pressure	$P_{\text{total}} = P_A + P_B + P_C + \dots$
Energy of photon	$E = h\nu$
Speed of light	$c = \nu\lambda$
Formal charge on an atom in a Lewis structure	$FC = \# \text{valence electrons} - \# \text{electrons in lone pairs} - \# \text{bonds}$
Molality	$m = \frac{\text{moles solute}}{\text{kg solvent}}$

Freezing point depression	$\Delta T_f = i K_f m$
Boiling point elevation	$\Delta T_b = i K_b m$
Osmotic pressure	$\pi = i M R T$
Raoult's law	$P_1 = X_1 P_1^\circ$
Equilibrium constant	$K_c = \frac{[C]_{eq}^c [D]_{eq}^d}{[A]_{eq}^a [B]_{eq}^b}$ where $aA + bB \rightleftharpoons cC + dD$
Reaction quotient	$Q = \frac{[C]^c [D]^d}{[A]^a [B]^b}$ where $aA + bB \rightleftharpoons cC + dD$
K_p from K_c	$K_p = K_c (RT)^{\Delta n}$
pH definition	$pH = -\log[H_3O^+]$
pOH definition	$pOH = -\log[OH^-]$
pH and pOH relationship	$pH + pOH = 14.00$
Ion product constant for water	$K_w = [H_3O^+][OH^-] = K_a \times K_b = 1.0 \times 10^{-14}$ at 25 °C
Acid ionization constant	$K_a = \frac{[H_3O^+][A^-]}{[HA]}$
Base ionization constant	$K_b = \frac{[OH^-][BH^+]}{[B]}$
pKa definition	$pK_a = -\log(K_a)$
Henderson-Hasselbalch equation*	$pH = pK_a + \log \frac{[A^-]}{[HA]}$
Arrhenius equation	$k = A e^{-\frac{E_a}{RT}}$
Two-point Arrhenius equation*	$\ln \left(\frac{k_2}{k_1} \right) = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$
First-order integrated rate law*	$\ln[A] = \ln[A]_o - kt$
Second-order integrated rate law*	$\frac{1}{[A]} = kt + \frac{1}{[A]_o}$
Zero-order integrated rate law*	$[A] = [A]_o - kt$

Heat equation	$q = mc\Delta T$
First law of thermodynamics*	$\Delta U = q + w$
Enthalpy change	$\Delta H^\circ = \Sigma n\Delta H_f^\circ \text{ products} - \Sigma n\Delta H_f^\circ \text{ reactants}$
Enthalpy change for a reaction from bond energies	$\Delta H_{\text{rxn}} = \Sigma(\Delta H \text{ bonds broken}) + \Sigma(\Delta H \text{ bonds formed})$ (breaking bonds-endergonic; forming bonds-exothermic)
Entropy change	$\Delta S^\circ = \Sigma nS_{\text{products}}^\circ - \Sigma nS_{\text{reactants}}^\circ$
Gibbs Free Energy change	$\Delta G^\circ = \Sigma n\Delta G_f^\circ \text{ products} - \Sigma n\Delta G_f^\circ \text{ reactants}$
Gibbs Free Energy change*	$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$
Gibbs Free Energy change*	$\Delta G^\circ = -RT\ln K$
Gibbs Free Energy change*	$\Delta G^\circ = -nFE^\circ$
Non-standard free energy	$\Delta G = \Delta G^\circ + RT\ln Q$
Electrical current	$I = \frac{\text{charge}}{\text{time}} = \frac{q}{t}$
Standard emf	$E^\circ_{\text{cell}} = E^\circ_{\text{red}} + E^\circ_{\text{ox}}$ where $E^\circ_{\text{ox}} = -(E^\circ_{\text{red}})$
Cell voltage*	$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} \ln Q$

Guided reading questions (by chapter):

Liquids, Solids and Intermolecular Forces

1. What causes a temporary dipole?
2. What are interactions that are based on temporary dipoles called? Are they relatively weak or strong?
3. What is responsible for the boiling point trends in Table 6.2?
4. What are isomers and how do they differ?
5. Define viscosity and how it relates to IMFs.
6. List the type of forces that polar molecules can experience. How do you know which type will happen?
7. What is so special about hydrogen bonds?
8. Define solubility and miscibility. How do you determine if a given solute will dissolve in a given solvent?
9. What are dipole-induced dipole interactions?
10. What does it mean to have competing intermolecular forces?
11. What is stronger-a covalent bond or an intermolecular force?

Or if you prefer something other than the textbook, try this:

<https://www.khanacademy.org/science/organic-chemistry/gen-chem-review/electronegativity-polarity/v/intermolecular-forces-and-molecular-bonds>

Solutions

1. How is solubility determined? What is the rule?
2. Define hydrophilic and hydrophobic.
3. Be able to label the parts of a phase diagram.
4. How is surface tension related to intermolecular forces?
5. How is viscosity related to increasing molar mass (all else equal?) (see Table 6.3)
6. What determines the rate at which molecules evaporate from liquid to gas?
7. What does it mean to achieve dynamic equilibrium?
8. Define vapor pressure and factors that would influence vapor pressure.
9. What is vaporization? What is the opposite of vaporization?
10. What is the relationship between temperature and the average kinetic energy of molecules?
11. Define the terms volatile and nonvolatile.
12. How does your body use the endothermic nature of vaporization for cooling?
13. What is the definition of boiling point?
14. Define homogeneous mixture. (what is another term used to describe it?)
15. Define solute, solvent, solubility and miscibility.
16. Check out Sample exercise 6.2. How is the identity of the solvent determined?
17. What are the three types of interactions that can occur in a solution containing one solute?
18. Why doesn't n-octane dissolve in water?
19. What are dipole-induced dipole interactions? Are they important? What is an example?
20. Define hydrophobic and hydrophilic interactions.
21. What is the general solubility rule?
22. Can a gas dissolve in a liquid? If so, provide an example.
23. What is the relationship stated in Henry's law?
24. What are factors that determine the solubility of a gas in a liquid?
25. What is Raoult's law and what information does it provide?

26. Review calculations of mole fraction if you can't remember.....
27. How does the presence of a solute affect the vapor pressure of a solvent?
28. How does the vapor pressure of a solution depend on the mole fraction of solvent?
29. Define colligative properties.
30. Define molarity and molality and compare these definitions.
31. When 1 mole of NaCl dissolves in water, how many total moles of particles are present?
32. Define nonvolatile solute and explain why vapor pressure must be handled differently when a volatile solute is present in a mixture.
33. Define boiling point elevation.
34. Define freezing point depression.
35. Why is molality used instead of molarity in these calculations?
36. What is the van't Hoff factor and why is it important when discussing the colligative properties?
37. How is the van't Hoff factor related to molecular (covalent) vs. ionic compounds?
38. Define semipermeable membrane and osmosis.
39. From what formula is the formula for osmotic pressure derived?
40. You should be familiar with the terms hypotonic, hypertonic and isotonic as they relate to solutions.
41. Why is drinking seawater harmful?
42. How can colligative properties be used to determine molar mass?

Thermodynamics

1. What is the sum of all the kinetic and potential energies in an object called?
2. Define thermodynamics and its first law.
3. What is the “system” and what are the “surroundings”?
4. Examine Fig. 9.13 and explain what it tells you about temperature changes during phase changes.
5. What is the heat capacity? Is it an extensive or intensive property?
6. How is the specific heat capacity of a substance different than the heat capacity?
7. Which sample will have a greater temperature change with a given amount of heat added: 10 g of water or 20 g of water?
8. What is the **molar** heat capacity? What are its units?
9. Does the process of melting require or release heat?
10. If the temperature of a sample is not changing (during a phase change), then what is the addition of heat “doing?”
11. What is the purpose of calorimetry? Define it.
12. What thermodynamic term is used to describe the magnitude of heat absorbed or released during a chemical reaction?
13. What is this thermodynamic term proportional to?
14. What is a thermochemical equation and what information does it provide?
15. Why might a coffee cup be useful to measure changes in enthalpy?
16. How can you directly calculate ΔH_{rxn} ? How can you indirectly calculate it?
17. What does it mean to be a state function? How does that relate to Hess’s law?
18. What happens to ΔH when we reverse a reaction?
19. What happens to ΔH when a reaction is multiplied (or divided) by some factor?
20. What is the ‘standard’ state as it pertains to standard enthalpies of reaction?
21. Describe a formation reaction. How many moles of product are formed in a formation reaction?
22. What is the standard enthalpy of formation for any substance in its standard state?

Spontaneity and Free Energy

1. What does it mean for a reaction to be spontaneous? Does that mean that the reaction is fast?
2. Is it true that all exothermic reactions are spontaneous? Why or why not?
3. Define entropy as best as you can. What is the second law of thermodynamics?
4. What is the third law of thermodynamics?
5. Define the standard molar entropy.
6. List some factors that affect entropy changes.
7. What factor is responsible for the difference in entropy of Bromine gas and Bromine liquid? What about ethane vs. propane (refer to Table 12.2 in text)?
8. Define the “universe” in thermodynamic terms.
9. How are changes in entropy values for reactions calculated?
10. What is the Gibbs free energy symbol and how is it defined? How is it calculated?
11. Describe the temperature dependence of a spontaneous reaction. (Pay careful attention to Table 12.3 in your text as it is very important)
12. What thermodynamic quantity (variable) do you use to determine whether or not a reaction is spontaneous?

Chemical Kinetics

1. What does the rate of reaction tell you?
2. How is the rate of a reaction measured?
3. What does delta (Δ) mean?
4. What happens to the concentration of product vs. time during a chemical reaction?
5. What happens to the concentration of reactant vs. time during a chemical reaction?
6. How are the rates of change of reactants and products related to each other?
7. In general, why does the rate of a reaction decrease as the reaction progresses?
8. What is the difference between average rate and instantaneous rate?
9. What happens to the rate of a reaction when the temperature decreases?
10. Define activation energy. How is the rate of a reaction related to activation energy?
11. Describe the collision model for chemical reactions.
12. What are the disadvantages to using the method of initial rates to determine the rate law for a reaction?
13. How do we determine the relationship between the rate of a reaction and the concentration of a reactant?
14. How do we determine the relationship between the concentration of a reactant and time?
15. For a first order reaction, what is plotted to yield a straight line?
16. For a second order reaction, what is plotted to yield a straight line?
17. For a zero order reaction, what is plotted to yield a straight line?
18. What is the half-life of a reaction? What is it related to?
19. For what order reaction is the half-life constant? (ie independent of initial concentration)
20. Please make careful note of Figure 13.15 and the thought process required to determine the order of a reaction from concentration and time data.
21. What are the three criteria for a reaction to occur?
22. How is activation energy related to the rate of a reaction?
23. How is the rate of a reaction affected by increasing temperature?
24. Define the parameters of the Arrhenius equation.
25. What is a transition state? How would you describe its energy?

26. How can you determine the activation energy of a reaction experimentally?
27. What is a reaction mechanism?
28. Define a reaction intermediate.
29. Define elementary step and uni, bi or termolecular reactions. Why are termolecular reactions so rare?
30. What must be true of a valid mechanism? Can it ever be proven true?
31. Check out Figure 13.23. Which step is rate-limiting? How do you know?
32. Can a rate law be determined directly from a balanced overall reaction? Why or why not?
33. What is the rate-determining step for a reaction involving people moving through an airport?
34. What is a reversible reaction and what does it mean when the rates of the forward reaction and reverse reaction are equal?

Chemical Equilibrium

1. What does it mean for a reaction to reach equilibrium? What does it mean for a reaction to lie far to the right? To the left?
2. What does it mean for a reaction to be reversible? Are all reactions reversible?
3. Define equilibrium constant. What does it mean if K is very large? Very small?
4. Where do the exponents come from in the equilibrium constant expression?
5. What do you look at to determine the direction of change for a reaction mixture containing R and P that is not at equilibrium?
6. Once the value of Q is known, how is that used to determine the direction of the reaction?
7. What is the value of Q when there are only products in a reaction?
8. What is the value of Q when there are only reactants in a reaction?
9. What happens to a reaction when $Q = K$? Does the reaction stop?
10. What do you know about liquids and solids when it comes to writing equilibrium constant expressions?
11. What does knowledge of an equilibrium constant allow you to determine?
12. What does the acronym RICE stand for?
13. What is the quadratic equation? How many solutions result from the quadratic equation? How do you know which one is the answer?
14. Is there a way to simplify the calculations resulting from a RICE table?
15. List at least 3 different ways to perturb a system at equilibrium.
16. What happens to the pressure in a reaction vessel when the volume is decreased? To what side of the equation does the reaction shift if the pressure of a reaction is increased?
17. When the temperature is changed in a reaction, what changes? Does Q or K change? How does the system respond?
18. How does a catalyst work? Does it affect the equilibrium position of a reaction?

Acids and Bases

1. Draw the Lewis structure of the hydronium ion.
2. According to the Bronsted-Lowry theory, what is an acid? A base?
3. If NH_3 is a base, what is its conjugate acid?
4. How is a strong acid defined? A weak acid?
5. What is K_a and what does it tell us about the strength of an acid?
6. Which is the strongest acid listed in Table 15.2?
7. Is water an acid or a base?
8. What is K_w ?
9. What digits are significant in a logarithm?
10. What does pH tell you?
11. Define strong acids and describe how the concentration of $[\text{H}_3\text{O}^+]$ is determined for a strong acid.
12. Define weak acids and describe how the concentration of $[\text{H}_3\text{O}^+]$ is determined for a weak acid. Which is more abundant strong or weak acids?
13. What is the formula for percent ionization?
14. If you have a solution of 0.10 M strong acid and 0.10 M weak acid, which will have the lowest pH?
15. Is there any way to get out of using the quadratic equation??? (review example 15.5-15.8)
16. Define monoprotic, diprotic and polyprotic acids. Do all the protons come off at once?
17. What does percent ionization tell you?
18. How are weak bases defined? How can you recognize a weak base?
19. Why are dilute acids treated differently than more concentrated ones in terms of calculation of pH?
20. Is the pH of a salt solution neutral? Why or why not?
21. What happens to an ionic compound when it is dissolved in water?
22. What can you say about the relative strength of the conjugate base of a weak acid? What if the acid is strong?
23. What is the product of $K_a \times K_b$?
24. How can you tell if a cation or anion will be acidic or basic?

Aqueous Equilibria

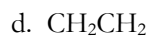
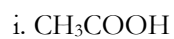
1. What is the so-called common ion effect? If additional conjugate base is present in a solution of a weak acid, is the ionization of that acid greater, less than or equal to what it would be in the absence of the additional conjugate base?
2. Write down the Henderson-Hasselbalch equation. What is it used for??? Are there any limitations of its use?
3. What is a buffer?
4. What is it made up of (specifically)?
5. How does a buffer respond to the addition of added acid or base? Are there limitations on a buffer's effectiveness? What are they?
6. What is a pH indicator and what role does it play in a titration?
7. What are the steps of a titration? What is a titration used for?
8. How does a titration curve differ based on the starting acid and base used in the titration?
9. What is the meaning of the equivalence point? What species is present in an acid/base titration at the equivalence point?
10. For titrations involving weak acids and bases, what is the significance of the $\frac{1}{2}$ equivalence point?

Electrochemistry

1. Define oxidation and reduction.
2. What is the oxidation number for an atom in a pure element? What is the oxidation number of oxygen in its compounds? Are there exceptions? What is the oxidation number of hydrogen in its compounds? Are there exceptions?
3. How does the oxidation number change in a redox reaction for a species that is oxidized? Reduced?
4. Define oxidizing agent and reducing agent.
5. What is the basis for balancing redox reaction equations?
6. Why are balancing redox reactions more complicated than other types of reactions? What is the special method for balancing redox reactions? (note: we will only balance redox in acid)
7. What is a device called that produces electricity through chemical reactions?
8. In the Zn half-cell, what keeps the positive charge from building up? What causes the positive charge anyway?
9. What is the driving force for electrical current flow?
10. What is the potential difference between the two electrodes in a cell called? How is a large difference in charge between the two electrodes generated?
11. What are standard cell potentials and what are the conditions of a standard cell?
12. Where does oxidation occur in an electrochemical cell? Where does reduction occur?
13. What is the role of a salt bridge in an electrochemical cell?
14. How is a cell diagram used to describe an electrochemical cell? How is an electrode like a platinum electrode dealt with in the cell notation?
15. How are standard reduction potentials defined? What does it mean to have a positive standard reduction potential? What does it mean to have a negative standard reduction potential?
16. For a spontaneous reaction, what is the sign of E_{cell}° ?
17. How can you determine if a redox reaction is spontaneous?
18. What is the relationship between E_{cell} and ΔG° ?
19. How are standard reduction potentials determined for half-reactions?
20. Does the concentration of species in a redox half-reaction affect E_{cell} ? How?
21. Describe the Nernst equation and when it is used. How is it derived? Does this equation look familiar to you?

Review 1

1. Draw Lewis structures of the following and determine if they are polar or nonpolar. Then list the type(s) of intermolecular forces present in each.



2. a) From the molecules in #1, list those that have polar bonds but are nonpolar.
b) From the molecules in #1, list those that do not have polar bonds but are polar.

3. Draw 2 possible resonance structures for NO_2^- . Assign formal charge for each atom in the polyatomic ion and based on formal charges, pick the “best” structure. If the structures are equivalent, indicate so.

4. Which of the following statements is TRUE of the resonance structures of NO_2^- ?

- A) The nitrite ion contains one N-O single bond and one N=O double bond.
- B) The nitrite ion contains two N-O bonds that each has a bond order of 1.5.
- C) The nitrite ion contains two N=O double bonds.
- D) The nitrite ion contains two N-O single bonds.
- E) None of the above are true.

5. Draw 3 resonance forms for the nitrate ion (NO_3^-). What can you say about the length of each N-O bond?

6. Choose the bond below that is **most** polar.

- A) C-N B) C-F C) C-O D) C-C E) F-F

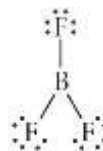
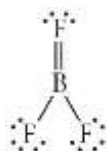
7. Choose the bond below that is **most** polar.

- A) H-I B) H-Br C) H-F D) H-Cl E) C-H

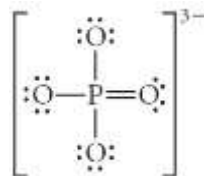
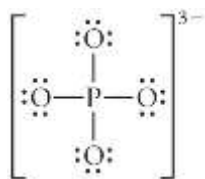
8. Choose the bond below that is **least** polar.

- A) P-F B) C-Br C) C-F D) C-I E) C-Cl

9. From the following two resonance structures for BF_3 , choose the **best** structure based on the assignment of formal charge to each atom.



10. Choose the best Lewis structure for PO_4^{3-} .



11. Choose the bond below that is the **strongest**.

- A) C-F B) C=O C) C-I D) I-I E) C≡N

12. Which element can expand its valence shell to accommodate more than eight electrons?

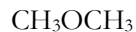
- A) N B) O C) Br D) He E) none of the above

Review 2

1. For each substance below:

a) if it is ionic, write the formulas of the component ions

b) if it is covalent, draw the Lewis structure



_____ 2. Which of the substances above, if any, have hydrogen bonding?

_____ 3. Which of the substances above has the highest vapor pressure?

4. Rank the compounds below from the compound that has the highest melting point to the compound with the lowest melting point.

A CH₃OCH₃

B HClO₃

C AlPO₄

D Ne

E NH₄I

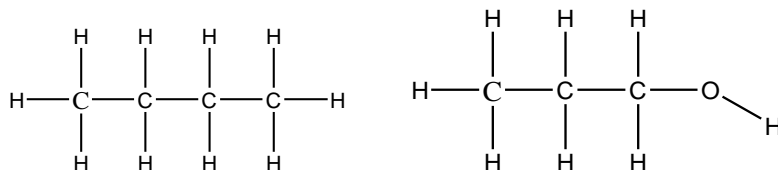
F SO₃

Strongest 1_____ 2_____ 3_____ 4_____ 5_____ 6_____ Weakest

Explain your reasoning below. This should include the identity of the strongest intermolecular force in each compound. Include also Lewis structure or ion formula, as appropriate.

Review 3

1. Which of the following compounds would have the greater viscosity? WHY?



2. Consider ethanol, $\text{CH}_3\text{CH}_2\text{OH}$, and dimethyl ether, CH_3OCH_3 , to answer the following questions.

- What type(s) of IMFs does ethanol exhibit?
- What type(s) of IMFs does dimethyl ether exhibit?
- Which of the two molecules has the stronger IMFs?
- Which of the two molecules has the greatest surface tension?

3. Which of the following substances will have the strongest intermolecular forces?

- a) H_2S b) CO_2 c) CH_3NH_2 d) Cl_2 e) Rn

4. In which of the following would dispersion forces be the **only significant** factors in determining the boiling point?

- I. Ar II. Li_2SO_4 III. CF_4 IV. Br_2 V. NH_3
- a) I, II, and III b) II, IV, and V c) I, III, and IV d) I, IV and V e) II and V

5. Which of the following boils at the **highest** temperature (assume all carbon chains are straight chains)?

- a) C_2H_6 b) C_3H_8 c) C_4H_{10} d) C_5H_{12} e) C_6H_{14}

6. Choose the substance with the highest boiling point.

- a) CH_4 b) KI c) CS_2 d) HF e) I_2

7. Place the following substances in order of **increasing** boiling point.

- $\text{CH}_3\text{CH}_2\text{OH}$ Ar CH_3OCH_3
- $\text{Ar} < \text{CH}_3\text{OCH}_3 < \text{CH}_3\text{CH}_2\text{OH}$
 - $\text{CH}_3\text{CH}_2\text{OH} < \text{Ar} < \text{CH}_3\text{OCH}_3$
 - $\text{CH}_3\text{CH}_2\text{OH} < \text{CH}_3\text{OCH}_3 < \text{Ar}$
 - $\text{CH}_3\text{OCH}_3 < \text{Ar} < \text{CH}_3\text{CH}_2\text{OH}$
 - $\text{Ar} < \text{CH}_3\text{CH}_2\text{OH} < \text{CH}_3\text{OCH}_3$

Review 4

- Molecule A has a molar mass twice that of molecule B.
 - If A and B have the same type of IMF's possible, will A or B evaporate faster if they are at the same temperature? Justify your response.
 - If A and B have the same evaporation rate at room temperature, what does this tell you about their IMF's?
- Consider the following molecules with similar molecular mass: formaldehyde ($\text{H}_2\text{C}=\text{O}$), methanol (CH_3OH), and ethane (CH_3CH_3). Rank these from lowest to highest evaporation rate.
- As the temperature increases, the evaporation rate _____ and the vapor pressure_____.
- A molecule with stronger IMF's has a _____ evaporation rate and a _____ vapor pressure.
- Rank the compounds below from the compound that has the highest vapor pressure at 15°C to the compound with the lowest vapor pressure at 15°C .

A C_8H_{18}

B C_4H_{10}

C PCl_3

D NH_3

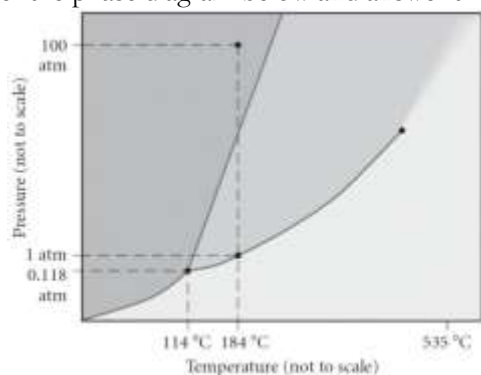
E BrF_5

F HClO_2

Highest 1_____ 2_____ 3_____ 4_____ 5_____ 6_____ Lowest

Explain your reasoning below. This should include the identity of the strongest intermolecular force in each compound.

Consider the phase diagram below and answer the questions that follow:



6. What phases exist at a temperature of 114 °C and a pressure of 0.118 atm? At this temperature, if the pressure is increased, what happens (name phase changes)?
7. Consider the phase diagram above. If the dashed line at 1 atm of pressure is followed from 100 to 500 °C, what phase changes will occur (in order of increasing temperature)?
- condensation, followed by vaporization
 - sublimation, followed by deposition
 - vaporization, followed by deposition
 - fusion, followed by vaporization
 - No phase change will occur under the conditions specified.
8. Which of the following statements is TRUE?
- Vapor pressure increases with temperature.
 - Hydrogen bonds are stronger than covalent bonds.
 - Vapor pressure increases with increasing surface area.
 - Dispersion forces are generally stronger than dipole-dipole forces.
 - None of the above are true.
9. Which substance below has the strongest intermolecular forces?
- A_2X , $\Delta H_{\text{vap}} = 39.6$ kJ/mol
 - BY_2 , $\Delta H_{\text{vap}} = 26.7$ kJ/mol
 - C_3X_2 , $\Delta H_{\text{vap}} = 36.4$ kJ/mol
 - DX_2 , $\Delta H_{\text{vap}} = 23.3$ kJ/mol
 - EY_3 , $\Delta H_{\text{vap}} = 21.5$ kJ/mol
10. Choose the substance with the lowest vapor pressure at a given temperature (note: lone pairs are not included in the drawings below nor are the geometries necessarily correct).
- a) $\begin{array}{c} \text{H} & \text{H} \\ | & | \\ \text{H}-\text{C}-\text{C}-\text{OH} \\ | & | \\ \text{H} & \text{H} \end{array}$
- b) HCl
- c) $\begin{array}{c} \text{H} & \text{H} \\ | & | \\ \text{HO}-\text{C}-\text{C}-\text{OH} \\ | & | \\ \text{H} & \text{H} \end{array}$
- d) $\begin{array}{c} \text{H} & \text{H} & \text{H} \\ | & | & | \\ \text{H}-\text{C}-\text{C}-\text{O}-\text{C}-\text{H} \\ | & | & | \\ \text{H} & \text{H} & \text{H} \end{array}$
- e) $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$

11. The ideal gas law is based on the kinetic molecular theory which doesn't hold true at high pressures and low temperatures. What two assumptions in the kinetic molecular theory are not true at high P and low T?

12. Place the following substances in order of **increasing** vapor pressure at a given temperature.

- NF₃ NH₃ BCl₃
- a) NH₃ < NF₃ < BCl₃ b) NF₃ < NH₃ < BCl₃ c) BCl₃ < NF₃ < NH₃
d) NH₃ < BCl₃ < NF₃ e) BCl₃ < NH₃ < NF₃

Review 5

1. In the examples below, list the strongest type of intermolecular force that can occur in the pure substance or in the mixture.

- a) LiI
- b) CH₃OH
- c) CH₃CH₃
- d) H₂ + H₂O
- e) CH₂F₂
- f) LiI + H₂O

2. Given the following substances, predict whether they will dissolve in the indicated solvent. List the strongest type of intermolecular force that can occur between the solute and solvent.

<i>Solute</i>	<i>Solvent</i>	<i>Dissolve? Yes or no</i>	<i>Strongest IMF</i>
KI	H ₂ O		
KI	CH ₃ CH ₂ CH ₂ CH ₂ CH ₃		
$\begin{array}{c} \text{O} \\ \\ \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CCH}_3 \end{array}$	H ₂ O		
$\begin{array}{c} \text{O} \\ \\ \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CCH}_3 \end{array}$	CH ₃ CH ₂ CH ₂ CH ₂ CH ₃		
NH ₃	H ₂ O		

3. Which of the following statements is TRUE?

- a) The solubility of a solid is not dependent on either temperature or pressure.
- b) The solubility of a solid is highly dependent on pressure.
- c) The solubility of a solid is highly dependent on both pressure and temperature.
- d) The solubility of a solid is highly dependent on temperature.

4. Select all true statements.

- a) The solubility of a gas in water increases with increasing pressure.
- b) The solubility of a gas in water decreases with increasing pressure.
- c) The solubility of a gas in water increases with increasing temperature.
- d) The solubility of a gas in water decreases with increasing temperature.
- e) None of the above statements are true.

5. A solution is formed at room temperature by vigorously dissolving enough of the solid solute so that some solid remains at the bottom of the solution. Which statement below is TRUE?

- a) The solution is considered unsaturated.
- b) The solution is considered supersaturated.
- c) The solution is considered saturated.
- d) The solution would be considered unsaturated if it were cooled a bit to increase the solubility of the solid.
- e) None of the above are true.

6. Which of the following compounds will be most soluble in pentane (C_5H_{12})?

- a) pentanol ($CH_3CH_2CH_2CH_2CH_2OH$)
- b) benzene (C_6H_6)
- c) acetic acid (CH_3CO_2H)
- e) ethyl methyl ketone ($CH_3CH_2COCH_3$)
- e) None of these compounds should be soluble in pentane.

7. Which of the following would be the best solvent to use to dissolve SO_3 ?

- a) H_2O
- b) C_6H_{14}
- c) $NaCl$
- d) any of these will work
- e) none of these will work

8. What volume of a 0.716 M KBr solution is needed to provide 30.5 g of KBr?

9. Show the three types of interactions that can occur in an aqueous solution of CH_3NH_2 .

10. What is the molarity of $C_2H_3O_2^-$ ions in a 1.0 M solution of aluminum acetate?

Review 6

1. Classify the following compounds as electrolytes (strong or weak) or non-electrolytes. If the compound is an electrolyte, give the number of moles of particles formed in solution per moles of formula units dissolved. For example, for one mole of NaCl formula units, 2 moles of particles are formed due to the dissociation of ions.

a. C_2H_6 _____

b. K_2SO_3 _____

c. HF _____

d. H_2SO_4 _____

e. $Ca(OH)_2$ _____

f. PCl_3 _____

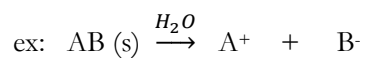
g. H_2CO_3 _____

h. C_6H_5OH _____

i. $RbNO_3$ _____

j. $HOCH_2CH_2OH$ _____

2. Write an equation to illustrate the behavior of the following substances in water. Make sure your equations are balanced!



a. $AgI (s)$

b. $K_3PO_4 (s)$

c. $Cu_3(PO_4)_2 (s)$

d. $Zn(OH)_2 (s)$

e. Barium acetate

f. Ammonium hydroxide

3. Calculate the mass of nitrogen dissolved in a 5.0 qt fish tank (at 25 °C) if the mole percentage of nitrogen in the air is 78 % and the total pressure is 1.0 atm (k_H for N_2 in water at 25 °C is 6.1×10^{-4} M/atm).

4. An aqueous solution of copper (II) chloride is made using 72.5 g of copper (II) chloride diluted to a total solution volume of 1.50 L. Calculate the molarity, molality and mass percent of the solution (assume a density of 1.05 g/mL for the solution).

5. What is the molarity of a 10.5 % by mass glucose ($C_6H_{12}O_6$) solution? (density of solution is 1.03 g/mL)

6. For the solutions below, calculate the **molarity of sodium ions** for each. Show all work.

A 0.513 g of Na_3PO_4 dissolved in 100.0 mL of solution

B 25.00 mL of 2.74 M NaCl diluted to make 250.0 mL of solution

C 0.0984 moles of Na_2SO_4 dissolved in 200.0 mL of solution

7. Rank the compounds below from the solution that has the highest vapor pressure to the solution with the lowest vapor pressure. All solutions are at 25°C. The vapor pressure of water at 25°C is 28.3 torr. The vapor pressure of CH₃OH at 25°C is 85.2 torr. The vapor pressure of CH₃CH₂OH at 25°C is 70.3 torr. The other substances are nonvolatile. Glucose has a molecular weight of 180.2 g/mol.

- A 0.0138 mol Na₂SO₄ in 25.00 g water
- B 0.0138 mol CaCl₂ in 25.00 g water
- C 0.0138 mol glucose in 25.00 g water
- D 0.0138 mol glucose in 25.00 g CH₃CH₂OH

Highest 1_____ 2_____ 3_____ 4_____ Lowest

Explain your reasoning below. First, show the calculations. Second, describe how you could make predictions about the order without doing a calculation.

8. Determine the van't Hoff factor for each compound when dissolved in water.

- a. Na₂SO₄
- b. K₂S
- c. Cr(NO₃)₃
- d. CH₃OCH₃
- e. CH₃COOH
- f. HClO₄

Review 7

1. Which of the following solutions will raise the BP of water the most? Explain your choice.

- a) 0.010 *m* FeCl₂
- b) 0.015 *m* LiNO₃
- c) 0.020 *m* CH₃CH₂OH
- d) 0.025 *m* CH₂Cl₂

2. A solution is made by placing 222.2 g of sodium phosphate into a volumetric flask, where 6.50 x 10² grams of water are added. At what temperature will this solution boil? At what temperature will this solution freeze?

3. Which of these aqueous solutions would be expected to have the highest boiling point?

- a) 0.100 *m* CaCl₂
- b) 0.200 *m* NaOH
- c) 0.050 *m* K₂SO₄
- d) 0.050 *m* Al₂(SO₄)₃
- e) 0.200 *m* CH₃OH

4. Which aqueous solution would have the lowest vapor pressure at 25 °C? (no calculation is required)

- a) 1 *m* NaCl
- b) 1 *m* Na₃PO₄
- c) 1 *m* sucrose, C₁₂H₁₀O₁₁
- d) 1 *m* MgCl₂
- e) 1 *m* glucose, C₆H₁₂O₆

Use the table below to answer the questions that follow:

solvent	Formula	K _b (°C/m)	K _f (°C/m)	Boiling point (°C)	Freezing point (°C)
water	H ₂ O	0.512	1.86	You know	You know
benzene	C ₆ H ₆	2.53	5.12	80.1	5.5

5. Which of the following would you expect to have the highest boiling point?

- a) pure water
- b) pure benzene
- c) acetone dissolved in benzene
- d) acetone dissolved in water

6. Which of the following would have the lowest freezing point?

- a) pure water
- b) 0.1 *m* CH₂F₂
- c) 0.1 *m* NiF₂
- d) both b & c have the same freezing point

7. Rank the substances/mixtures below from the highest boiling point to the lowest boiling point.

- A** pure water
- B** 0.10 *m* NaCl(*aq*)
- C** 0.10 *m* CCl₄ in benzene (C₆H₆)
- D** 0.20 *m* K₂S(*aq*)
- E** 0.02 *m* K₂CO₃(*aq*)
- F** 0.10 *m* CH₃OH(*aq*)

Highest 1 _____ 2 _____ 3 _____ 4 _____ 5 _____ 6 _____ Lowest

Explain your reasoning below without using or referring to the results of any calculations.

8. Calculate the molecular weights from the following scenarios:

a. A nonionizing solid that freezes at -1.20°C when 12.0 g are dissolved in 50.00 g water

b. 6.78 g of this substance dissolved in 35.00 g of benzene freezes at 3.09°C

9. Which of the following would you expect to have the highest value?

- a) heat of vaporization for a polar molecule
- b) heat of vaporization for a nonpolar molecule
- c) heat of fusion for a polar molecule
- d) heat of fusion for a nonpolar molecule

Review 8

1. Don't do a calculation, but estimate the final temperature when 10.0 g of aluminum at 75°C is allowed to come to thermal equilibrium with 10.0 g of gold at 25°C.

- a) 20°C b) 30°C c) 50°C d) 70°C e) 75°C f) 80°C

2. If 8.8×10^2 J of energy are added to 50.0 g of aluminum at 20.0°C, calculate (and show all work) the final temperature of the aluminum. The specific heat of aluminum is 0.21 cal/g °C.

3. If a 50.0 g piece of an unknown metal at 100.0°C is added to 50.0 g of water at 25.0 °C and the final temperature is measured to be 28.5 °C, what is the specific heat of the unknown metal?

4. Use tabulated data from your text book or the internet for this problem. If the same quantity of energy was added to each system below, rank from the highest final temperature of the system to the lowest.

- A. 20.0 grams of lead
- B. 10.0 grams of lead
- C. 20.0 grams of water
- D. 100.0 grams of water
- E. 20.0 grams of aluminum
- F. 40.0 grams of aluminum

Highest 1_____ 2_____ 3_____ 4_____ 5_____ 6_____ Lowest

Explain your reasoning below. (No true calculations--use approximations and/or proportional reasoning and explain with words.)

5. Which of the following substances (with specific heat capacity provided) would show the greatest temperature change upon absorbing 100.0 J of heat?

a) 10.0 g Ag, $C_{\text{Ag}} = 0.235 \text{ J/g}^\circ\text{C}$

b) 10.0 g H₂O, $C_{\text{H}_2\text{O}} = 4.18 \text{ J/g}^\circ\text{C}$

c) 10.0 g ethanol, $C_{\text{ethanol}} = 2.42 \text{ J/g}^\circ\text{C}$

d) 10.0 g Fe, $C_{\text{Fe}} = 0.449 \text{ J/g}^\circ\text{C}$

e) 10.0 g Au, $C_{\text{Au}} = 0.128 \text{ J/g}^\circ\text{C}$

6. How much energy is lost when cooling 36.0 g gaseous H₂O from 115°C to a liquid at 65°C?

The following physical data may be useful.

$\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$

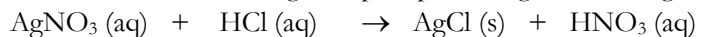
$C_{\text{liq}} = 4.18 \text{ J/g}^\circ\text{C}$

$C_{\text{gas}} = 2.01 \text{ J/g}^\circ\text{C}$

$C_{\text{sol}} = 2.09 \text{ J/g}^\circ\text{C}$

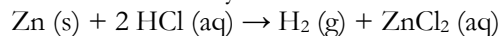
Review 9

1. The addition of HCl to a solution of AgNO₃ precipitates AgCl according to the reaction below:



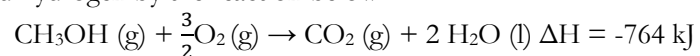
When 50.0 mL of 0.100 M AgNO₃ is combined with 50.0 mL of 0.100 M HCl in a coffee-cup calorimeter, the temperature changes from 23.40 °C to 24.21 °C. Calculate ΔH_{rxn} for the reaction as written. (Assume the density and specific heat capacity of the solution is that of water)

2. Zinc metal reacts with hydrochloric acid according to this balanced equation:



When 0.103 g of Zn (s) is combined with enough HCl to make 50.0 mL of solution in a coffee-cup calorimeter, all of the zinc reacts, raising the temperature of the solution from 22.5 °C to 23.7 °C. Find ΔH_{rxn} for this reaction as written (density and specific heat of solution are that of water).

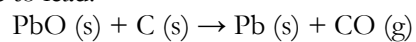
3. Methanol, CH₃OH is an efficient fuel with a high octane rating that can be produced from coal and hydrogen by the reaction below.



a. How much energy is evolved when 115.0 g of CH₃OH burns in excess oxygen?

b. What mass of O₂ is consumed when 925 kJ of energy is given off?

4. What is ΔH (in kJ) for the reaction as written below if 5.95 kJ must be supplied to convert 13.43 g of lead (II) oxide to lead?



Review 10

1. What is the enthalpy for the reaction ($Mn(s) + O_2(g) \rightarrow MnO_2(s)$) given the following information:



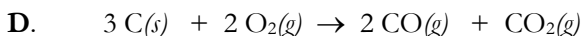
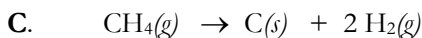
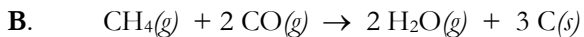
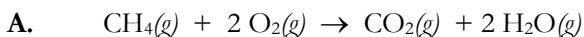
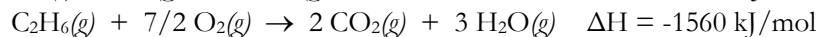
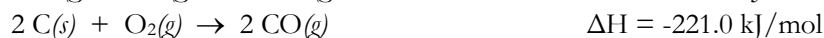
For questions 2 and 3, consider the reaction: $2 SO_2 + O_2 \rightarrow 2 SO_3$; $\Delta H = -196 \text{ kJ/mol}$

2. When this reaction occurs, you should observe

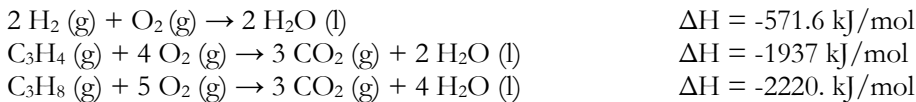
- a) temperature increasing b) temperature decreasing c) no temp change

3. When sulfur dioxide reacts with oxygen and 1.094 g of sulfur trioxide is made, how much energy is released or absorbed?

4. Use Hess's law and the equations given below to determine the enthalpy of reaction for the reactions in A-D. Not every equation will be used for each problem.



5. Given that,



Determine the enthalpy of the reaction: $\text{C}_3\text{H}_4 (\text{g}) + 2 \text{H}_2 (\text{g}) \rightarrow \text{C}_3\text{H}_8 (\text{g})$

6. An exothermic reaction has

- a) a negative ΔH and absorbs heat from the surroundings. An exothermic reaction feels warm to the touch.
- b) a positive ΔH and absorbs heat from the surroundings. An exothermic reaction feels warm to the touch.
- c) a positive ΔH and gives off heat to the surroundings. An exothermic reaction feels warm to the touch.
- d) a negative ΔH and gives off heat to the surroundings. An exothermic reaction feels warm to the touch.

7. Which of the following statements are FALSE?

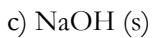
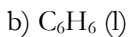
- a) If a chemical equation is multiplied by some factor, $\Delta H^\circ_{\text{rxn}}$ is also multiplied by the same factor.
- b) If a chemical reaction changes direction, then $\Delta H^\circ_{\text{rxn}}$ changes sign.
- c) If a chemical equation can be expressed as the sum of a series of steps, $\Delta H^\circ_{\text{rxn}}$ for the overall equation is the sum of the heats of reactions for each step.
- d) Standard states are all in the gas phase.

8. Which of the following processes is exothermic?

- a) The formation of dew in the morning.
- b) The melting of ice.
- c) The chemical reaction in a "cold pack" often used to treat injuries.
- d) The vaporization of water
- e) None of the above are exothermic.

Review 11

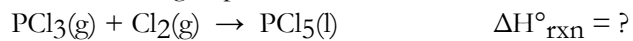
1. Write the formation reactions for the following substances, being sure to include all phases:



2. Use bond energies to calculate the ΔH for the combustion of ethanol (shown below). (Hint: first draw all Lewis structures and see notes for bond energy values)

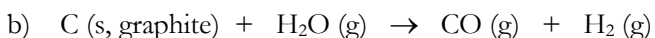
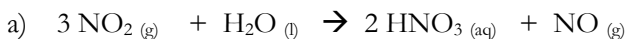


3. Use the bond energies provided to estimate $\Delta H^\circ_{\text{rxn}}$ for the reaction below.



<u>Bond</u>	<u>Bond Energy (kJ/mol)</u>
Cl-Cl	243
P-Cl	331

4. Using tabulated standard enthalpies of formation (see the appendix in your textbook), calculate ΔH° for the following reactions:

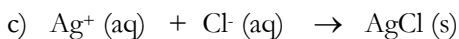
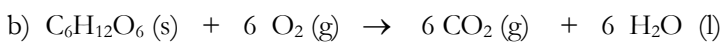
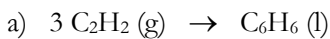


Classify the reactions above as endothermic or exothermic.

Review 12

1. The normal boiling point of the element rubidium is 688°C. The following equilibrium is important in rubidium vapor formation at that temperature: $2 \text{Rb}(l) \rightleftharpoons \text{Rb}_2(g)$. Predict the signs of ΔH° and ΔS° for this reaction.

2. WITHOUT doing a calculation, predict the sign of $\Delta S^\circ_{\text{rxn}}$ for the following reactions:



3. Rank the systems below from the one with the highest entropy to the one with the lowest.

A $\text{C}_6\text{H}_{12}(g)$ at 400 K in a 2.0 L flask

B $\text{C}_6\text{H}_{12}(g)$ at 700 K in a 2.0 L flask

C $\text{C}_6\text{H}_{12}(l)$ at 200 K in a 2.0 L flask

D $\text{CH}_4(g)$ at 200 K in a 2.0 L flask

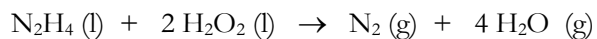
E $\text{CH}_4(g)$ at 700 K in a 2.0 L flask

F $\text{CH}_4(g)$ at 700 K in a 1.0 L flask

Highest 1_____ 2_____ 3_____ 4_____ 5_____ 6_____ Lowest

Explain your reasoning below.

4. a) Using the values of standard molar **entropies**, calculate the entropy change at 25 °C for the following reaction:



b) Does the sign of $\Delta S^\circ_{\text{rxn}}$ make sense? Explain.

c) Calculate the **enthalpy** change for the reaction at 25 °C using the values of standard enthalpies of formation.

d) Is this reaction endothermic or exothermic?

Review 13

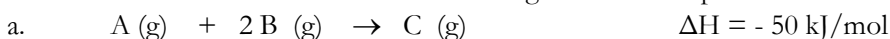
1. Calculate ΔG° at 45 °C for reactions for which

a. $\Delta H^\circ = 293 \text{ kJ/mol}$; $\Delta S^\circ = -695 \text{ J/mol}\cdot\text{K}$

b. $\Delta H^\circ = -86.6 \text{ kJ/mol}$; $\Delta S^\circ = -382 \text{ J/mol}\cdot\text{K}$

c. Tell whether reactions a and b are spontaneous.

2. Under what conditions would the following reactions be spontaneous?



Questions 3-5 refer to the following exothermic reaction: $2 \text{NO}_2(\text{g}) \rightarrow \text{N}_2\text{O}_4(\text{g})$

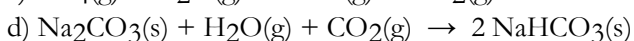
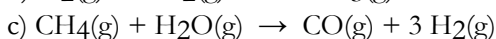
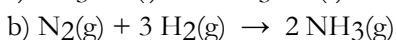
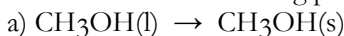
3. What is the sign on ΔH ?

4. What is the sign on ΔS ?

5. Based on your answers to 3 & 4, under what conditions will the reaction be spontaneous?

a) all conditions b) high temperatures c) low temperatures d) it will never be spontaneous

6. Which of the following processes have a $\Delta S > 0$?



e) All of the above processes have a $\Delta S > 0$.

7. Which of the following statements is TRUE?

a) Entropy is not a state function.

b) Endothermic processes decrease the entropy of the surroundings, at constant T and P.

c) Endothermic processes are never spontaneous.

d) Exothermic processes are always spontaneous.

e) None of the above are true.

8. Consider a reaction that has a positive ΔH and a positive ΔS . Which of the following statements is TRUE?

a) This reaction will be spontaneous only at high temperatures.

b) This reaction will be spontaneous at all temperatures.

c) This reaction will be nonspontaneous at all temperatures.

d) This reaction will be nonspontaneous only at high temperatures.

e) It is not possible to determine without more information.

9. Place the following in order of decreasing molar entropy at 298 K.



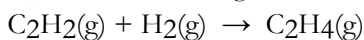
- a) $\text{Ar} > \text{N}_2\text{H}_4 > \text{HCl}$ b) $\text{Ar} > \text{HCl} > \text{N}_2\text{H}_4$ c) $\text{N}_2\text{H}_4 > \text{Ar} > \text{HCl}$
d) $\text{N}_2\text{H}_4 > \text{HCl} > \text{Ar}$ e) $\text{HCl} > \text{N}_2\text{H}_4 > \text{Ar}$

10. Place the following in order of decreasing standard molar entropy.



- a) $\text{NaCl(s)} > \text{NaCl(aq)} > \text{Na}_3\text{PO}_4(\text{aq})$
b) $\text{NaCl(aq)} > \text{NaCl(s)} > \text{Na}_3\text{PO}_4(\text{aq})$
c) $\text{Na}_3\text{PO}_4(\text{aq}) > \text{NaCl(aq)} > \text{NaCl(s)}$
d) $\text{NaCl(s)} > \text{Na}_3\text{PO}_4(\text{aq}) > \text{NaCl(aq)}$
e) $\text{NaCl(aq)} > \text{Na}_3\text{PO}_4(\text{aq}) > \text{NaCl(s)}$

11. Calculate $\Delta S^\circ_{\text{rxn}}$ for the following reaction. The S° for each species is shown below the reaction.



$$S^\circ(\text{J/mol}\cdot\text{K}) \quad 200.9 \quad 130.7 \quad 219.3$$

- A) +112.3 J/K B) +550.9 J/K C) -112.3 J/K D) +337.1 J/K E) -550.9 J/K

12. Estimate $\Delta G^\circ_{\text{rxn}}$ for the following reaction at 387 K.



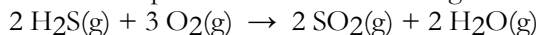
- A) +243 kJ B) -72.9 kJ C) +84.9 kJ D) -92.5 kJ E) -188 kJ

13. Use Hess's law to calculate $\Delta G^\circ_{\text{rxn}}$ using the following information.



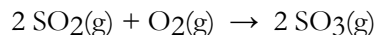
- A) -60.0 kJ B) +651.6 kJ C) -265.8 kJ D) +137.2 kJ E) +523.0 kJ

14. Above what temperature does the following reaction become nonspontaneous?



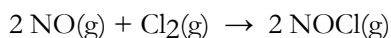
Review 14

1. Given the following balanced equation, determine the rate of reaction with respect to $[\text{SO}_2]$.



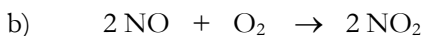
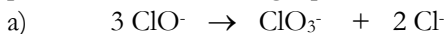
- a) Rate = $-\frac{1}{2} \frac{\Delta[\text{SO}_2]}{\Delta t}$ b) Rate = $+\frac{1}{2} \frac{\Delta[\text{SO}_2]}{\Delta t}$ c) Rate = $-\frac{2 \Delta[\text{SO}_2]}{\Delta t}$
d) Rate = $+\frac{2 \Delta[\text{SO}_2]}{\Delta t}$ e) It is not possible to determine without more information.

2. Given the following balanced equation, determine the rate of reaction with respect to $[\text{NOCl}]$. If the rate of Cl_2 loss is $4.84 \times 10^{-2} \text{ M/s}$, what is the rate of formation of NOCl ?



- a) $4.84 \times 10^{-2} \text{ M/s}$ b) $2.42 \times 10^{-2} \text{ M/s}$ c) $1.45 \times 10^{-1} \text{ M/s}$ d) $9.68 \times 10^{-2} \text{ M/s}$
e) $1.61 \times 10^{-2} \text{ M/s}$ f) $-9.68 \times 10^{-2} \text{ M/s}$

3. Express the average rate of reaction in terms of the rate of change of each reactant and each product in the following equations:



4. At some time, we observe that the reaction below is forming NO_2 at the rate of 0.048 M/s .



- a. What is the rate of change of O_2 per second ($\Delta[\text{O}_2]/\Delta t$) at this time?
b. What is the rate of change of N_2O_5 per second ($\Delta[\text{N}_2\text{O}_5]/\Delta t$) at this time?
c. What is the average rate of the reaction at this time?

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

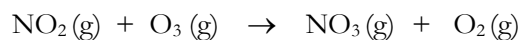
$$\text{Rate} = \frac{1}{2} \frac{\Delta[\text{N}_2]}{\Delta t} = \frac{\Delta[\text{O}_2]}{\Delta t} = -\frac{1}{2} \frac{\Delta[\text{N}_2\text{O}]}{\Delta t}$$

Review 15

1. Consider the data showing the initial rate of a reaction ($A \rightarrow \text{products}$) at different concentrations of A. What is the order of the reaction? Write a rate law for the reaction, including the value of the rate constant, k.

[A], M	Rate (M/s)
0.12	0.0078
0.16	0.0104
0.20	0.0130

2. The following kinetic data were obtained for the reaction:



Experiment	[NO ₂], M	[O ₃], M	initial rate (M/s)
1	2.1×10^{-3}	7.0×10^{-4}	6.3×10^{-3}
2	2.1×10^{-3}	1.39×10^{-3}	1.25×10^{-2}
3	3.8×10^{-3}	7.0×10^{-4}	1.14×10^{-2}
4	6.6×10^{-3}	1.8×10^{-4}	?

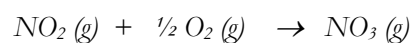
a. Write the rate law for this reaction (show your work).

b. What is the order of the reaction (overall)?

c. Determine the value of the rate constant (with the correct units).

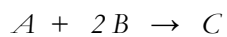
d. Predict the rate of the reaction in experiment 4 based on the determined rate law.

3. Use the data below to determine the rate law and rate constant for the reaction:



	[NO ₂], M	[O ₂], M	Rate (M/s)
Trial 1	1.2	0.6	3.37
Trial 2	1.8	1.8	7.58
Trial 3	1.2	1.2	3.37
Trial 4	3.4	1.8	27.1

4. Given the following reaction data, determine the experimental rate law for the reaction below.

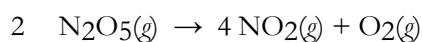


Trial	[A] (M)	[B] (M)	Rate (M/s)
1	0.200	0.200	0.0465
2	0.250	0.200	0.0581
3	0.200	0.450	0.0698
4	0.125	0.125	0.0230

Review 16

1. Using $t_{1/2} = t$ and $[A] = \frac{1}{2} [A]_0$, solve the zero, first and second order integrated rate equations for half-life ($t_{1/2}$).

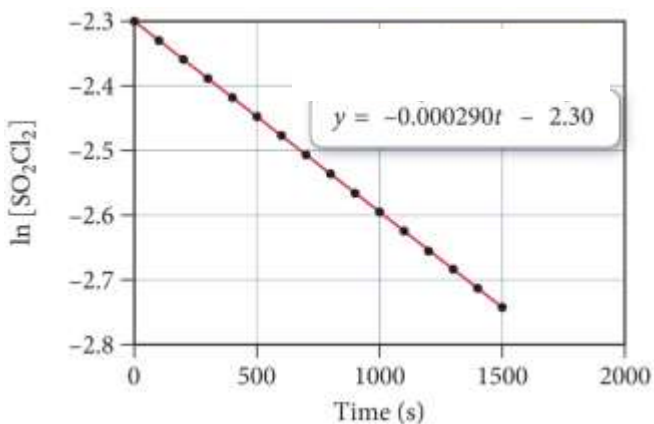
2. The decomposition of dinitrogen pentoxide is described by the chemical equation:



a. If the rate of disappearance of N_2O_5 is equal to 1.60 mM/min at a particular moment, what is the rate of appearance of NO_2 at that moment?

b. What is the average rate of the reaction?

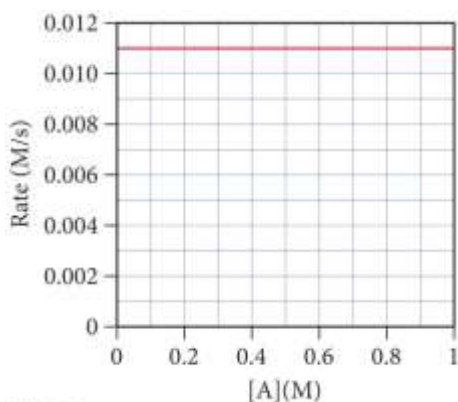
3. Answer the questions below based on the following equation and graph below:



a. Write the rate law expression for the reaction given. Be sure to answer as completely as possible.

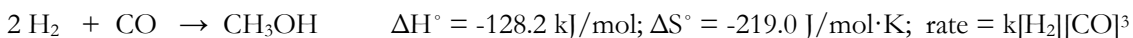
b. What is the half-life for the reaction?

4. A graph of rate vs. concentration for a given reaction generates the data shown below.



- What is the order of the reaction with respect to [A]? (1 pt)
- Based on your answer to part a, what are the units of k for this reaction? (1 pt)

The questions below refer to the following reaction and information given about it:



- What is the overall order of the reaction?
- If the rate with respect to hydrogen gas is 0.20 M/s, what is the rate with respect to carbon monoxide?
- If the concentration of hydrogen gas is doubled (CO remains constant), how will the rate of reaction change? (please answer with: up or down and by what factor)
- What is the ΔG° value for this reaction at 25°C?
- At what temperature values is this reaction spontaneous?

10. Rank these reactions from greatest to least overall order of reaction.

A. $k = 15 \text{ s}^{-1}$

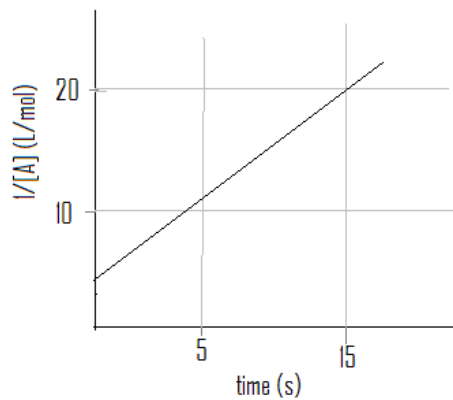
B. $\text{rate} = 0.15 [\text{A}][\text{B}]$

C. constant half-life of 20 seconds

D. $\text{rate} = 1.5 \times 10^3 \text{ M/s}$
(at all concentrations)

E. $k = 0.12 \text{ M}^{-3}\text{min}^{-1}$

F.



Greatest 1 _____ 2 _____ 3 _____ 4 _____ 5 _____ 6 _____ **Least**

Explain your reasoning below.

Review 17

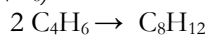
1. a. For a given 1st order reaction, the concentration of reactant is determined to decrease from 2.00 M to 1.50 M in 64.0 minutes. What is the rate constant for the reaction?

b. Calculate the time required for one half of the sample to decompose.

2. The first order rate constant for the dissociation of A is 0.173 s^{-1} . Calculate the time needed for the concentration of A to decrease to $\frac{1}{25}$ of its initial concentration.

3. The first-order decomposition of N_2O at 1000 K has a rate constant of 0.76 s^{-1} . If the initial concentration of N_2O is 10.9 M, what is the concentration of N_2O after 9.6 min?

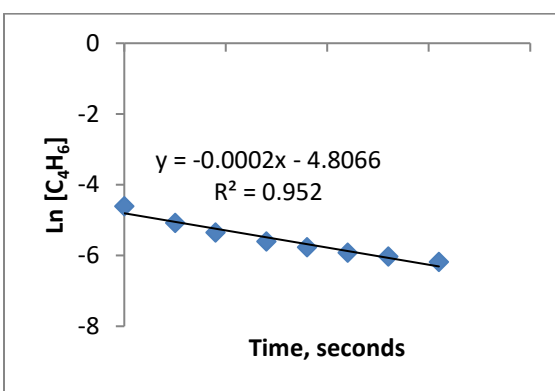
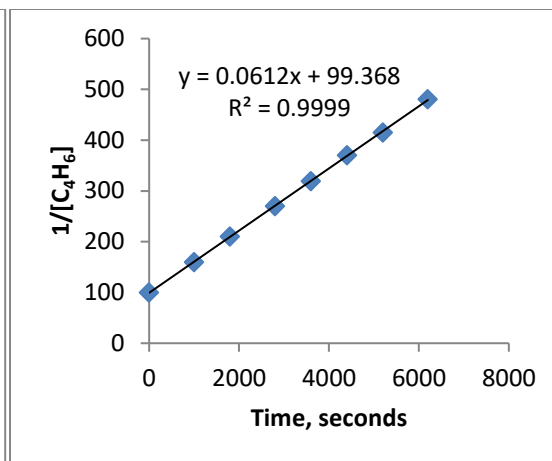
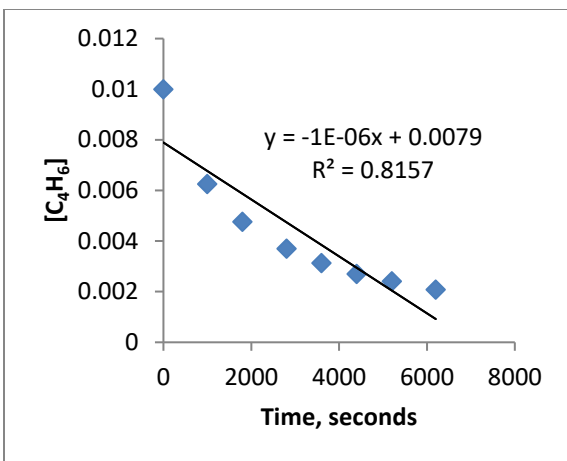
4. Butadiene (C_4H_6) reacts to form its dimer according to the equation below:



The following data were collected for this reaction at a given temperature:

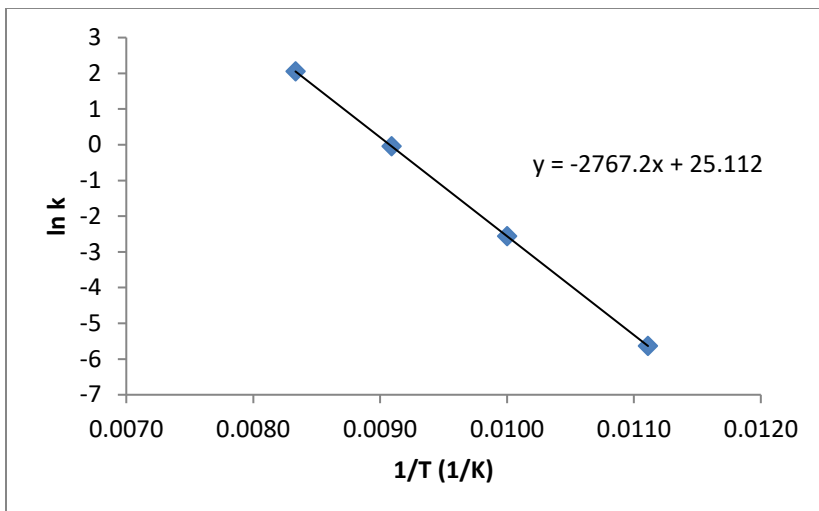
time, seconds	$[\text{C}_4\text{H}_6]$, M
0	0.01
1000	0.00625
1800	0.00476
2800	0.0037
3600	0.00313
4400	0.0027
5200	0.00241
6200	0.00208

From this data, the following graphs were generated:



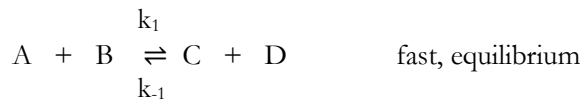
- Is this reaction zero, first or second order? How can you tell?
- What is the value of the rate constant for the reaction (watch your units!!!)?
- What is the rate law for the reaction (include your calculated value of k)?
- What is the half-life for the reaction under the initial conditions of this experiment?
- Using the appropriate integrated rate law, calculate the amount of time required for 0.01 M butadiene to be reduced down to 0.0015 M.

5. From the information provided in the graph, determine the activation energy (including appropriate units) for the reaction below: $\text{N}_2\text{O}_5(\text{g}) \rightarrow \text{NO}_3(\text{g}) + \text{NO}_2(\text{g})$



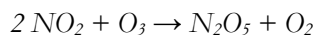
Review 18

1. Using the following elementary steps, answer the questions below.



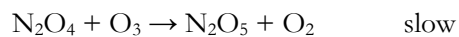
- What is the overall or bulk reaction?
- List any intermediates in the reaction mechanism (if any).
- List any catalysts in the reaction mechanism (if any).
- Write the rate law for the reaction (remember the rate law cannot contain reaction intermediates).

2. The reaction of nitrogen dioxide and ozone, has been studied at 231 K. Given the following data, determine the experimental rate law.

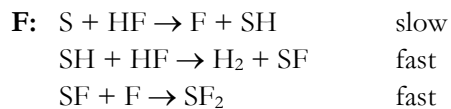
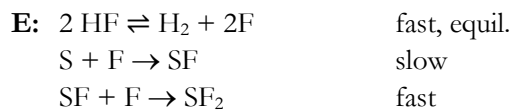
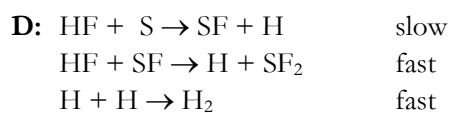
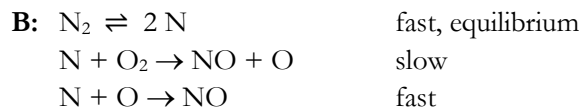
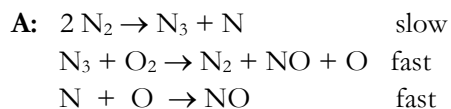


Experiment	Initial [NO ₂], M	Initial [O ₃], M	Initial rate (M/s)
1	1.33 x 10 ⁻²	2.35 x 10 ⁻²	3.50 x 10 ⁻⁶
2	5.32 x 10 ⁻²	2.35 x 10 ⁻²	1.40 x 10 ⁻⁵
3	1.33 x 10 ⁻²	7.83 x 10 ⁻³	1.17 x 10 ⁻⁶

Is the following proposed mechanism (shown below) consistent with the given kinetic data? You must SHOW your work to get credit on this problem.

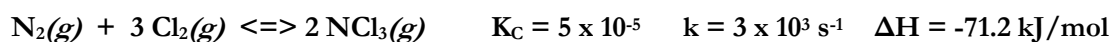


3. Possible mechanisms for two reactions are given below. Determine the rate law predicted by each of the mechanisms.



Review 19

The following questions refer to the reaction below. Answer in the blank. Write clearly!



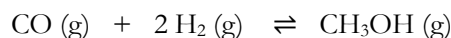
- _____ 1. The overall reaction order is
a. zero b. first c. second d. third e. more than 3
- _____ 2. If the reaction were run in a smaller container, the equilibrium concentration of N_2 will be
a. higher b. lower c. the same d. insufficient info to predict
- _____ 3. If more Cl_2 is added to the reaction at equilibrium, the concentration of N_2 will
a. increase b. decrease c. stay the same d. there is insufficient info to predict
- _____ 4. If the reaction were run at a higher temperature, the equilibrium concentration of N_2 will be
a. higher b. lower c. the same d. insufficient info to predict
- _____ 5. If the reaction were run at a higher temperature, the rate of reaction would
a. increase b. decrease c. stay the same d. insufficient info to predict
- _____ 6. To calculate K_P for this reaction, you would use the formula $K_P =$
a. K_C b. $K_C(RT)$ c. $K_C(RT)^{-2}$ d. $K_C(RT)^2$ e. $K_C(RT)^3$
- _____ 7. In a potential energy diagram of this reaction,
a. the products would be at a higher level than the reactants
b. the products would be at a lower level than the reactants
c. the products and reactants would be nearly equal
d. you can't draw a potential energy diagram for this reaction
- _____ 8. The reaction above will make
a. large amounts of products quickly b. small amounts of products quickly
c. large amounts of products slowly d. small amounts of products slowly
- _____ 9. The rate-determining step in a mechanism is the
a. first step b. the first step after an equilibrium (if there is one) c. the last one
d. the one with the highest rate constant e. the one with the lowest rate constant
- _____ 10. A catalyst increases the rate of reaction by
a. changing the mechanism b. lowering the activation energy
c. orienting the molecule properly d. all of these e. none of these

Review 20

1. Which of the following statements is FALSE?

- a) When $K \gg 1$, the forward reaction is favored and essentially goes to completion.
- b) When $K \ll 1$, the reverse reaction is favored and the forward reaction does not proceed to a great extent.
- c) When $K \approx 1$, neither the forward or reverse reaction is strongly favored, and about the same amount of reactants and products exist at equilibrium.
- d) $K \gg 1$ implies that the reaction is very fast at producing products.
- e) None of the above.

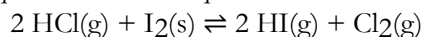
2. Consider the reaction below that is performed in a 2.00 L container.



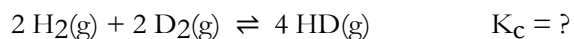
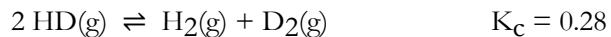
An equilibrium mixture of this reaction at a certain temperature was found to have 0.150 moles of CO, 0.50 moles of H₂ and 5.50 moles of CH₃OH. What is the value of the equilibrium constant, K, at this temperature?

3. Determine the value of K for the following reaction if the equilibrium concentrations are as follows:

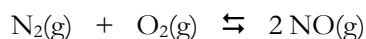
$[\text{HCl}]_{\text{eq}} = 0.13 \text{ M}$, $[\text{HI}]_{\text{eq}} = 5.6 \times 10^{-16} \text{ M}$, $[\text{Cl}_2]_{\text{eq}} = 0.0019 \text{ M}$.



4. The equilibrium constant is given for one of the reactions below. Determine the value of the missing equilibrium constant.



5. For the reaction below, the value of the equilibrium constant at 25°C is 1×10^{-30} . What is the feasibility of this reaction (in terms of NO production) at the specified temperature?



6. For the reaction $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$, $K = 1.8 \times 10^4$ at a particular temperature. If the equilibrium concentrations of N_2 and NH_3 are 0.015 M and 2.00 M, what is the equilibrium concentration of H_2 ?

7. Consider the reaction $2\text{CH}_4(\text{g}) \rightarrow \text{C}_2\text{H}_2(\text{g}) + 3\text{H}_2(\text{g})$, where $K_P = 22.4$.

a) Is the system at equilibrium if $P_{\text{CH}_4} = 4.00 \text{ atm}$, $P_{\text{C}_2\text{H}_2} = 4.00 \text{ atm}$, and $P_{\text{H}_2} = 4.00 \text{ atm}$? If not, which direction will the reaction proceed to reach equilibrium? Justify your response.

b) Is the system at equilibrium if $P_{\text{CH}_4} = 10.0 \text{ atm}$, $P_{\text{C}_2\text{H}_2} = 10.0 \text{ atm}$, and $P_{\text{H}_2} = 10.0 \text{ atm}$? If not, which direction will the reaction proceed to reach equilibrium? Justify your response.

8. Calculate K_p for the reaction of hydrogen gas and chlorine gas to form hydrogen chloride gas at 55 °C if the concentration of the gases at equilibrium are $[H_2]=4.0 \times 10^{-2} \text{ M}$, $[Cl_2] = 4.2 \times 10^{-4} \text{ M}$ and $[HCl] = 3.2 \times 10^{-1} \text{ M}$.

9. Give the direction of the reaction under standard conditions, if $K \ll 1$.

- a) The forward reaction is favored.
- b) The reverse reaction is favored.
- c) Neither direction is favored.
- d) If the temperature is raised, then the forward reaction is favored.
- e) If the temperature is raised, then the reverse reaction is favored.

10. Which of the following statements are TRUE?

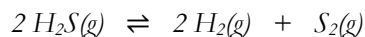
- a) Dynamic equilibrium occurs when the rate of the forward reaction equals the rate of the reverse reaction.
- b) The equilibrium constant for the forward reaction is equal to the equilibrium constant for the reverse reaction.
- c) A reaction quotient (Q) larger than the equilibrium constant (K) means that the reaction will favor the production of more products.
- d) Dynamic equilibrium indicates that the amount of reactants and products are equal.
- e) All of the above are true.

Review 21

1. When 0.200 moles of N_2 reacts with 0.200 moles of H_2 in a 1.00 L container (to produce NH_3), the $[N_2]$ at equilibrium is 0.190 M. What are the equilibrium concentrations of H_2 and NH_3 and what is the value of K_c ?

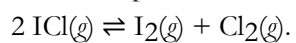
2. The reaction $NiO(s) + CO(g) \rightleftharpoons Ni(s) + CO_2(g)$ has $K_c = 4.00 \times 10^1$. If a mixture of solid nickel (II) oxide and 0.350 M CO come to equilibrium, what will be the equilibrium concentration of carbon dioxide?

3. The system below reaches equilibrium in a 2.6 liter container.



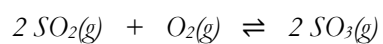
At equilibrium the container holds 1.82 mole H_2S , 0.62 mole H_2 and 0.36 mole S_2 . Calculate the equilibrium constant, K .

4. At a certain temperature the equilibrium constant, K_c , equals 0.11 for the reaction:



What is the equilibrium concentration of ICl if 0.45 mol of I_2 and 0.45 mol of Cl_2 are initially mixed in a 2.0-L flask?

5. Three moles of pure SO_3 are placed in an 8.00 liter flask at 1150K. At equilibrium, 0.580 mol O_2 has formed. Calculate K_c for this reaction at 1150K.

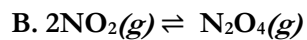


Review 22

1. Using the conditions given, calculate equilibrium concentrations for all species in reactions A-C. The temperature is 500°C.

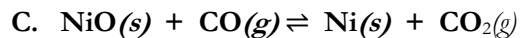


$K_C = 2.4 \times 10^{-4}$ initial: 0.01 mol Ag_2SO_4 in 1.0 L solution



$K_C = 164$

initial: 0.055 mol NO_2 in 1.0L



$K_C = 4.0 \times 10^{-3}$

initial conditions: 1.0 mol NiO and 1.0 mol CO in 1.0 L flask.

2: For the following reaction at 25 °C, $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2 \text{NO}_2(\text{g})$

Substance	ΔH_f° (kJ/mol)	S° (J/mol•K)
$\text{N}_2\text{O}_4(\text{g})$	9.16	304.4
$\text{NO}_2(\text{g})$	33.2	240.1

- Calculate the standard Gibbs free energy of the reaction.
- Using your answer to part a, calculate the equilibrium constant for the reaction.
- If 2.0 moles of NO_2 and 0.5 moles of N_2O_4 are placed in a 1.0 L container, in which direction does the reaction spontaneously proceed?
- Prove your answer to part c by calculating the Gibbs free energy of the reaction under the conditions in part c.

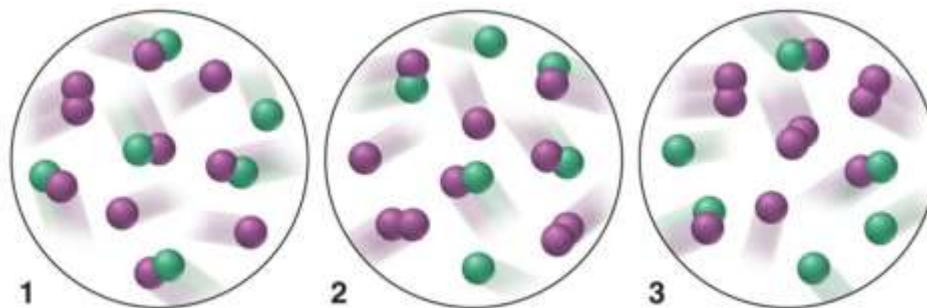
Example 3: For $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2 \text{NO}_2(\text{g})$, $K_c = 4.61 \times 10^{-3}$ at 25 °C. If at equilibrium, $[\text{NO}_2] = 0.0236 \text{ M}$ in a 2.26 L flask, how many grams of N_2O_4 are present?

Example 4: Consider the following reaction: $\text{CO}_2(\text{g}) + \text{C}(\text{s, graphite}) \rightleftharpoons 2\text{CO}(\text{g})$

A reaction mixture initially contains 0.56 atm CO_2 and 0.32 atm CO . Determine the equilibrium pressure of CO if K_P for the reaction at this temperature is 2.25.

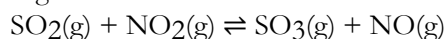
Review 23

1. Consider the reaction and scenario below and answer the questions that follow (note: you MUST be able to see the colors to solve this problem):



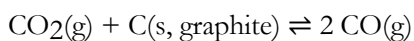
- If $K = 2$ at the temperature of the reaction, which scene represents the mixture at equilibrium?
- Will the reaction mixtures in the other two scenes proceed toward reactant or toward products to reach equilibrium?
- For the mixture at equilibrium, how will a rise in temperature affect $[Y_2]$ and K ?
- How will a decrease in pressure influence the mixture at equilibrium?

2. Consider the following reaction at equilibrium and answer the questions that follow as shift right, shift left or no change.



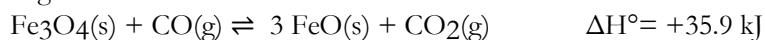
- Adding more SO_2
- Removing NO
- Increasing the pressure by decreasing the volume
- Addition of a catalyst

3. Consider the following reaction at equilibrium and answer the questions that follow as shift right, shift left or no change.



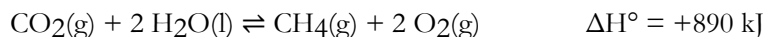
- a) Adding graphite
- b) Decreasing the pressure by increasing the volume
- c) Adding CO
- d) Removing CO₂

4. Consider the following reaction at equilibrium and answer the questions that follow as shift right, shift left or no change.



- a) Adding FeO
- b) Adding CO
- c) Increasing the temperature
- d) Increasing the pressure by reducing the volume
- e) Removing CO₂

5. Consider the following reaction at equilibrium. Answer the following questions based on what happens to the concentration of oxygen. (ie, if the following stresses are applied to the reaction, does the concentration of O₂ increase, decrease or stay the same?)



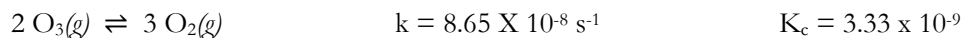
- a) Increasing temperature
- b) Increasing the amount of carbon dioxide
- c) adding a catalyst
- d) Increasing the pressure by decreasing the volume
- e) Addition of Argon, an inert gas

6. For the reactions below, calculate the equilibrium constant using the information given.



D. Without doing a calculation, what conclusions could you make (if any) about the equilibrium constants for each reaction?

7. Consider the following reaction to answer questions A-E below:



A. If a container with equal amounts of both ozone (O_3) and oxygen gas were allowed to come to equilibrium...

- a. you quickly end up with a lot of oxygen
- b. you would quickly end up with a lot of ozone
- c. you would slowly end up with a lot of oxygen
- d. you would slowly end up with a lot of ozone

B. Write the K_c expression for the reaction.

C. Write the K_p expression for the reaction

D. Yes or No—Are the values of K_c and K_p the same?

E. Write the rate law for this reaction.

8. Which of the following is not true for ΔG_{rxn} ?

- a) If $\Delta G^\circ_{\text{rxn}} > 0$, the reaction is spontaneous in the forward direction.
- b) If $Q = 1$, then $\Delta G_{\text{rxn}} = \Delta G^\circ_{\text{rxn}}$.
- c) If $\Delta G^\circ_{\text{rxn}} = 0$, the reaction is spontaneous in the reverse direction.
- d) If $\Delta G^\circ_{\text{rxn}} > 0$, the reaction is spontaneous in the reverse direction.
- e) Under equilibrium conditions, $\Delta G_{\text{rxn}} = 0$.

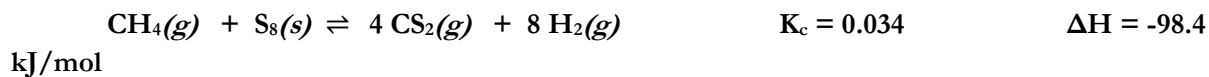
9. Choose the statement below that is TRUE.

- a) $K > 1$, $\Delta G^\circ_{\text{rxn}}$ is positive b) $K < 1$, $\Delta G^\circ_{\text{rxn}}$ is negative c) $\Delta G^\circ_{\text{rxn}} = 0$ at equilibrium
d) $\Delta G_{\text{rxn}} = 0$ at equilibrium e) None of the above statements are true.

10. If $\Delta G < 0$, which of the following statements are true?

- a) The reaction is spontaneous.
b) The reaction is not spontaneous.
c) The reaction is at equilibrium.
d) Not enough information is given.

Consider the following reaction to answer questions 11-15:



11. Write the K_c expression for the reaction.

12. Write the K_p expression for the reaction

13. If carbon disulfide is added to the system at equilibrium, how will the concentration of hydrogen gas change?

14. How will the value of the equilibrium constant change if the temperature is increased?

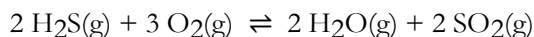
15. If 2.0 moles of CH_4 , 5.0 mol of S_8 , 0.30 mol CS_2 and 0.020 mol of H_2 were added to a 2.0 L container, in what direction does the reaction spontaneously move?

16. Consider the endothermic reaction $2A(s) + B_2(g) \rightleftharpoons 2AB(g)$, where $B_2(g)$ is in its standard state. The standard Gibbs Free Energy of formation for $A(s)$ and $AB(g)$ are $-197.6 \frac{\text{kJ}}{\text{mol}}$ and $-200.7 \frac{\text{kJ}}{\text{mol}}$, respectively, at 298.15 K.

- a. Determine the Standard Gibbs Free energy change for the reaction.
- b. Write the equilibrium expression, K_C , for the reaction.
- c. Determine the equilibrium constant, K_C , for the reaction.
- d. When a mixture of 179.5 g of A is placed in a 1.50 L container with 3.00 moles of B_2 and 4.50 moles of AB , determine the equilibrium concentration of AB .

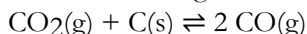
Review 24

1. Consider the following reaction at equilibrium. What effect will adding more H₂S have on the system?



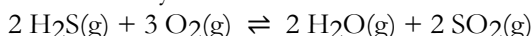
- a) reaction shifts left b) No change c) The equilibrium constant will decrease.
d) The equilibrium constant will increase. e) reaction shifts right

2. Consider the following reaction at equilibrium. What effect will adding some C have on the system?



- a) No effect b) The equilibrium constant will decrease. c) reaction shifts left
d) The equilibrium constant will increase. e) reaction shifts right

3. Consider the following reaction at equilibrium. What effect will increasing the volume of the reaction mixture have on the system?



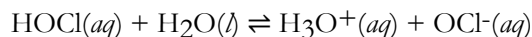
- a) increases [SO₂] b) No effect c) decreases [SO₂]
d) The equilibrium constant will decrease. e) The equilibrium constant will increase.

4. Consider the following reaction: $\text{CuS}(\text{s}) + \text{O}_2(\text{g}) \rightleftharpoons \text{Cu}(\text{s}) + \text{SO}_2(\text{g})$

An equilibrium mixture contains 1.2 M SO₂. Determine the equilibrium concentration of O₂ if K_c for the reaction at this temperature is 1.5.

- a) 1.8 M b) 0.96 M c) 1.3 M d) 0.80 M e) not enough information is provided

5. The reaction below contains 4 species: a Bronsted-Lowry acid, Bronsted-Lowry base, a conjugate acid and a conjugate base. Label each.



6. Write the formula of the conjugate base of HPO₄²⁻ _____

7. In the following reaction, HClO₃ is acting as a Bronsted-Lowry acid and CO₃²⁻ is acting as a Bronsted-Lowry base. Predict the products of this reaction (pay careful attention to the charges).



8. Label the following as acid, base or salt.

- a) HCNO _____
b) NaBrO _____
c) NH₃ _____
d) KC₂H₃O₂ _____

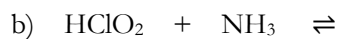
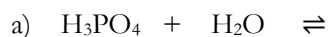
Review 25

1. Fill in the table to the right to complete the conjugate acid/base pairs.

Acid	Base
HClO ₄	
	CO ₃ ²⁻
	H ₂ PO ₄ ⁻
C ₆ H ₅ NH ₃ ⁺	
	H ₂ O
HC ₄ H ₇ O ₂	
H ₃ PO ₄	
	SO ₃ ²⁻

2. The species H₂PO₄⁻ is amphoteric. Write the 2 reactions showing the behavior of H₂PO₄⁻ in aqueous solution.

3. Complete the following reactions. The first species listed is the acid in these reactions.



4. For the reaction in #3c, show the movement of electrons by drawing the Lewis structures of the acid and base and using arrows to describe how the reaction occurs.

5. Predict whether the following ions or molecules can act as either a Lewis acid or base and explain why.



Review 26

1. Fill in the table below.

Solution	$[\text{H}_3\text{O}^+]$	$[\text{OH}^-]$	pH	pOH
A			3.2	10.8
B	1.70×10^{-11}			
C		3.33×10^{-8}		
D			9.35	
E		4.31×10^{-5}		
F				6.78

2. Circle the stronger acid in each pair and provide an explanation for your answer.

- a) HClO_2 or HClO_3
- b) HClO_2 or HBrO_2
- c) CH_3COOH or CBr_3COOH
- d) CBr_3COOH or CH_2BrCOOH
- e) LiH or HF
- f) H_2S or H_2Se

3. Using tabulated data as necessary, answer the questions below.

a) In each of the following pairs of acids, circle the stronger acid.

HI vs. HF

HNO_2 vs. $\text{HC}_7\text{H}_5\text{O}_2$

H_3BO_3 ($K_a = 5.4 \times 10^{-10}$) vs. HBrO ($\text{p}K_a = 8.55$)

b) In the above pairs of acids, circle the acid that would produce a solution with the lowest pH (assume the concentrations are the same).

c) Which solution will have the lower pH, 0.1 M HF or 1.0 M HF? Why?

Knowing that weaker acids produce stronger conjugate bases, pick the stronger bases below.

d) In each pair, pick the stronger base:

I⁻ and F⁻

C₂H₃O₂⁻ and CHO₂⁻

C₆H₅O⁻ and C₇H₅O₂⁻

4. **True or False.** The factor that relates the concentration of H⁺ in a pH 5 solution to a pH 8 solution is 1000 (or 0.001, depending on how you set up the fraction).

5. Rank 0.10 M aqueous solutions of the acids below from strongest to weakest. You may find the appendix of K_a values in your text helpful. (You should not need a calculator for this question!)

- A** CH₃COOH
- B** HClO₄
- C** HNO₃
- D** HNO₂
- E** HCNO

Strongest 1_____ 2_____ 3_____ 4_____ 5_____ Weakest

6. For each of the substances below, 1) write the reaction that describes their behavior in water and 2) calculate the pH of a 0.10 M aqueous solution of the substances.

a. CH₃COOH

b. HClO₄

c. HNO_2

d. HOCN

7. Calculate the percent ionization of 0.15 M HIO (hypoiodous acid, can be written as HOI).

8. Calculate the K_a of 0.045 M solution of a weak acid if the pH of the solution is 4.02.

9. Arrange the following 0.10 M solutions in order of increasing pH. The lowest pH should be given #1.

NaOH

NH₃

HC₂H₃O₂

Ba(OH)₂

10. A 0.10 M solution of a weak acid (HA) ionizes 0.85%.

a) Calculate the pH of the acid.

b) Calculate the K_a of the acid.

Review 27

1. Classify the following as acid, base or salt. **For the salts, determine the parent acid and base and classify the salt as acidic, basic or neutral.**

- a) HBrO
- b) NaClO_2
- c) $\text{C}_3\text{H}_7\text{NH}_2$
- d) $\text{C}_5\text{H}_5\text{NHBr}$
- e) NH_4NO_3

2. Put a check mark in front of each statement that is true about an aqueous solution of these substances.

- _____ a. A stronger acid will have a higher pH
- _____ b. A stronger acid will have a higher K_a .
- _____ c. A stronger base will have a higher pH.
- _____ d. HNO_3 is the strongest of all the strong oxyacids.

3. Determine whether the following salts are acidic, basic or neutral. Write the appropriate equilibrium reactions for salts that react with water.

- a) K_2CO_3
- b) $\text{C}_6\text{H}_5\text{NH}_3\text{Cl}$
- c) LiF
- d) BaSO_4
- e) Na_3PO_4
- f) NH_4NO_3
- g) NH_4F

4. Calculate the pH of each solution below. You may find Tables of K_a and K_b values (in your text) helpful.

a. 0.020 M CH_3NH_2

b. 0.010 M CsOH

c. 0.50 M $\text{C}_5\text{H}_5\text{N}$

d. 0.050 M HCl

e. 0.200 M LiF.

f. 0.350 M NH_4Br .

Review 28

1. Rank the following compounds from lowest to highest pH assuming the concentration of each compound is 0.10 M.

- a) HCl
- b) $\text{Sr}(\text{OH})_2$
- c) BaF_2
- d) $\text{HC}_2\text{H}_3\text{O}_2$
- e) NH_4Cl

2. Consider a solution that contains 0.13 M benzoic acid and 0.12 M sodium benzoate.

- a. Calculate the pH using an ICE table.

- b. Calculate the pH using the Henderson-Hasselbalch equation.

- c. Is it valid to use the Henderson-Hasselbalch equation here? Why or why not?

3. Which of the following solutions is a good buffer system?

- A) A solution that is 0.10 M NaCl and 0.10 M HCl
- B) A solution that is 0.10 M HCN and 0.10 M LiCN
- C) A solution that is 0.10 M NaOH and 0.10 M HNO_3
- D) A solution that is 0.10 M HNO_3 and 0.10 M KNO_3
- E) A solution that is 0.10 M HCN and 0.10 M NaCl

Explain why you chose your answer.

4. If the pK_a of $HCHO_2$ is 3.74 and the pH of an $HCHO_2/NaCHO_2$ solution is 3.74, which of the following is TRUE?

- A) $[HCHO_2] > [NaCHO_2]$
- B) $[HCHO_2] = [NaCHO_2]$
- C) $[HCHO_2] < [NaCHO_2]$
- D) $[HCHO_2] < <[NaCHO_2]$
- E) It is not possible to make a buffer of this pH from $HCHO_2$ and $NaCHO_2$.

Explain why you chose your answer.

5. A 1.0 L buffer solution is 0.10 M in HF and 0.050 M in NaF. Which of the following actions would destroy the buffer?

- a) Adding 0.050 mol of HCl
- b) Adding 0.050 mol of NaF
- c) Adding 0.05 mol of NaOH
- d) None of the above
- e) More than one of the above

Explain your answer.

6. Which of the following acids would be the best choice to combine with its sodium salt to make a solution buffered at pH 2?

Chlorous acid $pK_a = 1.95$

Nitrous acid $pK_a = 3.34$

Formic acid $pK_a = 3.74$

Hypochlorous acid $pK_a = 7.54$

Review 29

Consider the titration of 25.00 mL of 0.100 M NaOH with 0.152 M HNO₃.

1. What is the titration reaction?
2. What is the volume of titrant required to reach the equivalence point?
3. What are the primary species present at 0 mL (ie what's in the beaker when 0 mL of HNO₃ has been added)? What is the pH at 0 mL of titrant added?
4. What are the primary species present at the equivalence point?
5. **Approximately**, what is the pH at the equivalence point? (no calculation required)
6. Sketch the titration curve for this experiment. Label each axes, show where pH = 7 and the equivalence point is.
7. What would be a good indicator to use for this experiment, and what would be the color at the equivalence point?

Consider the titration of 20.0 mL of 0.18 M CH₃NH₂ with 0.100 M HCl.

1. What is the titration reaction?
2. What is the volume of titrant required to reach the equivalence point?

3. What is the volume at half-way to the equivalence point ($V_{e/2}$)?
4. What is/are the primary species present at $V_{e/2}$?
5. What is the pH at $V_{e/2}$? (this should be an exact number)
6. Sketch the titration curve for this experiment. Label each axes, show where $\text{pH} = 7$ and the equivalence point is.
7. **Approximately**, what is the pH at the equivalence point?
8. What would be a good indicator to use for this experiment, and what would be the color at the equivalence point?
9. What is the primary species present at 50.00 mL of HCl added and what is the approximate pH?

For each of the titrations described below, rank the pH at the volume of titrant described from highest pH to lowest pH. You may find a table of equilibrium constants helpful.

- A** The titration of 25 mL of 0.10 M HNO_3 with 0.10 M NaOH
volume titrant = equivalence point
- B** The titration of 25 mL of 0.10 M HNO_3 with 0.10 M NaOH
volume titrant = 0 mL
- C** The titration of 25 mL of 0.10 M HF with 0.10 M NaOH
volume titrant = equivalence point
- D** The titration of 25 mL of 0.10 M HF with 0.10 M NaOH
volume titrant = half of the equivalent point volume
- E** The titration of 30 mL of 0.10 M $\text{C}_5\text{H}_5\text{N}$ with 0.10 M HCl
volume titrant = half of the equivalent point volume
- F** The titration of 30 mL of 0.10 M $\text{C}_5\text{H}_5\text{N}$ with 0.10 M HCl
volume titrant = 35 mL

Highest pH 1 _____ 2 _____ 3 _____ 4 _____ 5 _____ 6 _____ Lowest pH

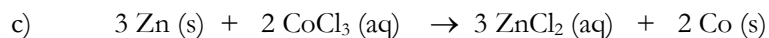
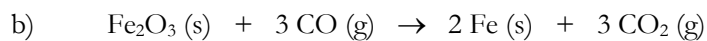
Explain your reasoning below (Show the titration reaction! Use back if needed. What substances are present in significant concentration?). You should not have to perform any calculations.

Assign oxidation numbers to the underlined element in the following compounds.

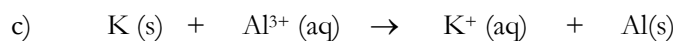
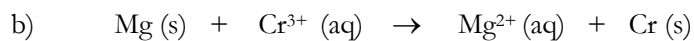
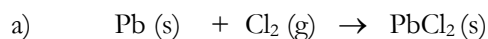
- a. S₈ _____ h. Zn _____
- b. Cr₂O₇²⁻ _____ i. HNO₂ _____
- c. BaSO₄ _____ j. Li⁺ _____
- d. S₂O₃²⁻ _____
- e. SO₃²⁻ _____
- f. KMnO₄ _____
- g. H₂CO _____

Review 30

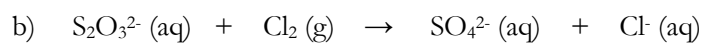
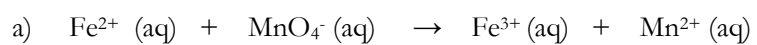
1. In the following reactions, identify the species that is oxidized and the species that is reduced. Also identify the oxidizing agent and the reducing agent.



3. Balance the following redox reactions using the half-reaction method and determine the number of electrons transferred in each reaction. Also identify the oxidation and reduction half-reactions.



4. Balance the following redox reactions in **acidic solution**.

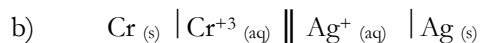
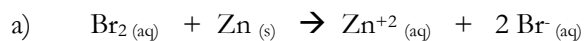


Review 31

1. a) Calculate the standard potential (E°_{cell}) of a voltaic cell made of Mg/Mg²⁺ and Cu/Cu²⁺ half-cell reactions at 25°C.

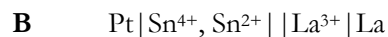
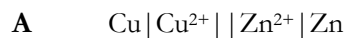
b) Write the balanced net equation.

2. Calculate the cell potential for the following reactions under standard conditions; indicate if each is spontaneous or not.



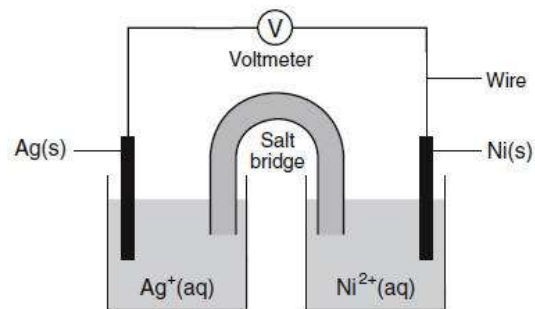
3. For the reactions in #2, calculate the standard free energy change for each. Does your answer in #3 agree with your answer in #2 (ie in regards to spontaneity)?

4. Determine the balanced overall reaction for the following electrochemical cells. (Write reaction as described by the cell.). For each reaction, 1) determine the number of electrons transferred 2) calculate E°_{cell} and 3) calculate the standard gibbs free energy change



5. Consider the voltaic cell to the right and answer the questions that follow:

- What is the overall reaction?
- What is the anode? The cathode?
- What species is oxidized? reduced?
- What should the voltmeter read?
- What is the ΔG° for the reaction if all species in aqueous solution are at a concentration of 1.0 M?



f) Write the shorthand notation for the cell.

Review 32

1. Using a table of standard reduction potentials, which is the strongest oxidizing agent? Ag^+ , Cl_2 or Zn^{2+}
2. Using a table of standard reduction potentials, which is the strongest reducing agent? Au , Br^- or Pb
3. Will electrons flow spontaneously from Ca to Br_2 ? Why or why not?
4. Will electrons flow spontaneously from Cl^- to Ni^{2+} ? Why or why not?
5. A voltaic cell consists of a Pb/Pb^{2+} and a Cu/Cu^{2+} half-cell at 25°C . The initial concentrations of Pb^{2+} and Cu^{2+} are 0.0500 M and 1.50 M respectively. What is the cell potential?
6. Calculate the molar solubility of Ag_2SO_4 in pure water if the K_{sp} is 6.9×10^{-15} .