Things to Know – Chem 1A

Chapter 1

1. What is matter? What is chemistry?
2. Macroscopic vs. submicroscopic ("molecular view")
3. Atom vs. molecule (what's the difference?)
4. States of matter (how do the observable properties relate to the arrangement and motion of the molecules in each?)
5. Classification of matter: pure substances (elements or compounds), mixtures (homogeneous or heterogeneous). How do the observable properties differ? How do the particles differ? Understand and be able to explain the differences. Be able to draw a molecular picture of each. Give an example of each.
6. Mixtures can be purified – how?
7. Compounds can be separated into elements by a chemical reaction (not physical methods.)
8. Physical vs. chemical properties, examples
9. Extensive vs. intensive properties, examples
10. Physical change vs. chemical change
11. Metric system, meaning of prefixes: kilo-, centi-, milli-, micro-, nano-
12. Temperature conversions (°C, K, °F)
13. Accuracy vs. precision. Can you have one without the other?
14. Interpolation of measurements
15. Significant figures: counting, rounding off, determining the uncertainty in a measurement.
16. Dimensional Analysis (conversion factors) – practice lots of these!
17. Conversions between squared or cubed units (square or cube the entire conversion factor)
18. Density: how to calculate it ($d = m/v$), how to use it as a conversion factor (2.70 g/cm^3 means $2.70 \text{ g}/ 1 \text{ cm}^3$)
19. Percent: how to calculate it, how to use it as a conversion factor (12.3% X by mass means 12.3 g X/ 100 g total)

Chapter 2

1. Postulates of Dalton's atomic theory
2. How atomic theory explained the law of conservation of mass and the law of constant composition
3. Law of multiple proportions
4. Know, in general terms, how early experiments led to conclusions about the structure of the atom. Know who discovered what, and the order of discoveries.
5. What did Rutherford's experiment show? How? Be able to explain this one.

6. Structure of atom: protons, electrons, neutrons. (Mass, charge, and location)
7. Atomic number, mass number, isotope symbols (know how to interpret and how to write isotope symbols)
8. Isotopes – definition and examples
9. Calculating weighted-average atomic masses

$$\textbf{Atomic mass} = (\textbf{abundance A})(\textbf{mass A}) + (\textbf{abundance B})(\textbf{mass B}) + ...$$

In the above equation, use abundances as decimals. The decimal abundances add up to 1 (exactly).
10. General stuff about the periodic table: period, group, metals, nonmetals, semimetals, main-group, transition elements, inner-transition elements, alkali, alkaline earth, halogens, noble gases.
11. Why is the periodic table arranged the way it is?
12. Diatomic elements
13. Compounds – molecular vs. ionic – what do they consist of?
14. What is an ion? Cation? Anion?
15. Know the normal charges of main-group ions.
16. Transition metals – charges can vary (except Ag^+ and Zn^{2+})
17. Know the formulas, charges, and names of polyatomic ions.
18. There are no molecules in ionic compounds!
19. What holds ionic compounds together?
20. Naming and writing formulas for ionic compounds (metal + nonmetal): name cation, then anion. <u>No prefixes!</u> $Fe_2(SO_3)_3$ is iron (III) sulfite, K_2O is potassium oxide. Compounds must be uncharged overall.
21. Naming and writing formulas for acids
    Binary acids: HCl is hydrochloric acid. Oxoacids: $HClO_3$ is chloric acid.
22. Naming and writing formulas for binary molecular compounds (2 nonmetals): Use prefixes. N_2O is dinitrogen monoxide, NO_2 is nitrogen dioxide. Remember, there are no ions in binary molecular compounds.
23. Organic compounds are carbon-based, and they have a totally different naming system. They can be very long and complicated molecules.

Chapter 3

1. Writing and balancing chemical equations for reactions
2. Be able to identify combination reactions, decomposition reactions, and combustion reactions.
3. For a combination reaction involving a metal and a nonmetal, be able to predict the formula of the product. (The product will be ionic. Think of the charges on the ions.)
4. Be able to write the equation for any combustion reaction.
5. Be able to calculate the molecular weight, formula weight, or molar mass of any substance. (These terms all mean basically the same thing.)

6.   Be able to find the mass percent of each element in a compound.

$$\text{mass \% A} = \frac{\text{mass of A}}{\text{total mass}} \times 100$$

7.   What is a mole? What is Avogadro's number?
8.   Be able to convert grams to moles to molecules and vice versa.
9.   Also be able to calculate how many atoms of a particular element are in a particular compound. (Look at the formula of the compound to get a conversion factor. For example, in 1 mole of $C_4H_{10}O$, there are 4 moles C, 10 moles H, and 1 mole of O.)
10.   Find the empirical formula of a compound from masses or mass percents. (Convert all masses to moles of each element, then divide all of them by the smallest number of moles to get the mole ratio. Make sure you have the lowest whole number ratio.)
11.   What is the relationship between the empirical formula and the molecular formula?
12.   Be able to determine molecular formula from the empirical formula and the approximate molar mass.
13.   Determining empirical formula from combustion data – understand the concepts – why does this work?
      Compound containing C, H, and O:
      a. From mass of CO_2 produced, calculate the mass of C.
      b. From mass of H_2O produced, calculate the mass of H.
      c. Total mass compound – mass C – mass H = mass O in compound.
      d. Convert the masses of all elements to moles.
      e. Take mole ratio to get the formula.
14.   Hydrocarbons (same type of problem as above) shortcut:
      a. From mass of CO_2 produced, calculate the moles of C.
      b. From mass of H_2O produced, calculate the moles of H.
      c. Take the mole ratio to get the formula.
15.   Information you can get from a balanced equation: mole ratio of molecules that react.
16.   Conversions: g of X ⟷ mol of X ⟷ mol of Y ⟷ g of Y
      Be able to do any variation. You need a balanced equation. If an equation is given, balance it first (don't assume that it's already balanced).
17.   Limiting reactant problems: if you are given specific amounts of two or more different reactants, you must first determine which is limiting.
      a. Find moles of each.
      b. Compare the stoichiometric ratio from the balanced equation to the mole ratio that you have.
      c. Which is limiting? Which is in excess?
      d. Finish the problem using the number of moles of the LR that you have.
      e. Be able to determine how much of the excess reactant is left over (first, calculate how much was used up in reacting with the LR, then subtract from the initial amount).

18. Percent yield, actual yield, theoretical yield – what are they? What does it mean if the percent yield is less than 100%? More than 100%?
19. Be able to calculate the percent yield or use it as a conversion factor.

Chapter 4

1. Solution, solute, solvent definitions
2. Strong vs. weak vs. nonelectrolytes – behavior in solution
3. Why is water a good solvent?
4. Solubility rules (will be provided) – how to use them
5. Writing equations: predicting products for precipitation, acid-base, and gas forming reactions
6. Be able to write total and net ionic equations. Also be able to write the net ionic equation without first writing the regular and the total ionic equations.
7. Remember: weak acids are not written as separate ions. Only soluble strong electrolytes exist as ions.
8. Memorize the 7 strong acids: HCl, HBr, HI, HNO_3, $HClO_3$, $HClO_4$, and H_2SO_4 for its first ionization.
9. Properties of acids and bases, meaning of monoprotic, diprotic, triprotic.
10. Acid-base reactions: what are the products? Be able to write equations.
11. For a polyprotic acid in excess base: remove all H^+ ions.
    Ex: $H_2SO_4 + 2NaOH \rightarrow 2 H_2O + Na_2SO_4$
12. Reaction between ammonia and an acid – predict products
13. Gas forming reactions to know:
    a carbonate + acid $\rightarrow$ $CO_{2\,(g)}$ + $H_2O_{(l)}$ + ionic compound
    a bicarbonate + acid $\rightarrow$ $CO_{2\,(g)}$ + $H_2O_{(l)}$ + ionic compound
    a sulfide + acid $\rightarrow$ $H_2S_{(g)}$ + ionic compound
    ammonium salt + strong base $\rightarrow$ $NH_{3\,(g)}$ + $H_2O_{(l)}$ + ionic compound
    Be able to write the equations for any of the above reactions.
14. Definitions of oxidation, reduction.
15. Be able to find the oxidation number of any atom (in an element or compound)
16. Given a reaction, determine what is oxidized and what is reduced (first find oxidation numbers).
17. If an element appears in a reaction, the reaction must be a redox reaction.
18. Understanding the activity series and how to determine relative activities:
    $A + BC \rightarrow B + AC$ If this happens, A is more active than B.
19. Molarity = # moles solute/ L solution. Be able to calculate molarity or use it as a conversion factor.
20. Concentration of ions: 0.1 M $Na_2SO_{4\,(aq)}$ is 0.2 M Na^+ and 0.1 M SO_4^{2-}.
21. Describe how to make a solution of a given molarity. (Use a volumetric flask. Describe the process in words.)

22.   Dilutions – you can use $M_cV_c = M_dV_d$, where V_d is the total final volume of solution ($V_d = V_c$ + water added). Why? Because # moles of solute doesn't change.
23.   Titrations – no limiting reactant. Need a balanced equation. Start with the substance you can find moles of. Convert to moles of the other substance. Finish the problem: could find molarity, volume, mass %, molar mass, etc. **DO NOT** use $M_1V_1 = M_2V_2$. Why not?
24.   Be able to do limiting reactant problems involving solutions. Write the net ionic equation, find concentrations of all ions present after the reaction.

Chapter 10

1.   Properties of gases
2.   Pressure – what is it? How is it measured? How does atmospheric pressure vary? What are gas molecules doing to exert pressure?
3.   How does a barometer work? What does it measure? Why use Hg as the liquid?
4.   How does an open-end manometer work? What about a closed-end manometer? What does each one measure?
5.   Boyle's Law: $P_1V_1 = P_2V_2$ As pressure increases, volume decreases. (Used if n and T are constant.) **Why**? Explain at the molecular level.
6.   Charles' Law: as T increases, volume also increases. V is proportional to temp in Kelvin. (Used if n and P are constant.) **Why**? Explain at the molecular level.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

7.   Gay-Lussac's Law: as T increases, P increases (assuming n and V are constant). **Why**? Explain at the molecular level.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

8.   Combined gas law (used when n is constant but P, V and T are changing)

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

9.   Avogadro's law: volume is proportional to moles of gas (assume P and T are constant). **Why**? Explain at the molecular level.

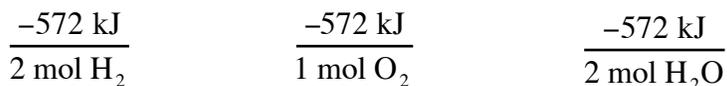$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

10.   What is STP? What is the molar volume of an ideal gas at STP? (22.4 L/mole : can use it as a conversion factor.)

11. Reacting gases: if measured at the same T and P, volume is proportional to moles. (Can use volume ratios instead of mole ratios – this is an important shortcut!)

12. Ideal gas law $PV = nRT$, $R = 0.08206$ L•atm/K•mol Units used must match the units of R.

13. Be able to derive any of the other gas laws from $PV = nRT$ (put constants on one side, variables on the other)

14. Given the density of a gas at a specific temperature and pressure, calculate MM of a gas (# g/# mol) Approach: assume 1 L. Find the number of moles of gas in 1 L, then divide # g/L by #mol/L.

15. Given the molar mass of a gas, P, and T, calculate density. Approach: find the number of moles in 1 L, then convert to g.

16. Stoichiometry problems involving gases, including limiting reactant problems

17. Gas mixtures - partial pressures $P_{total} = P_A + P_B + P_C + ...$,

18. $P_{total} = n_{total}RT/V$ $P_A = n_ART/V$ $P_A = X_AP_{total}$

19. Collect a gas over water – subtract vapor pressure of water from total pressure – why?

20. In Lab 13 calculation, why did you have to subtract an extra correction factor from P_{total} to get P_{H2}?

21. Kinetic-Molecular theory – **how it explains behavior of gases**

22. Velocity distribution of gas molecules (Boltzmann distribution) – low vs. high temp. (What does the graph look like? Are all the molecules in a sample traveling the same speed?)

23. Using the equation below: what units are used? What is u_{rms}?

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$

24. Diffusion or effusion problems – comparing rates, times, # moles, etc. – reason it out! Make sure your answer makes sense. Molar mass ratio goes under the square root symbol. Practice these types of problems.

25. Nonideal gases – under what conditions do gases exhibit nonideal behavior? Why?

26. Van der Waals equation – what does it take into account?

Chapter 5

1. Energy, work, conservation of energy. Kinetic vs. potential energy. Types of potential energy

2. Units of energy

3. system, surroundings, closed system

4. heat, thermal energy (molecular level explanation for things next to each other ending up at the same temp.)

5. First law of Thermo: $\Delta E = q + w$. What is included in E?

6. Sign conventions for q and w

7. Endothermic vs. exothermic

8.   State function vs. path function. What is the significance of state functions?

9.   Know how to calculate w for a given reaction. $w = -P\Delta V = -\Delta n_{gas}RT$

10.   Enthalpy (ΔH) – heat transferred at constant pressure. Heat transferred at constant volume = ΔE. $\Delta E = \Delta H - P\Delta V$ or $\Delta E = \Delta H + w$

11.   Thermochemical equations – implied amounts (kJ per x moles)

12.   Stoichiometry using ΔH – how much energy is given off per x g of reactant? Example: $2\ H_2 + O_2 \rightarrow 2\ H_2O\ \ \Delta H = -572\ kJ$ means:

$$\frac{-572\ kJ}{2\ mol\ H_2} \qquad \frac{-572\ kJ}{1\ mol\ O_2} \qquad \frac{-572\ kJ}{2\ mol\ H_2O}$$

13.   Specific heat (C) vs. heat capacity (also C) (units?) They are different!

14.   What does it mean if a substance has a high specific heat?

15.   Use $q = Cm\Delta T$ or $q = C\Delta T$ for temperature changes.

16.   Coffee-cup calorimetry – constant P (open container) $q_P = \Delta H$
$q_{calorimeter} = q_{solution}$ (Usually the heat capacity of the syrofoam cup itself is ignored.) We often have to assume that the solution in the calorimeter has a density and heat capacity close to those of water, since the solution consists of mostly water.

17.   Bomb calorimetry – rigid container, so $\Delta V = 0$, so $q_V = \Delta E$.
$q_{calorimeter} = (q_{water} + q_{bomb})$ Sometimes, the heat capacity of the entire calorimeter is given with the water already included. In that case, $q_{calorimeter} = C_{calorimeter}\Delta T$.

18.   For both of the above types of calorimetry, $q_{calorimeter} = -q_{reaction}$! In order to get ΔH or ΔE, divide by the number of moles of the limiting reactant. Then check the stoichiometry of the reaction – adjust for coefficients.

19.   Manipulating thermochemical equations (reverse a rxn, multiply a rxn by N, add reactions) – what happens to ΔH in each case?

20.   Hess's Law: given several reactions and their ΔH values, determine the ΔH value for a new reaction by rearranging and adding up the given reactions.

21.   Heat of formation – know how it's defined, be able to write the corresponding reaction.

22.   Given $\Delta H°_f$ values, calculate ΔH_{rxn} (products – reactants)

23.   OR Given $\Delta H°_f$ values and ΔH_{rxn}, calculate a missing $\Delta H°_f$ value. (Don't forget: for elements, the $\Delta H°_f = 0$. Why?)

24.   Phase changes - constant temp (Why? Be able to explain at the molecular level). Be able to draw a heating curve.

25.   Calculating total energy involved in a process that has phase changes and temperature changes (calculate them separately, then add)

26. ΔH_{vap}, ΔH_{fus} – What are they? Which one is higher? Why?

27.   Which phase changes have a positive q? Which have a negative q?

28.   Lab 15 calculation!

29.   Heat involved in phase changes practice worksheet!

30.   Heat is also involved in making a solution and diluting. (given calorimetry data, calculate ΔH of solution.) ex: $NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$

Chapter 6

1. Wave properties: wavelength, amplitude, velocity, frequency
2. Types of electromagnetic radiation – which ones are high energy? Which are low energy? Which color of visible light has the highest frequency and which color has the lowest frequency?
3. $c = \upsilon\lambda$
4. What is a node?
5. What is a standing wave? What is the significance of a standing wave?
6. What is quantization?
7. Failures of classical physics: blackbody radiation, photoelectric effect. What new assumption was used to resolve these issues?
8. General relationship between E, υ, and λ.
9. Calculate E per photon or E per mole of photons, calculate wavelength from E, etc.
10. Bohr model for H: allowed energy levels for the electron are:

$$E = -2.18 \times 10^{-18}\text{J}\left(\frac{1}{n^2}\right)$$

$$\Delta E_{e^-} = -2.18 \times 10^{-18}\text{J}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

11. What are possible values of n in the above equation?
12. $\Delta E_{electron} = E_{photon}$
13. Ground state vs. excited state
14. Energy levels get closer together as you go up.
15. emission vs. absorption (sign of ΔE)
16. Calculate E or λ of a certain transition (for H) – use above equation. Which transitions occur in the visible region?
17. Why are only certain wavelengths of light emitted/absorbed?
18. Understand the connection between what the electron is doing and the light wavelength involved.
19. Spectrum of H (for other atoms, the spectrum is more complicated.)
20. Why do different atoms emit different colors of light?
21. De Broglie equation: used for matter, not light

$$\lambda = \frac{h}{mu}$$

22. Uncertainty principle: we can't know where the electron is, but we can get info on the probability of finding it in a certain region.
23. Ideas behind quantum mechanics
24. What is a wavefunction? What is the physical significance of ψ^2?
25. Quantum numbers: n, l, m_l, m_s: possible values, physical meaning for each
26. Know the general shape of s, p, and d orbitals
27. Sketching orbitals – include nodes. n-l-1 radial/spherical nodes, l nodal planes/angular nodes

28. For the H atom, the energy of the orbitals depends only on n.
29. For "many-electron" atoms, the energy depends on n and l.
30. Electron spin is quantized: $m_s = +1/2$ or $-1/2$.
31. Pauli exclusion principle, Aufbau principle, Hund's rule (what are they?)
31. Ground state vs. excited state vs. impossible configurations
32. Order of energy: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, etc. Know this order of energy (use the periodic table)
33. Be able to write the electron configuration for any atom.
34. Arrow/orbital diagrams, spectroscopic notation, noble gas notation
35. Paramagnetic, diamagnetic, ferromagnetic – what is the difference? How can you tell the difference experimentally?
36. Exceptions to filling order: Cr, Cu, others.
37. Elements in the same group have similar chemical properties. Why?

Chapter 7

1. When was the periodic table developed? How was it arranged? Who got credit for it, and why?
2. Understand and be able to explain screening in your own words. Be able to use a plot of Ψ^2 vs. radius to point out which orbitals shield which other orbitals. What is effective nuclear charge? Which electrons are shielded the most? Least?
3. Size (radius) – know the trend (atoms get smaller as you go across → and larger as you go down). Be able to explain **why**. Include shielding in your explanation.
4. Sizes of ions: Na^+ is smaller than Na. Cl^- is larger than Cl. When comparing isoelectronic ions, the one with more protons is smaller. Be able to explain why.
5. Ionization energy – know the exact process this refers to. Be able to write the equation. IE gets larger as you go across → and smaller as you go down. Explain why. Be able to explain the exceptions. (Look at the electron configurations before and after.)
6. What does it mean if an atom has a high ionization energy?
7. First IE, second IE, third IE: be able to write the equations for the exact processes. Be able to explain the trend.
8. Electron affinity – know the exact process. Atoms have a greater electron affinity as you go across. Be able to explain exceptions.
9. Filled or half-filled p or d subshells are especially stable.
10. Electron configurations of ions:
    Main-group – like noble gases – filled shell
    Transition metal ions – lose electrons from the outermost s level first
11. Properties of metals, nonmetals, and metalloids
12. metal oxide + H_2O → metal hydroxide
13. metal oxide + acid → salt + water
14. nonmetal oxide + H_2O → acid

15.   nonmetal oxide + base → salt + water
16.   Given the reactants, be able to write the products of the above reactions
       (and then balance the equation.)
17.   General properties of groups 1A, 2A, 7A, 8A, hydrogen, oxygen, and sulfur

Chapter 8

1.   Valence electrons – what are they?
2.   Dot symbols of elements
3.   Octet rule –what is it? What is the significance?
4.   What is an ionic bond?
5.   Lattice energy – what process does it refer to? What does its magnitude
       tell you? How is it determined?
6.   Predicting relative lattice energies of ionic compounds: high charges and
       small ions give the highest lattice energies.
7.   Covalent vs. ionic bond, polar vs. nonpolar covalent bond – what are the
       electrons doing in each?
8.   Normal number of bonds for groups 4, 5, 6, and 7
9.   Electronegativity trend: en increases as you go across and as you go up.
       Know this electronegativity trend: F, O, N = Cl, Br, C = S = I, H = P
10.   Bond polarity is determined by the <u>difference</u> in electronegativity.
11.   Drawing Lewis structures - determine skeleton structure (remember the
       rule for oxoacids – H is never attached to the central atom) Make sure you
       have the correct total number of valence electrons. Satisfy the octet rule if
       possible. Know when it's OK for the octet rule not to be satisfied.
12.   Minimize formal charge, if possible. (Be able to calculate formal charge for
       each atom)
13.   Remember that Be and B can have less than an octet. C, N, O, and F must
       have octets. Elements in the third period and beyond can have expanded
       octets.
14.   Choosing between non-equivalent resonance structures: minimize formal
       charge, fewer atoms that have formal charge, smaller magnitude of charge,
       negative formal charge on the more electronegative atom or positive
       formal charge on the less electronegative atom.
15.   Be able to draw all equivalent resonance structures, if applicable. (What
       are resonance structures and why are they used?)
16.   Bond length vs. bond strength vs. bond order – how do they relate?
17.   Bond dissociation energies – know how to use them to estimate ΔH_{rxn}.
       Why is this ΔH_{rxn} an estimate? For an **exothermic** reaction: a negative
       ΔH it corresponds to forming <u>stronger</u> bonds overall. The products are
       more stable than the reactants.
18.   Understand how to draw Lewis structures for larger organic compounds.
       (Make sure each atom has its normal number of bonds and lone pairs.)

Use the "condensed structural formula" as a guide for writing the Lewis structure.

19. Given the skeleton structure for a long organic molecule, be able to fill in any multiple bonds and lone pairs.

Chapter 9

1. Predicting shapes of molecules using VSEPR (need Lewis structure): count the total number of groups of electrons on the central atom (remember that double and triple bonds count as one group of electrons), then determine the number of lone pairs on the central atom. Know all shapes and be able to sketch each one.

2. Predict shape and bond angles. (Remember, lone pairs take up more space than bonding pairs, and multiple bonds take up more space than single bonds.)

3. Be able to predict whether molecules are polar or nonpolar:
   a. Are the individual bonds polar?
   b. Is the molecule asymmetric in 3 dimensions? (need to know shape)
   If yes to both questions, it's polar. If no to either question, it's nonpolar.

4. Potential energy curve for H_2 – explain the meaning of the different parts of the curve. H_2 is more stable than two individual H atoms.

5. Valence-Bond Theory: a bond is a result of overlap of orbitals. Orbitals must <u>hybridize</u> to point in the correct direction.

6.
Hybridizations:

# groups e- on central atom	Hybrid orbital set	Bond angle	Shape (if no lone pairs)
2	sp	180°	Linear
3	sp²	120°	Trigonal planar
4	sp³	109.5°	Tetrahedral
5	sp³d	120°, 90°	Trigonal Bipyramidal
6	sp³d²	90°	Octahedral

7. "Describe the bonding using valence-bond theory":
   a. Get the Lewis structure. How many groups of electrons on the central atom?
   b. What hybridization corresponds?
   c. Draw arrow diagrams for the unhybridized and then the hybridized central atom. (If needed, adjust electrons for formal charge on the central atom). Indicate lone pairs, sigma bonds, and pi bonds.

8. Draw a picture showing overlap of hybrid orbitals on the central atom with the appropriate atomic orbitals on outer atoms. Label the types of orbitals overlapping.

9. Be able to explain the difference between sigma and pi bonds. Draw each.

10.   How can you tell whether a molecule has delocalized π orbitals? Any molecule or ion with a delocalized π orbital system is especially stable.

11.   Molecular orbital theory: atomic orbitals combine to form molecular orbitals. If the orbitals are added, you get a bonding MO which has a lower energy than the original orbitals. If the orbitals are subtracted, the result is an antibonding MO which has a higher energy than the original orbitals.

12.   Be able to explain the difference between a bonding and an antibonding molecular orbital. Be able to draw a picture of the electron density in each (for both sigma and pi molecular orbitals).

13.   Bond order = $0.5 \times (n_b - n_a)$

14.   Be able to write the MO arrow diagram and the MO electron configuration for any diatomic molecule or ion in periods 1 or 2. You will need to remember the order of filling of molecular orbitals. (σ_{2s}, σ^*_{2s}, π_{2p}, σ_{2p}, π^*_{2p}, σ^*_{2p}) Remember that for O and F the order of filling is slightly different (σ_{2s}, σ^*_{2s}, σ_{2p}, π_{2p}, π^*_{2p}, σ^*_{2p}).

15.   Be able to determine if diatomic molecules are paramagnetic or diamagnetic. Be able to calculate the bond order and use it to compare bond lengths, strengths, and stability.

Chapter 11

1.   Kinetic-Molecular Theory, differences between solids, liquids, and gases

2.   Intermolecular Forces: types and strengths
•ion-dipole   (ion + polar molecule)   strongest
•dipole-dipole   (2 polar molecules)
•hydrogen bonding   (H attached directly to N, O, or F)
•induced dipole-induced dipole (London dispersion forces) (all molecules)

3.   Understand that IMF's are much weaker than ionic or covalent bonds.

4.   Stronger IMF's lead to: higher ΔH_{vap}, higher bp, lower vp, higher surface tension, higher viscosity. Reason: Strong intermolecular forces mean that it is difficult to separate the molecules from each other.

5.   Comparing molecules' IMF's:
a. Pure substances: look at MM, polarity, ability to H-bond. If they have similar MM, then the polarity and ability to H-bond are most important. If very different MM, the one with the higher MM will often have the strongest IMF's.
b. Different substances: MM doesn't matter. Similarity matters most.

6.   Surface tension, viscosity (high IMF's - higher surface tension and viscosity. What else does viscosity depend on?)

7.   Explain the idea behind surface tension.

8.   Know the six possible phase changes and whether heat is absorbed or released for each. Remember that phase changes happen at a constant temperature.

9.   What is vapor pressure? What does it depend on, and how? What does "volatile" mean?

10. Be able to explain why vp increases as T increases.
11. Relationship between boiling point and atmospheric pressure, concept behind using a pressure cooker
12. Phase diagrams (know shape and regions):
    At a given P and T, what phase(s) is (are) stable?
    What is the bp or mp at a given pressure?
    What is the vp at a given temperature?
    How could you liquefy the substance (given P and T, phase diagram or critical point)?
    Will the substance sublime at a given pressure?
    Which is more dense, solid or liquid?
    What is the triple point? The critical point?
13. Be able to calculate the relative percent humidity of an air sample. Be able to determine the dew point under a certain set of conditions. (See handout)

Chapter 12

1. Crystalline vs. amorphous solid – what's the difference? Give an example of each.
2. Understand and be able to explain the differences between molecular solids, network solids, metallic solids, and ionic solids.
3. Given the formula of a solid, decide what type it is and describe the type of forces holding the particles together. Explain why diamond and graphite have such different properties.
4. Which solids are held together most strongly? Least strongly?
5. mp, conductivity, hardness comparisons of different types of solids

Chapter 13

1. solute, solvent, solution definitions
2. $\Delta H_{solution}$ can be either + or -.
3. Dissolving is a physical process, but in some cases dissolving can be the result of a chemical reaction.
4. Definition of solubility, in quantitative terms
5. Saturated, unsaturated, supersaturated – definitions. Be able to tell when a solution is saturated or unsaturated, given the solubility at a specific temperature.
6. Determining relative solubility: "like dissolves like". Look at similarity of the two substances. For organic molecules, think about polarity, ability to hydrogen bond, and length of the nonpolar section. (How long is the hydrocarbon part compared to the polar/hydrogen bonding part?) Molar mass does not matter.

7. Gas solubility depends on P of gas above liquid.
Henry's Law $S = kP$ (Application: carbonated drinks)
8. Solubility depends on T.
Gases: less soluble at higher T.   (Application: dissolved O_2 and "thermal pollution")
Solids: most are more soluble at higher T.
9. Concentration units: M, m, X_A, mass %, ppm, ppb
For each unit, know how to calculate it and how to rewrite it as a conversion factor. Be as specific as possible!
10. Be able to convert from m to M or M to m (given the density of solution)
11. Be able to convert from mass % to m (or from m to mass %)
12. Colligative properties: What are they? What do they depend on and what don't they depend on?
13. vapor pressure lowering: $P_A = X_A P_A^\circ$ vp solution < vp of pure solvent
remember- X_A is mole fraction of **solvent**!
If both components are volatile, $P = X_A P_A^\circ + X_B P_B^\circ$
14. Boiling point elevation: $\Delta T_b = K_b m$ bp solution > bp pure solvent
15. Freezing point depression: $\Delta T_f = K_f m$ fp solution < fp pure solvent
16. Osmotic pressure: $\pi = MRT$
17. Water will flow toward the solution which has a higher concentration of solute (to try to equalize the concentrations on each side of the membrane).
Application: cell membranes
hypotonic, hypertonic, and isotonic solutions: what will happen to cells?
18. All colligative properties depend on the concentration of solute <u>particles</u> in the solution. Ionic compounds split apart in solution. The number of particles they split into can be predicted from the formula. However, in real life, ion pairs form in the solution. This reduces the effective number of solute particles, and the effective concentration of solute particles can't be predicted exactly. The formation of ion pairs is more prevalent in more concentrated solutions.
19. Using colligative property data to determine molar mass of solute:
a. find moles of solute   b. put #g/#mol, divide to get # g/mole.
20. Colloids - not true solutions - particles are very large.
How do soaps work?