

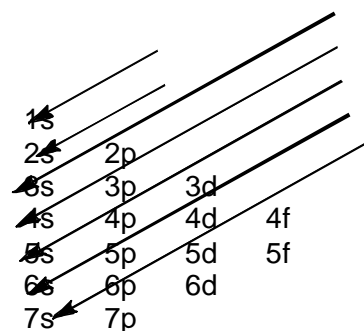
Exam 2 Review

Chapter 3.2-3.11 (25, 27,29, 35, 45, 49, 51, 55, 57, 59, 121, 123, 125, 12, 71, 73,79,81, 85, 93, 97, 99, 111, 113, 115, 133 and 135)

1. The photoelectric effect
 - A. Know the 4 properties of the photoelectric effect
 - B. Einstein's postulations
 - C. $E = h\nu$ where h is planck's constant $6.626 \times 10^{-34} \text{ J sec}$
 - D. Remember that the energy calculated in C is J/photon. Be able to determine the number of photons or total energy given the number of photons
 - E. $E_{\text{total}} = E_{\text{kin}} + E_{\text{binding}}$
2. Absorption and Emission Spectra
 - A. Absorption spectra is a measure of the visible light spectrum that the element absorbs
 - B. Emission spectra is the spectrum emitted when visible light is shown through an element
3. Bohr's Theory
 - A. e^- can only have specific (quantized) energy values
 - B. light is emitted as e^- moves from one energy level to a lower energy level
4. Properties shared by all electrons
 - A. de Broglie: If things that behave like a wave (light) have particle characteristics then things that behave like particles should also have wave characteristics
 - B. de Broglie equation: $\lambda_{\text{particle}} = h / mv$ where h is Planck's constant , m is mass in Kg and v is the speed
 - C. Heisenberg's uncertainty principle: The more accurately we know the position of an electron, the more uncertain we are of it's motion

Quantization and Quantum numbers

1. Each electron in an atom has a unique set of quantum numbers
2. Principle Quantum number (n) indexes energy and size and has positive integer values
3. Azimuthal quantum number (l)
 - A. indexes shape of atomic orbital and is limited by the value of n . It can have values of zero and positive integers $\leq (n-1)$
 - B. numerical value of l is linked to an orbital letter ie: $0=s, 1=p, 2=d, 3=f$
 - C. when naming orbitals use the number of n and the letter associated with l , ie: when $n=3$ and $l=1$ the orbital is $3p$
4. Magnetic quantum number (M_l) indexes direction. It is limited in its value by l , values can be negative and positive integers including the value of zero. As l increases by 1, 2 additional values of M_l become possible
5. Spin orientation quantum number (M_s): can either be up ($+\frac{1}{2}$) or down ($-\frac{1}{2}$)
6. Terminology when dealing with ground state electron configuration (terms to know)
 - a. Pauli Exclusion principle
 - b. Aufbau Principle
 - c. Degenerate orbitals
 - d. Isoelectric (Isoelectronic)
 - e. Conditions for Atomic ground state
 - f. Order of orbital filling



1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d

7. Electron Configuration
 - a. Know the difference between valence electrons and core electrons
 - b. Three ways to represent electron configuration
 1. list set of all quantum numbers
 2. short hand notation: Write the quantum number then orbital symbol with the number of electrons in each orbital superscripted

s max 2 e-	p max 6 e-
d max 10 e-	f max 14 e-

8. Diagrammatic representation: showing electrons as arrows
 - a. Hund's Rule
 - b. Near degenerate orbitals
 - c. Write the configuration of ions
 - d. Magnetic properties of atoms (diamagnetic and paramagnetic)

Chapter 4.1-4.5 (17, 23, 25, 27, 29, 31, 33, 35, 41, 43, 47, 49, 55, 57, 61, 63, 67, 69, 75, 77, 79, 83, 85, 87, 89, 91, 93, 95, 97, 99, 101, 103, 105, 107, 109 and 111)

1. Ionization Energy: the amount of energy required to completely remove an electron.
 - A. increases from left to right across a row of the periodic table
 - B. decreases from top to bottom down a column of the periodic table
2. Screening (Shielding) electrons cancel out part of the attraction of another electron to the nucleus.
 - A. As n increases the screening effect decreases
 - B. as l increases the screening effect of those electrons decrease
 - C. electrons in lower orbitals screen electrons in higher orbitals but not the other way around.
 - D. electrons in degenerate orbitals do not screen each other
 - A. Effective Nuclear Charge (Z_{eff}) The charge an electron feels from the nucleus with the consideration of screening.
 - B. Electron Affinity: measures the likelihood of an atom gaining an electron
3. Periodicity of Atomic properties
 - A. As principle quantum number increases the atomic orbitals become larger and less stable
 - B. Figure stability decreases from top to bottom down a column and increases left to right in a row of the periodic table
 - C. Atomic radii decreases from left to right in a row and increases from top to bottom of a column in the periodic table
 - D. Ionization Energy increases from left to right in a row and decreases from top to bottom in a column of the periodic table
 - E. Electron Affinity: measures the energy change when an electron is added to an atom. Becomes more negative from left to right across a row of the periodic table
 - F. Why do ions form (ie: why do we have Na^+ and not Na^{2+} ?)

Chapter 5.1, 5.2, 5.3, 5.5, 5.7, 5.8 and 5.9 (1, 11, 35, 37, 39, 41, 43, 45, 45, 47, 67, 69, 75, 77, 79, 83, 85, 87, 89, 91, 97, 103, 107, 109, 115, 119, 121, 127, 129 and 131)

1. know the difference between a molecule and a compound
2. Be able to write a molecular formula
3. know how to determine molar mass
4. Be able to calculate percent composition from molar mass
5. Be able to determine molecular formulas when given percent composition and molar mass

Bonding

1. Ionic Bonds: electrons are transferred from one species to another
 - A. usually between a metal and a non-metal
 - B. Lattice energy is the energy required to break an ionic bond
2. non-polar covalent bonds the equal sharing of electrons ie: F_2 , H_2 , N_2
3. Polar covalent bonds: unequal sharing of electrons- the more electronegative atom hogs the electron density
4. Bond Length (Bond distance): The distance between nuclei of bound atoms
5. Bond Energy: The energy required to break a bond
6. Octet Rule: electrons are shared or transferred between atoms in a molecule so that each atom has an octet of electrons
7. Exceptions to the octet rule
 - A. Hydrogen only forms one bond. It only shares 2 electrons
 - B. Beryllium only forms two bonds
 - C. Boron only forms 3 bonds
 - D. Atoms that can have d orbitals (elements in and above the 3rd energy level) can have more than 8 electrons around it.

Chapter 6 (23, 25, 29, 31, 35, 37, 39, 41, 53, 57, 59, 73 and 79)

Lewis Structures

1. Lewis Theory
2. Lewis Symbols
 - A. Only Valence electrons are shown
 - B. put 1 electron around the elemental symbol before doubling them up. This illustrates hunds rule
3. Lewis Structures
 - A. Single, double and triple bonds are all 1 region of electron density
 - B. Central atoms vs terminal atoms
 - C. Formal Charge = # of valence e – $\frac{1}{2}$ (# of bonding e) – # of non-bonding electrons
 - D. Resonance structures
 - E. Molecules want to have the lowest number of formal charges it possibly can and so atoms such as S and P will use their expanded octet to alleviate formal charges

Chapter 7.1-7.5 (7, 9, 11, 13, 19, 21, 35, 37, 39,43, 45, 47, 49, 79 and 83)

Molecular shapes

1. Be able to draw a Lewis structure and locate regions of electron density around a central atom
2. What does Steric number tell you? (you are responsible for steric numbers 2,3,4,5 and 6) you just need to memorize the tables in this se
3. Use VSEPR theory to determine molecular shape and give VSEPR notation
4. Bond angles
5. Be able to determine whether a molecule is polar or nonpolar

Hybridization

1. Valence Bond Theory: When atoms bond they do so by overlapping orbitals
2. Hybridization
 - A. Applies to all atoms except hydrogen
 - B. Identify which atomic orbitals are involved in forming a hybrid orbital
 - C. Know the relationship between hybrid orbitals and steric number
3. Types of bonds
 1. sigma bonds are single bonds that form between hybridized orbitals
 2. pi bonds are double or triple bonds that form from overlapping p orbitals or overlapping d-p orbitals
 3. effects of multiple bonds on bond length and bond energy
 4. Resonance structures and extended pi systems