Review Sheet for Final Exam

Mole- amount of a substance
Molecular Weight- mass in grams per mole
Mass number- number of protons + number of neutrons
Atomic number- number of protons = number of electrons
Element- matter composed of only one type of atom
Compound- matter composed of 2 or more kinds of atoms in definite proportion
Molecule- matter composed of 2 or more atoms covalently bonded together
Anion- a negatively charged ion
Cation- a positively charged ion
Salt- any compound formed by an anion and a cation

Molarity (M) - moles of solute/ 1 Liter of solution  (mol/ V)
   mol= VM

Density (g/mL) - mass/ volume

Light Emissions- each element has a unique line spectrum.
-relationship between light and energy
   E= hv
   c= λν
   E= hc/λ

E= energy (J); h= Planck’s constant (6.62x10^{-34} Js); v= frequency (s^{-1});
c= speed of light (3.00x 10^{8} m/s); λ= wavelength (m)
- visible light
   450-700nm on a spectrum
   The higher the wavelength, the lower the energy, the lower the wavelength, the higher the energy.

**electromagnetic spectrum**
Electrons absorb energy to jump up to a higher energy level because it is harder to keep an electron farther away from the nucleus so more energy is needed. This causes certain energies of light to be absorbed. When the electron wants to go back to a lower energy level, it may emit energy in the form of a photon.

Colors that we see are colors that are being emitted or the colors being reflected. If a color is emitted, that means that the other colors are being absorbed and there should be a peak on the spectrum at those colors wavelength values.

**-Photosynthesis**

When light strikes a molecule of chlorophyll, it absorbs energy and the electrons are excited and move to a higher energy level. The electron relaxes and releases energy in the form of chemical potential energy. It makes glucose and this allows it to get rid of the extra energy so that it can return to its ground state.

## Orbitals and Quantum Numbers

- There are three quantum numbers, n, l, and ml with integer values.
  - n, principal quantum number = 1, 2, 3…
    1. Determines orbital energy and specifies the electron shell.
    2. Defines orbital size.
    3. In atoms with more than one electron, two or more electrons can have the same n value…they have the same electron shell or electron level.
    4. Describes the shell that the electron is assigned to in an atom.
  - l, angular momentum quantum number = 0, 1, 2, 3…, n-1
    1. Gives orbital’s subshell.
    2. Determines orbital’s shape.
    3. Describes the subshell within the shell.
  - ml, magnetic quantum number = integers from –l to l, including 0.
    1. Gives the orbital’s orientation in space.
    2. Number of ml values specifies the number of orientations that exist for that subshell and the number of orbitals in the subshell.
    3. Describes the orientation of the orbital within the subshell.

<table>
<thead>
<tr>
<th>Value of n</th>
<th>Possible values of l</th>
<th>Possible values of ml</th>
<th>Corresponding Orbital Shape</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0</td>
<td>0</td>
<td>s</td>
</tr>
<tr>
<td>2</td>
<td>1</td>
<td>-1,0,1</td>
<td>p</td>
</tr>
<tr>
<td>3</td>
<td>2</td>
<td>-2,-1,0,1,2</td>
<td>d</td>
</tr>
<tr>
<td>4</td>
<td>3</td>
<td>-3,-2,-1,0,1,2,3</td>
<td>f</td>
</tr>
</tbody>
</table>

**n** = the number of subshells in a shell

2l+1 = number of ml values

n^2 = the number of orbitals in shell
A fourth quantum number $m_s$, was introduced to represent the spin of the electron. Its value is either $\frac{1}{2}$ or $-\frac{1}{2}$.

Nodal planes or surfaces are imaginary planes where there is zero probability of finding an electron. Nodal planes can be found in between areas of opposite shading and increase energy of an orbital.

Electron Configurations

- Lowest energy levels are filled first because ground state is the lowest energy configuration.
- Use either the diagonal rule or the periodic table method to find the order orbitals fill.
- Place electrons in orbitals using atomic numbers remembering that “s” orbitals can hold up to 2 electrons, “p” orbitals can hold up to 6 electrons, and “d” orbitals can hold up to 10.
- Valence electrons-outer most electrons. These electrons like to be as spread out as possible.
  
  Ex. Carbon =
  
  \[
  \text{Ex. Carbon = } [\text{He}]2s^22p^2
  \]

- Nobel gas cores can be used to shorten the configuration.
  Using the noble gas closest to the element (with a lower atomic number), start configuration from there.
  
  Ex. Carbon = [He]2s^22p^2

- Hund’s Rule- electron configuration will maximize the number of unpaired electrons. The electrons will only pair up when there are no longer any empty orbitals available in that subshell.
  - Paramagnetic-unpaired electrons (react to magnetic fields)
  - Diamagnetic- all paired electrons (don’t react)

- Half filled and fully filled subshells are the most stable (elements in “d” block under Cr and Cu).

Periodic Trends
- Ionization Energy- the energy to remove an electron
  Increases across a period left to right and decreases down a group because the positive nuclear charge is greater across a period so it holds stronger and the distance from the nucleus is greater down a group because of the number of orbitals.

- Atomic Radius-size of atom
  Decreases across a period left to right and increases down a group because as you go across a period, the positive nuclear charge is stronger so it holds the electrons closer to it making it smaller and as you go down a period, more energy levels are added making it bigger.

- Electron Affinity- the energy given off when gaining an electron
  Increases across a period left to right and decreases down a period because as you go across a period, the effective nuclear charge increases making it harder to ionize the atom, which makes it more attractive to electrons (proportionally related to ionization energy).
  - Effective Nuclear Charge- nuclear charge experienced by a specific electron in a multi electron system. Typical \( Z_{\text{eff}} \) for valence electrons is slightly greater than the number of valence electrons.

Electronegativity(x)- the ability of an atom to attract electrons in a bond.
  Can’t be measured, only inferred
  Increases across a period and decreases down a group.
  - Cs is the lowest (0.8) and F is the highest (4.0)
To test for polarity, find the difference between the electronegativities of the two atoms in the bond…
- $\Delta x \leq 0.4$, then there is a non-polar bond
- $0.5 \leq \Delta x \leq 1.5$, then the bond is a covalent polar bond
- $\Delta x \geq 1.5$, then there is an ionic bond
(Carbon – Hydrogen bonds are always non-polar bonds)

Drawing Lewis Dot Structures
1. Count the valence electrons on the individual atoms (remember charges).
2. Place the least electronegative atom in the center (unless H) and symmetrically arrange other atoms around it.
3. Connect central atom with outer atoms using dashes to represent 2 electrons.
4. Draw in lone pairs around the outer atoms, satisfying the octet rule.
5. If there are still electrons unaccounted for, place them as lone pairs around the central atom (Octet rule can be broken if the central atom has a valence d sub-shell).
6. Calculate formal charges of atoms individually using the equation
   \[ FC = (#\text{of valence e-}) - (#\text{of bonds}) - 2(#\text{of lone pairs}) \]
7. Rearrange bonds creating double bonds, etc. if the central atom has a positive charge. If charge can’t be minimized, put negative charge on most electronegative atoms.

- Resonance Structures- represents bonding in a molecule of ion when a single Lewis structure doesn’t completely demonstrate bonding.
  -represented using double headed arrows.
  -usually where double bonds are present.
  -resonance forms differ ONLY by electron location

Valence Shell Electron Pair Repulsion Theory (VSEPR)- used to draw accurate shapes of molecules.

<table>
<thead>
<tr>
<th># of groups</th>
<th>Molecular Shape</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Linear (180°)</td>
</tr>
<tr>
<td>3</td>
<td>Trigonal Planar (120°)</td>
</tr>
<tr>
<td>4</td>
<td>Tetrahedral (109.5°)</td>
</tr>
<tr>
<td>5</td>
<td>Trigonal Bipyramidal (120° and 90°)</td>
</tr>
<tr>
<td>6</td>
<td>Octahedral (90°)</td>
</tr>
</tbody>
</table>

-lone pairs can affect the shape of the molecule.

<table>
<thead>
<tr>
<th># of groups</th>
<th># of lone pairs</th>
<th>Molecular Shape</th>
</tr>
</thead>
<tbody>
<tr>
<td>3</td>
<td>1</td>
<td>Bent (&lt;120°)</td>
</tr>
<tr>
<td>4</td>
<td>1</td>
<td>Trigonal Pyramidal (&lt;109.5°)</td>
</tr>
<tr>
<td>4</td>
<td>2</td>
<td>Bent</td>
</tr>
<tr>
<td>5</td>
<td>1</td>
<td>See Saw</td>
</tr>
<tr>
<td>5</td>
<td>2</td>
<td>T- Shaped</td>
</tr>
<tr>
<td>5</td>
<td>3</td>
<td>Linear</td>
</tr>
<tr>
<td>6</td>
<td>1</td>
<td>Square Pyramidal</td>
</tr>
</tbody>
</table>
Acids and Bases - the two major classes of compounds; they react with each other to gain resonance. Polarity of bonds are critical for reactions.

-Acids- substances that increases the concentration of H+ in an aqueous solution.
- Strong Acids- HCl, HBr, HI, HNO₃, HClO₄, H₂SO₄

-Bases- substances that when dissolved in water increases hydroxide concentration.
- Bitter taste and slippery
- Strong Bases- Group 1 bases (not including H)

- Reaction- acid gives up H, bases take H
- Amphoteric- may act as a base or an acid, i.e. H₂O

Bond Order- # of bonds between 2 specific atoms. If the molecule has resonance, count bonds and divide by the number of resonance structures.

Bond Length- distance between two bonded atoms.
- Depends on atoms involved (atomic radii) and bond order
- Higher bond order, shorter bond, i.e. Triple bonds have the shortest bonds

Hybridization- atomic orbitals mix to form new atomic orbitals that overlap to better show the shape of the molecule.

<table>
<thead>
<tr>
<th># of electron groups</th>
<th>Hybrid Orbital Type</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>$sp^3$</td>
</tr>
<tr>
<td>3</td>
<td>$sp^2$</td>
</tr>
<tr>
<td>4</td>
<td>$sp^3$</td>
</tr>
<tr>
<td>5</td>
<td>$sp^3d$</td>
</tr>
<tr>
<td>6</td>
<td>$sp^3d^2$</td>
</tr>
</tbody>
</table>

- Single bonds = sigma bond
- Double bonds = sigma and pi bond
- Triple bonds = sigma bond and 2 pi bonds

Naming
1. Binary Ionic Compounds- cation first, anion last (total charge of 0)
- Cation- name of element
- Anion- element name with –ide ending
Ex. LiBr- lithium bromide
Ex. Na₂O- sodium Oxide
- Hydrates- regular naming with (prefix)hydrate on the end. If we don’t know the number of waters, it’s just hydrate.
2. Polyoxoanions
- Chlorate- ClO₃⁻
- Chlorite- ClO₂⁻ loses an O
- Hypochlorite- ClO⁻ loses another O
- Chloride- Cl⁻ loses another O
- PerChlorate- ClO₄⁻ gains an O

3. Acids-based on anionic piece
- if it ends in –ate, it becomes –ic acid
- if it ends in –ite, it becomes –ous acid
- if it has a prefix, it stays
- if it ends in –ide, it becomes hydro ___ ic acid

4. Binary non-metal
- less electronegative atom first
- second ends in –ide
- add prefixes to them based on suffix, but if the first only has 1, leave it off.

(memorize ions worksheet)

Organic Chemistry
-Hydrocarbons-molecules containing carbon and hydrogen only.
  Formula for saturated hydrocarbons- CₙH₂ₙ₊₂ (if no double bonds or rings)
  -Naming straight hydrocarbons

<table>
<thead>
<tr>
<th># of C’s</th>
<th>Root Name</th>
<th>Suffix</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Meth-</td>
<td>-ane</td>
</tr>
<tr>
<td>2</td>
<td>Eth-</td>
<td>1</td>
</tr>
<tr>
<td>3</td>
<td>Prop-</td>
<td>1</td>
</tr>
<tr>
<td>4</td>
<td>But-</td>
<td>1</td>
</tr>
<tr>
<td>5</td>
<td>Pent-</td>
<td>1</td>
</tr>
<tr>
<td>6</td>
<td>Hex-</td>
<td>1</td>
</tr>
<tr>
<td>7</td>
<td>Hept-</td>
<td>1</td>
</tr>
<tr>
<td>8</td>
<td>Oct-</td>
<td>1</td>
</tr>
<tr>
<td>9</td>
<td>Non-</td>
<td>1</td>
</tr>
<tr>
<td>10</td>
<td>Dec-</td>
<td>1</td>
</tr>
</tbody>
</table>
-count the number of carbons in the longest carbon chain and add the appropriate suffix—ane for single bonds, -ene for double, and –yne for triple.

-to show branches, put the number of the carbon that has the branch, the number of C’s coming off of it and the suffix –yl.

-put the names in alphabetical order

Ex. 3ethyl,2methyl-pentane

-Saturation- only alkanes are saturated. Alkanes are all sp³ hybridized, alkenes are sp² hybridized and alkynes are sp hybridized.

Degree of unsaturation- \(\left[2 + 2\text{C} + \text{N} - \text{H} - \text{X}\right]/2\)

-number that you get is the number of either pi bonds or rings

-Functional Groups-common group of atoms in an organic molecule.

1. Aromatic Hydrocarbons- can be fairly stable because of resonance.

2. Alcohols- have a relatively high boiling point and tend to be polar.

-ending of –ol after prefix and must indicate which C has the OH attached. Ex. 1propen-4-ol

3. Ethers- oxygen attached to two carbons, low boiling point, not water soluble.

4. Amines- replace Hs on NH₃ with a C…basic so they react with acids, not water soluble and offensive odors.

-primary, secondary and tertiary refers to # of Cs.

5. Adehydes and Ketones- pleasant odors

6. Carboxylic Acids- water soluble

7. Esters- smell really good

-ended regularly, but followed by acetate

8. Amides- N attached to three C groups, one of which is C=O.

(memorize functional group diagrams)

Polymers-giant molecules made from monomers (small molecules)

2 methods to make polymers:

1. Condensation polymers-made by combining monomers and spitting out small molecules.

2. Addition polymers- break a double bond and form new sigma bonds.

-condensation reactions- both ends are reactive so that they can continue to react.

involves the substitution of a carbonyl
-natural polymers include DNA, protein, starch and rubber.

-addition reactions-break pi bonds and combine monomers with new sigma bonds

-Isomers-compounds with the same formula, but a different structure.
  1. structural isomers- different connectivity
  2. stereo isomers- same connectivity
     -geometric isomers. Ex. cis and trans
  3. optical isomers-isomers with non-super imposable mirror image- chiral
     (chiral- a C with 4 different groups attached)

-anti addition- adds from opposite sides because of the size of the atom results in a trans isomer.

-combustion Reactions-adds O2 to a hydrocarbon and the results in CO2 and H2O
-hydrolysis-reaction with water (mostly esters and amides)
   -hydrolysis of esters result in carboxylic acids and alcohols
   -hydrolysis of amides result in carboxylic acids and amines

Finding the Limiting Reactant
  1. Balance the equation
  2. Compare moles of each reactant to the product through a molar ratio
  3. Which ever produces the least amount of moles is the limiting reactant
     (grams can also be compared by using the molecular weight)

Finding the Empirical Formula-lowest number ratio of atoms in a moleule
Ex. CxHy + zO2→ xCO2 + yH2O…with grams of CO2 and H2O and grams of the organic sample given.
  1. Take grams of CO2 and divide by its molecular weight.
  2. Do a molar ratio between moles of CO2 and moles of C in CO2 to get moles of C.
  3. Repeat steps 1 and 2 with H2O to get moles of H
  4. Convert moles of C and moles of H into grams of both using the molecular weights of carbon and hydrogen.
  5. Set the grams of C, H and O equal to the total grams in the organic sample.
  6. Through subtraction, the grams of O can be found.
  7. convert grams of O into moles of O using molecular weight.
  8. Divide all three mole amounts by the smallest number.
  9. Whole numbers should be obtained, plug in for the empirical formula.

Percentage Yield
$$\text{Percentage Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$
Thermodynamics - the study of heat and work

1st Law of Thermodynamics - Total energy of the universe is constant...energy can’t be created or destroyed.

- Energy - the capacity to do work (E)

1. Thermal Energy - energy of random motion of molecules/atoms; can be measured in the form of temperature (kinetic).
   - Exothermic - process where system transfers E to surroundings.
   - Endothermic - process where system absorbs E from surroundings.

\[ q = m C \Delta T \]

\( q = \) Energy (J)
\( m = \) mass (g)
\( C = \) specific heat of material (J/gK or J/g°C)
\( \Delta T = \) change in temperature (°C or K)...Temp \( \text{final} \) – Temp \( \text{initial} \)

Because of the 1st law of thermodynamics: \( q_x + q_y = 0 \) or \( q_x = -q_y \)

- Soooo… \( m_x C_x \Delta T_x = - m_y C_y \Delta T_y \)

- To prevent energy loss into the environment, experiments are usually done in a calorimeter - an insulated device designed not to lose heat.
  - the calorimeter needs to be made account for in the equation…
  \( q_x + q_y + q_{\text{cal}} = 0 \)
  …however, the calorimeter is usually considered to be equal to zero if it is a coffee cup calorimeter.

- Energy is needed to change both temperature and to change state of matter so there are values for heat of fusion and heat of vaporization.
  - heat of fusion - the heat required to convert a substance from a solid at its melting point to a liquid.
  - heat of vaporization - the heat required to convert liquid at its boiling point to gas.
  So, \( q = n \Delta H_{\text{vap}} \) or \( q = n \Delta H_{\text{fus}} \) (This equation is generally applicable to any reaction.)

- Work is also a form of energy.
  \( \Delta E = \Delta q + \Delta w \)
  \( q = \) heat; \( w = \) work
  …because we have constant pressure, \( w = -P \Delta V \).

- Enthalpy (H) - the heat capacity of a substance at a constant pressure…we care about \( \Delta H \).
  - means that at a constant pressure, \( q = \Delta H \)...and that \( \Delta E = \Delta H - P \Delta V \).
  - if \( \Delta V \) is small, then we can consider \( \Delta E \) to be equal to \( \Delta H \) (\( \Delta E = \Delta H \)).
  - depends on the molar amounts of reactants and products so making 3 mol of something makes an enthalpy change that is three times as big as making 1 mol of it.

- Negative values of \( \Delta E \) and \( \Delta H \) specify that energy from an exothermic reaction; positive values mean that energy is from an endothermic reaction.
*State Functions- values defined only by final and initial states (represented by capital letters).

- Because $\Delta H$ is a state function, it is the same for a single step or multi step process…the value is the same no matter how you get it.

-In general, $\Delta H_{\text{rxn}} = \Sigma [\Delta H_f \text{ products}] - \Sigma [\Delta H_f \text{ reactants}]$
* If there is a coefficient in the equation, the $\Delta H_f$ must be multiplied by that coefficient before inserting it into the equation.
* All elements in their normal states have a $\Delta H_f$ of 0.
* These reactions usually occur because they will be able to get to a lower energy state afterwards.

2. Chemical Energy- energy associated with bonds (potential).

-Different bonds have their own $\Delta H$’s.

$\Delta H = E \text{ in} - E \text{ out}$  or  $E \text{ from broken bonds} - E \text{ from formed bonds.}$
- when bonds are broken, they have positive energy because it takes energy to break a bond. To form a bond, it takes energy, so the value is negative.

* $C \equiv C$ is the strongest bond and has the most bond energy.
  $C - C$ - 346, $C = C$ - 610, $C \equiv C$ - 835…in the case of carbon-carbon bonding, each additional bond is worth less energy.