

Exam #2 Study Guide

Covering Course Book Units/Chapter 5-7

=== FOCUS ON CALCULATIONS ===

Exam will cover the math/calculations in chemistry of

- Avogadro's Number: atoms/molecules/particles
- moles / mass (grams) conversions
- molar mass of atoms, compounds, molecules & how to determine
- percent composition from mass AND from chemical formula
- calculating empirical formula (Lab 5A!) and EFM (empirical formula mass)
- percent water mass in hydrate compound
- Energy: conversion between **joule** and **calorie**
- Using the equation $q = c_p \times \Delta T \times m$: when given 3 of the 4, solve for the 4th
- Converting temperature (T) between Celsius (°C) and Kelvin (K)
- Understanding calculations on LAST PAGE (7.E) exercises section 7.5-7.6
- Conversions, Scientific Notation, Significant Digits, Decimal Places: too many still do NOT understand this part, so expect INTENSE homework drilling on this!!

=== CHAPTER 5 ===

- Avogadro's Number: particles, atoms, molecules
- Molar Mass: of atoms, molecules, how to determine
- Percent Composition: by mass, by use of chemical formula
- Empirical formula: what is elemental analysis, how to determine empirical formula: this was the purpose of Experiment 5A
- Percent of water by mass in a hydrate compound
- What is EFM?
- Determining molecular formula from empirical formula mass

=== CHAPTER 6 ===

- How do the patterns of the Periodic Table help us understand the properties of matter and of the giving stability and a lowest energy (ground) states
- The octet rule: why atoms lose or gain electrons to form stable energy states and acquire valence shell electron numbers that look like the noble gas elements, whether of the current period (PT row) or the previous period (PT row)
- The octet rule for formation of atoms as ions in ionic compounds
- The octet rule for formation of atoms as part of molecules in non-ionic compounds
- Lewis structures, use of dot diagrams
- What part of the atom does a Lewis structure represent: valence shell electrons of atoms

- Using Lewis structures to understand movement of electrons in formation of ionic compounds
 1. one electron FROM a metal cation TO a nonmetal anion
 2. two (or more) electrons FROM a metal cation TO two (or more) nonmetal anions
 3. two (or more) electrons FROM two (or more) metal cations TO one nonmetal anion
- What is the formula unit?
- Lewis structures in both ionic and covalent compounds
- The rules to build Lewis structures
 1. Get (sum) total number of electrons from ALL atoms in compound (molecule) AND **add** electrons if the compound/molecule is a NEGATIVELY charged ion by the number on the ion OR **subtract** electrons if the compound/molecule is a POSITIVELY charged ion by the number on the ion
 2. A. Arrange the atoms to show the central atom(s) and their terminal atoms around the central atom
B. keep track of the “unused” total of electrons in building the Lewis structure
 3. Add a single bond—a line or two dots for electrons—between all the atoms to be bonded at a minimum to form the molecule, usually between the central atom and the terminal atoms
 4. Making **octets**: for period 2 and 3 elements, they will want 8 electrons (4 pairs of electrons) in their valence shell. Particularly for atoms C, N, O, and the Group 17 halides. Terminal hydrogen (H) atoms get a duet, not octet. All these added electrons likely to be nonbonding (lone) pairs
 5. Leftover electrons: places electron pairs around the central atom(s) since it/they will likely need it
 6. A. Final check on octets: this is usually for the central atom, which may NOT have a complete octet. Until you get familiar with recognizing Lewis structures having full octet, draw circles around the dots of bonding and nonbonding pairs of ALL atoms to show that an octet is surrounding the atoms.
B. If an octet is missing on any central atom, move one or more lone pair(s) from one or more of the terminal atoms to form a double bond, a triple bond, or two double bonds to complete the octet for ALL atoms
 7. Do a final check of all the manipulations
- Look at all the examples of creating Lewis structures: carbonate ion, water molecule, formaldehyde, ammonium ion
- Exercises to come: acetylene (C_2H_2), propadiene ($\text{H}_2\text{C}=\text{C}=\text{CH}_2$)
- Resonance
 1. What is it? How can carbonate CO_3^{2-} , ozone O_3 , nitrate (NO_3^-) be used as an example of it.
 2. Bond lengths: how are bond lengths different for single, double, triple bonds?
 3. How does analyzing bond lengths prove resonance?
 4. What is bond/electron pair delocalization? What does the term “half-bond” refer to and how does it apply to understanding resonance

- What is the purpose of formal charge? To what does formal charge apply to? How can knowing to formal charge be used to find the ionic charge, if any, on a molecule (as a polyatomic cation or polyatomic anion)? How does formal charge demonstrate the “ownership” of electrons in bonds and lone pairs?
- What are the exceptions to the octet rule?
 - A. What is a sextet or quartet? What central atoms of molecules show octet rule exceptions?
 - B. What molecules have an odd number of electrons in their Lewis structures? What does an odd number of total electrons imply about these molecules?
- VSEPR
 - A) what is it? What does the acronym stand for?
 - B) What feature of a molecule does it affect (two words, one being with “s” and other begins with “g”, used to describe same thing)?
 - C) What are the two “components” of VSEPR?
 - D) What is an “electron group”? How does it affect VSEPR and the molecule?
 - E) Know the difference between geometry affected by electron groups and the presence of all atoms affecting geometry of molecule
- Molecular geometry:
 - A) What geometry would a molecule with two (terminal) electron groups around a central atom have? What bond angle?
 - B) Three terminal electron group molecules with three atoms and two atoms + lone pair: what geometry? What bond angles?
 - C) Four terminal electron group molecules: four atom around central atom bond angle and geometry, three atoms + 1 lone pair, 2 atoms + 2 lone pairs
 - D) How does a nonbonding (lone) pair have an effect in VSEPR and affecting bond angles where terminal atoms exist in the molecule?
 - E) what do terms “linear”, “trigonal planar”, “bent”, “tetrahedral”, and “trigonal pyramidal” mean, and what features of a molecule should exist for these molecular geometries?
- Shape prediction: given a molecule, what analysis can you use to predict shape/geometry around a central atom, or even more than one central atom that will have a certain geometry?
- What is BOND polarity? What is electronegativity? How can differences in electronegativity in two bonded atoms be used to understand the determination of BOND polarity
- How do differences in electronegativity classify types of bonding between atoms: ionic and covalent, with the latter subdivided into polar and nonpolar covalent; therefore three types?
- What examples of nonpolar covalent, polar covalent, and ionic can you recognize and even give?
- What is MOLECULAR polarity (and how is it different from BOND polarity)?
- How do the features of symmetry and asymmetry in molecular structure and its geometry make a molecule polar or nonpolar, and how does your ability to show

Lewis structures and the effects of VSEPR contribute to understanding if a molecule is polar or nonpolar

- How can you indicate/symbolize both BOND and MOLECULAR polarity using the delta-plus or delta-positive (δ^+) and delta-minus or delta-negative (δ^-) indicators, or the crossed-arrow symbols? What part of the polarity of a molecule does the crossed end of the crossed-arrow diagram show, and what does the arrow head show?
- Features contributing to molecular (not bond) polarity: asymmetry, bonded atoms with large differences in electronegativity, the presence of nonbonding (lone) pairs
- Dipole: a quantity which is a vector. A vector has two characteristics: magnitude (reflective of electronegativity difference) and direction
- CORRECTION/CLARIFICATION TO LECTURE: A “dipole” as the word alone is a vector. But the “dipole moment” is NOT a vector: it is the magnitude part alone, and has no direction characteristic. No student will be penalized if this somehow becomes an exam question. The lecture slide will be fixed.
- Dipoles apply to bonds (between atoms) and also to molecules as a whole. The molecule as a whole dipole is usually the sum of the individual bond vectors. In mathematics, the summation of vectors involves the head-to-tail placement of the dipoles (vectors) representing the individual bond dipoles and then drawing a net or resultant vector from the tail of the first dipole (vector) to the head of the last dipole vector placed.
- Understand examples of polar (H-X, where X is Group 17; NH_3 , CHCl_3) and nonpolar (BF_3 , CCl_4) molecules. What makes them polar or not polar? Look at the books 5 steps also for assessing polar or nonpolar in molecules
- What are the properties of polar molecules: dipole alignment in external electric field (created by [high] voltage); higher boiling points for liquids, higher melting points for solids; solubility (ability to dissolve) in polar solvents like water
- INTERMOLECULAR FORCES
- Electrostatic-type weak forces between molecules that make these molecules/compounds capable of being liquids and solids
- Dipole-dipole interactions: a type of intermolecular force in which molecules have a (permanent) dipole because they are polar (type) molecules. The interaction between these molecules involves the molecules orienting or aligning themselves to maximize attraction between the opposite partial charge character of the dipole and to minimize repulsive forces in the dipole-dipole interactions. These arrangements enable solids to form
- In liquids formed by molecule dipole-dipole interactions, molecules move more freely and change orientations, but most of the time, the interactions will be more attractive and less repulsive to retain the property of being liquid
- Dipole-dipole interactions are weak in that the interaction (attraction or repulsion) will fall off if any distance occurs during movement
- Dipole-dipole interactions occur in molecules that have polarity (seen in dipole moments), not in nonpolar molecules. These interactions can increase melting and

boiling points showing the bulk of molecules will want to be solids or liquids. There is a direct correlation between molecular polarity (non-zero dipole moment) and higher boiling points and other physical properties

- London Dispersion Forces (LDF): an interaction which explains why/how nonpolar molecules/compounds can become liquids in the absence of a molecular dipole.
- LDF is thought to be caused or to occur because of electrons in orbit with atoms being in constant movement, particularly the outer/valence shell. This movement creates charge separation (distancing) of a very instantaneous and momentary nature, and a charge separation is a dipole, no matter how weak in magnitude. This momentary dipoles **induces** a momentary dipole in a neighboring molecule, and that induced dipole induces another dipole in yet another molecule. This development of a series of induced dipoles in numerous molecules can create attractive forces no matter how weak to enable nonpolar molecules in bulk to make a liquid phase and even solid phase possible, although at very low temperatures (well below room temperature), and having physical properties like a measurable boiling point
- Very large atoms like Xe are more likely to show these instantaneous dipoles which are electrons densely forming on one end of the atomic sphere because valence shell electrons are shielded by inner (core) shell electrons from the nucleus: the high positive charge of a nucleus would inhibit these momentary dipole formations. The ability of momentary dipoles to induce other dipoles is called polarizability
- As the mass of a molecule increases, it has a greater potential of the intermolecular forces causing molecules to form a liquid as opposed to staying as a gas. This is because the larger mass creates more points for intermolecular interactions, these points being dipoles or by LDF
- The molecular shape also is a factor in creating more points of intermolecular force interactions. The shape with a smaller surface area—such as a molecule that is more spherical than linear or rod-like—will reduce interactions. These more spherically shaped molecules will be less likely to be liquids, and have lower boiling points, compared to the more rod-like linear molecules.
- Hydrogen Bonding (H-bonding): this type of intermolecular force is stronger than any of the previous intermolecular forces mentioned. It is called H-bonding because the hydrogen atom the feature of this bond. The H atom is bonded in a polar covalent bond to an atom with a greater electronegativity (usually N, O, and F atoms; but also S and Cl atoms as well). There is a strong bond dipole that gives the H atom significantly positively charged character. This positively charged character of the H atom allows the H atom to be attracted often to the nonbonding (lone) pair electrons belonging to a N, O, F, S, or Cl atom on a neighboring molecule, since lone pairs are by nature highly negatively charged. The N/O/F/S/Cl atoms that are polarly covalently bonded to the H atom are called “hydrogen bond donors”, while the lone pair electrons of the other atom become the “hydrogen bond acceptor”. The H-bonding interaction happens with numerous molecules, creating a bridging-like interaction. Hydrogen bonding is what makes H₂O into the liquid water and into the solid ice at the appropriate temperatures

=== CHAPTER 7 ===

- Properties of solids: the particles (atoms, molecules) have fixed positions with little to no motion; there is a definite shape & volume. Solids are hard: crystals or rock, but can be soft, such as fats or wax. Solids are usually brittle, and crystals are often 3D arrays of positive and negative ions (ionic solids). Sometimes these solids do have crystalline arrays, but can be crushed into amorphous powders
- Liquids have properties of being able to overcome
- Cohesion: force/attraction of molecules for themselves, of their kind. What makes a water droplet spherical and not spread out
- Adhesion: force/attraction of molecules for other types/kinds of molecules: what makes water want to wick on to a paper towel
- Surface Tension: the resistance by a liquid's surface to an external force that makes molecules minimize their surface area, with molecules having strong cohesion. Measured in J/m^2 (a measure of work [energy] to increase a surface area against cohesive forces) or dyne/cm (a one-dimensional line of force at the boundary of liquid [like a needle that floats on water])
- Capillary action (capillarity): movement of liquid through or along the surface of solid material, caused by a combination of cohesion, adhesion, and the surface tension of the liquid, with forces that can even overcome gravity.
- Water shows strong adhesion to walls of glass because glass has silanol (hydroxyl) groups that interact through H-bonding; water climbs above its level in the container and forms a concave meniscus
- Mercury in contrast because of its metallic bonding nature will favor cohesion and not interact with glass hydroxyls in an adhesive way. In the capillary, mercury will not climb to the level of its container, and the surface tension promotes a convex meniscus
- Viscosity is the resistance of any liquid (fluid) to flow; viscosity creates more internal friction or a "thickness" to the liquid. A device called a viscometer measures the time it takes for any liquid to flow through tubing with gravity
- Chemicals with certain functional groups (like hydroxyl [-OH] groups increasing H-bonding) that increase the strength of intermolecular forces can increase viscosity. Adding another -OH group to ethanol, which already has one -OH group, will increase viscosity by 15 times
- Crystalline Solids: there are 4 classes
 1. Ionic: formed classically by the ionic compounds (metal monatomic cations with nonmetal monatomic anions but also including polyatomic types). Crystals are hard, brittle, have high melting points. Not electrically conductive as crystal solids, but the molten (liquid) form conducts electricity and especially when the ions dissolved in aqueous solution excellent conductive
 2. Metallic: bars and ingots of those shiny metals made of gold, iron, sodium, potassium, and all the atoms of the many elements we call metals on the metal side of the Periodic Table are held together by this type of **metallic bonding**. The atoms are all generally cationic (positively charged) as they release the valence

shell electrons to move freely, especially when two electrodes (one positive, one negative) are touched to the bar/ingot with an external voltage: this causes a “sea” of electrons to move in an electric current making the solid metal bar/ingot electrically conductive. These solid forms of the pure metal atoms will have high melting points that are related to the number of valence electrons they can donate: aluminum (3 valence electrons) has a higher melting point than magnesium (2 valence electrons) which has a higher melting point than sodium (1 valence electron) as an example. Also the melting point of a solid metal will be higher with the atoms of elements having a smaller atomic radius: therefore pure lithium (Li) with smaller atomic radius in its Group has a higher melting point than pure sodium (Na) metal solid which has a higher melting point than pure potassium (K) metal solid.

3. Covalent Network: these solids have their atoms bonded by covalent bonds and are not ionic or composed of ions (cations or anions) at all like in metal cation-nonmetal anion crystal solids. The classic example is when forms of pure carbon bond into solids like diamond and graphite: these have different molecular geometries as nonmetal single elements solids (diamond is tetrahedral pure carbon in a network, graphite is composed of sheets or layers of carbon bonded in a trigonal planar geometry. Other examples of covalent network solids those with more than one element like silicon carbide (SiC) which is called carborundum or silicon dioxide (SiO₂) called quartz. None of these solids will be electrical conductors
 4. Molecular: these are basically solids (solid forms) of molecules that often can have liquid or gaseous phases. They form solids because of their ability to have stronger intermolecular forces of various kinds (hydrogen bonding, dipole-dipole interactions, dipole-induced dipole or LDF interactions) at temperatures that are usual, but which can also artificially be created. Example molecules are H₂O (forming ice), I₂ (iodine crystals), dry ice (CO₂) at -77°C.
- In talking about energy and work (work is a form of energy as a force being applied over a distance), and in talking about the conservation of energy, we have to talk about the **universe** being composed of a **system**, which is the thing under study or in observation, and the **surroundings**, which interacts with the system in which energy is added to or taken from the system, since energy is conserved and cannot be created or destroyed
 - Reactions of compounds (atoms, molecules) including changes of phases (solid to liquid to gas) will involve the absorption of energy INTO the system FROM the surround and these are endothermic (energy [often as heat] from surroundings to system) or exothermic (energy from system to surroundings).
 - Be able to understand whether the boiling of water, the burning of gasoline, the formation of ice on a pond, the vaporization (evaporation) of water [different from the boiling of water!], the condensation of water as a vapor, or the melting of gold requires the INPUT of energy to cause that change in matter or the RELEASE of energy. Is it endothermic or exothermic?

- Understand what kinetic energy is (energy applied to a mass to cause motion), potential energy (or “stored energy”) which is energy related to a mass having a certain position in which a force is applied to it that could move it by a distance, or where the energy is because of the composition of the matter, in bonds between atoms or molecules that can be broken in chemical reactions (this is called chemical potential energy), releasing the energy. The burning of fuel of any kind (paper, gasoline) is the release of chemical potential energy.
- Energy is a quantity, having a number and units of calories or joules. 1 calorie is the energy required to heat 1 gram of water (H₂O as a liquid) by 1°C. All else is related to that standard. The joule can be determined using the conversion factor that 1 calorie (cal) = 4.184 joule (J)
- Heat capacity: the amount of heat energy (q) as measured in joules or calories that raises the temperature of any substance by 1°C or 1K.
- Specific heat capacity: the amount of heat energy (q) required to raise 1 gram of a substance by 1°C. Note the difference in definition of heat capacity and specific heat capacity: the latter specifies the amount of a substance (1 gram) but the other did not specify the amount as a mass...think about it.
- The equation you should know: $q = c_p \times m \times \Delta T$
- This is heat/energy (in cal or J) equals the product of the specific heat capacity (in cal or J per gram per °C or K), the mass (in grams and IS a measurable quantity), and the difference between two temperatures, an initial temp and final temp. This temp difference that you calculate also becomes a measurable quantity. The heat (q) value can be a measurable quantity. Of those four values, three of them are measurable quantities: and measurable quantities means you have to be careful of significant digits for multiplication/division operations and decimal places for addition/subtraction operations!
- You will be asked problems using this equation. You will be given 3 of the values, find the 4th using algebra. Keep track of significant digits
- Phase transitions: there are terms (nouns/verbs) that are associated with changes of phase. Understand and know them
- Enthalpies and differences in enthalpies (ΔH). ΔH_{vap} as heat of vaporization and phase from liquid to gas, ΔH_{fus} as heat of fusion and phase from solid to liquid, ΔH_{sub} as heat of sublimation and phase from solid to gas
- Understand the difference between boiling and evaporation and vaporization. They are different. And condensation
- What are heating and cooling curves: what do they demonstrate? Energy added to matter in the solid, liquid, and gaseous phase can involve changes in temperature to the matter. But at the points of temperature where phases change from solid to liquid and liquid to gas (and in the reverse), energy is added to the system but there is NO temperature change (it is isothermal): at these temperatures (which are melting and boiling points), the energy is used to cause the phase change and while occurring at these unchanging temperatures.
- In H₂O in going from ice to water to steam, there are five parts to the heating energy input curve: (i) raising/changing the temperature of the ice through heat application,

(ii) the melting of ice at 0°C (no temp change) to water, (iii) the change in temperature of water, (iv) the boiling of water at 100°C (no temp change) to steam, and (iv) the (super)heating of steam. This is true for all substances: atoms or molecules