

Things to Know – Chem 1A – Exam 3

Chapter 6

1. Pauli exclusion principle, Aufbau principle, Hund's rule (what are they?)
2. Ground state vs. excited state vs. impossible configurations
3. Order of energy: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, etc. Know this order of energy (use the periodic table)
4. Be able to write the electron configuration for any atom.
5. Arrow/orbital diagrams, spectroscopic notation, noble gas notation
6. Paramagnetic, diamagnetic, ferromagnetic – what is the difference? How can you tell the difference experimentally?
7. Exceptions to filling order: Cr, Cu, others.
8. 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 \rightarrow 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 \rightarrow 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 + $\text{H}_2\text{O} \rightarrow$ metal hydroxide
13. metal hydroxide + acid \rightarrow salt + water
14. nonmetal oxide + $\text{H}_2\text{O} \rightarrow$ acid
15. nonmetal oxide + base \rightarrow 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. (Charges have a bigger effect on LE.) Be able to rank compounds in order of lattice energy and explain your reasoning.
7. Covalent vs. ionic bond, polar vs. nonpolar covalent bond – what are the electrons doing in each? How can you tell the difference?
8. Know the normal number of bonds and lone pairs for groups 4, 5, 6, and 7 (and realize that there are plenty of exceptions – these are preferences, not requirements).
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 difference in electronegativity. (larger difference, more polar.)
11. Drawing Lewis structures - determine skeleton structure (remember the rule for oxyacids – H is never attached to the central atom). Make sure you have the correct total number of valence electrons. Satisfy the octet rule. Know when it's OK for the octet rule not to be satisfied. Be able to draw the Lewis structure for anything.
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. In resonance structures, electrons are arranged differently, but atoms do not switch positions.
17. Bond length vs. bond strength vs. bond order – how do they relate?
18. 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 corresponds to forming stronger bonds overall. The products are more stable than the reactants. (opposite for endothermic)
19. Be able 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.
20. 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 hybridize to point in the correct direction.

Hybridizations:

# groups e- on central atom	Hybrid orbital set	Bond angle	Shape (if no lone pairs)
2	sp	180°	Linear
3	sp^2	120°	Trigonal planar
4	sp^3	109.5°	Tetrahedral

6. “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. Indicate lone pairs, sigma bonds, and pi bonds.
 - d. 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. Your picture should show the correct approximate bond angles and molecular shapes.
7. Be able to explain the difference between sigma and pi bonds. Draw each.
8. How can you tell whether a molecule has delocalized π orbitals? Any molecule or ion with a delocalized π orbital system is especially stable.
9. 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.
10. 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).
11. Bond order = $0.5 \times (n_b - n_a)$
12. 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}^*).
13. 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. Be able to calculate the energy needed for various phase changes.
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 (at the molecular level).
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.
4. Explain why diamond and graphite have such different properties.
5. Which solids are held together most strongly? Least strongly?
6. mp, conductivity, hardness comparisons of different types of solids

Additional Items:

1. Estimating values from a graph – include all certain digits and the first uncertain digit. Include units.
2. Given a graph of a straight line, calculate the slope. Do not use data points for the slope calculation, since there is often scatter in the data (so the data points are usually not exactly on the line).

Chapter 13

1. solute, solvent, solution definitions
2. $\Delta H_{\text{solution}}$ 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? How many hydrogen bonding groups are there per carbon atom?) 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 be able to rewrite it as a conversion factor. Be as specific as possible! Be able to use these conversion factors in a problem.
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: $P_A = X_A P_A^\circ$ $p_{\text{solution}} < p_{\text{pure solvent}}$
remember- X_A is mole fraction of **solvent**! (a common mistake is to use the mole fraction of the solute here.) Be able to do problems involving this equation (and the other equations in this section).
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 that 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 particles 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.
21. How do soaps work? Explain.