***STUDY GUIDE EXAM 3***

***(LIST Most of the major concepts taught for Chapters 7,8,9 and 10)***

***Chapter 7 Electronic Structure of Atoms***

**wavelength** (λ): distance between successive peaks; generally given in m, cm, or nm

– Know 1 m = 109 nm

**Use equation: c =** λν

**speed of light, c: 2.998**×**108 m/s** (will be given)

**frequency** (ν): number of crests passing by a given point in 1 s; given in 1/s=hertz (Hz)

**wavelength** (λ): distance between successive peaks; generally given in m, cm, or nm

– Know 1 m = 109 nm

Relate energy, frequency, and wavelength:

**E = h**ν **= hc/**λ

**Dual Nature of the Electron**

– Electron can behave like a wave or a particle

**Planck and Quantum Theory**

– proposed energy is absorbed and emitted as bundles = **quanta**

– single bundle of energy = **quantum**

**Einstein and the Photoelectric Effect**

– Be able to describe the Photoelectric Effect

– Know this provided experimental evidence for light existing as particles = **photons**

**Bohr Theory of the Atom**

* **Electrons move in quantized orbits called “energy levels” around nucleus**
* **ground state:** electron in lowest energy level
* **excited state:** electron in higher energy level
* When atom absorbs energy, electron jumps from lower energy to higher energy level.
* When electron drops from a higher energy to a lower energy level, it releases energy, in some cases as light → atomic emission spectra
* Know limitations of the Bohr Model
* Recognize energy levels further from the nucleus are closer together
* Calculate the ∆E for an ni→ nf *(where ni > nf )* electronic transition within an hydrogen atom using the Bohr equation

$$∆E=\left(-2.18×10^{-18}J\right)\left(\frac{1}{n\_{f}^{2}}-\frac{1}{n\_{i}^{2}}\right)$$

$$\frac{1}{λ}=\left(1.097×10^{-2}\frac{1}{nm}\right)\left(\frac{1}{n\_{f}^{2}}-\frac{1}{n\_{i}^{2}}\right)$$

* Calculate the wavelength *(or frequency)* when given the energy and vice versa using the E = hν equation *(… this includes knowing the definition of and the mathematical relationships between λ (wavelength)) and ν (frequency).*

Wavelength (λ): distance between successive peaks; generally given in m, cm, or nm

– Know 1 m = 109 nm

Frequency (ν): number of crests passing by a given point in 1 s; given in 1/s =hertz (Hz)

$$ν=\frac{c}{λ}$$

* Discuss the failures of classical physics in explaining electronic structure and why a new model, i.e. quantum theory, was necessary to fill the gaps left by classical physics.
* Discuss the contributions of the following scientists to the development of quantum theory/quantum mechanics: Planck, Einstein, Bohr, de Broglie, Heisenberg, and Scrodinger. (This includes being able to state the success and the failure of Bohr’s postulates in explaining line spectra of the hydrogen atom/electronic structure.
* Be aware of the relationships between the contributions of Planck, Einstein, and Bohr (i.e energy is absorbed or released *( emitted)* in packets called quanta *(Planck)* or photons where E photon = hν and this theory is seen in the photoelectric effect *(Einstein)* and that the energy of electrons are quantized and the energy of an electronic transition can be calculated using the Bohr equation *(Bohr)).* ……. and relationships for de Broglie, Heisenberg, and Schrodinger (…. the electron is a standing matter wave that is not a point object but a series of points that extend into space but it is also a particle *(de Broglie, λ = h/m*v *)* ……… and the electron has no definite location, i.e ∆x for the hydrogen electron is greater than the size of the hydrogen atom *( Heisenberg)* ……..and that the standing matter wave can be described by a wave equation which can be solved for wave functions to give four quantum numbers *(Schrodinger)* )… and these quantum numbers represent the electronic address of an electron.

 Planck and Quantum Theory

Proposed energy is absorbed and emitted as bundles = quanta

– single bundle of energy = quantum

Einstein and the Photoelectric Effect

Be able to describe the Photoelectric Effect and calculate the threshold energy, total energy,kinetic energy and the velocity of the electron

 *de Broglie Matter wave*

Electrons have dualistic nature: sometimes behaving as particles and sometimes behaving as waves; deBroglie and  = h/mv;

Uncertainty principle: there is a limit as to how precise we can know the position and momentum of an electron in an atom: (x) x (mv) ≥ h/4π

Schrodinger’s wave equation can be solved to yield position probabilities and energies of electrons in atoms

* Give the possible values of the *l* and the *m*  quantum numbers when given the value of the principal number*(*

n = 1, 2, 3, …; l = 0, 1, …, n – 1; ml = - l to 0 to – l

* + Write the values of n, l and ml for each orbital in a specific p or d subshell. (This includes being aware of the specific characteristic of an orbital given by each quantum number.( i.e., n gives size and energy , *l* describes the shape or type of orbital ( s, p, d, f, etc. ), m*l* describes the orientation, e.g. px py and pz etc.) and ms identifies the electron in the orbital by its spin . You should also know that it requires four quantum numbers to identify an electron in an atom and that quantum numbers 1-3( i.e, n. *l*  and ml ) identify the orbital and the ms quantum number identifies the electron,
	+ Know the shapes of s,p,d and maybe f orbitals

***Chapter 8 (Periodic Properties of the Elements)***

* Use the periodic table to determine the number of valence electrons for main group elements *(This includes recognizing valence electrons are the electrons used in forming chemical bonds )*
* Recognize the correct filling order for electron in orbitals ( i.e aufbau principle ( low energy to high energy), Hund’s rule for degenerate orbitals ( fill empty orbitals first) , and Pauli exclusion principle ( maximum of two electrons per orbital with opposite spins)
* Write the condensed electron configuration for an atom. (*e.g the condensed electron configuration for Cl is [Ne]3s2 3p5,*
* *Be able to draw orbital diagrams of electron configurations*
* State the difference between the terms ground state and excited state when referring to electronic transitions in the hydrogen and multi electron atom.
* Write electron configurations for monoatomic ions .
* Graph , explain, or discuss the relationship between Zeff and atomic radius
* Use the periodic table to arrange a series of elements according to atomic radius, ionization energy, electron affinity and metallic character.
* Use the periodic table to arrange a series of monoatomic ions according to atomic radius.
* Compare the atomic radii of atoms and their ions
* Calculate the Zeff  for an atom where Zeff = Z – S.
* Explain the difference between ionization energy and electron affinity
* Graph the trend in ionization energy for elements with atomic numbers 1- 54 *(see graph at the end of the document )*
* Explain the large increase from I1 to I2 for the alkali metals where I1 to I2 are the first and second ionization energies.

***Chapter 9 Basic Concepts of Chemical Bonding***

* Recognize the characteristics of or the substances containing either of the three major types of bonds ie. ionic, covalent (polar covalent), and metallic (e.g What are the bond types in KCl ? KOH? SO2? HNO3? an iron rod ?
* Calculate the lattice energy of an ionic compound using the Born-Haber cycle.
* Lattice energy and Born-Haber cycle describes the energetics associated with the formation of one mole of an ionic compound from its elements
* Predict relative lattice energies
* Use the concepts of Columb’s law
* Understand the trends in bond length and strength
* Draw Lewis structures for molecular compounds or polyatomic ions.
* Draw acceptable resonance structures when give a molecule that has two or more resonance structures, e.g. write a resonance structure for CO32-. ( This includes being able to assign formal charge to determine the best structure
* formal charge is the charge associated with an atom in a molecule = # valence electrons - # electrons assigned to atom in the molecule; formal charge can lead to a more stable Lewis structure when several alternatives are found
* Recognize the difference between the terms resonance structure and resonance hybrid.
* Resonance structures: equivalent Lewis structures that include multiple bonds; bond strengths and bond distances are averaged in resonance scenarios; delocalized electrons are shared between many electrons in resonance structures
* Explain the difference between localized and delocalized electrons and determine if a set of pi electrons or unshared pair electrons are localized or delocalized, e.g. unshared pairs in H2O localized or delocalized? Pi bond in CH2=CH2 localized or delocalized? Bonds in benzene localized or delocalized? Are any electrons in O3 delocalized?
* Recognize the relationship between electronegativity values *( i.e values for two atoms)* and bond polarity ( one is used to calculate the other… Also this includes understanding the definition of polarity/dipole and recognizing that a polar bond is said to have a dipole )
* Classify a bond as either polar, nonpolar or ionic based on an evaluation using electronegativity values or by periodic trend. Also this includes being able to draw the polarization related symbols δ+ δ- and  and being able to predict the behavior or polar/nonpolar molecules in an electric field)
* Calculate the ΔHrxn using bond enthalpies. Also this includes recognizing that bond enthalpies will always have a positive sign because bond breaking is an endothermic process.

***Chapter 10 (Molecular Geometry and Bonding Theories)***

* Use VSEPR Theory to predict the orbital geometry and the molecular geometry ( i.e shape) of molecules or polyatomic ions and determine the approximate bond angle.
* Orbital or Electron domain geometries; linear, trigonal planar, tetrahedral, trigonal bipyramidal; and octahedral shapes
* Molecular shapes are determined by the numbers of lone pairs (LP) and bond pairs (BP) of electrons: linear; trigonal planar and bent; tetrahedral and trigonal pyramidal and bent; trigonal bipyramidal and seesaw and T-shaped and linear; octahedral and square pyramidal and square planar
* Differentiate the terms bond dipole and dipole moment.
* Use electronegativity values, Lewis structures and VSEPR to determine if a molecule is a polar or nonpolar
* Determine the hybridization of a bonded atom ;
* Hybridization of atomic orbitals: in order to explain the shapes of molecules, we have to hybridize atomic orbitals; s + p = two sp hybrids, directed at 180° with respect to each other; s + p + p = three sp2 hybrids, directed at 120° with respect to each other; s + p + p + p = four sp3 hybrids, directed at 109.5° with respect to each other; s + p + p + p + d = five sp3d hybrids, directed according to trigonal bipyramidal geometry; s + p + p + p + d + d = six sp3d2 hybrids, directed according to octahedral geometry
* Evaluate a molecule or polyatomic ion to determine if the pi electrons or localized or delocalized.
* Describe the bonding using the concepts of hybridization including showing how hybridization works using orbital diagrams of ground state, excited state and mixing of orbitals from the VBT perspective.
* *Understand the role hybridization in defining the shape of a molecule and how are sigma* *and pi electrons /electron charge density distributed in this molecule.*
* Give the number of sigma (σ), pi ( π), and unshared pairs in a molecule.
* Draw an orbital picture view of a sigma or pi bond from the perspective
* Draw the molecular orbitals ( bonding and antibonding) for the following atomic orbital combinations; (1) s and s, (2) p and p head-to-head , and (3) p and p sideways.
* Compare the strengths and weaknesses of Lewis structures /resonance theory, VSEPR VBT, VBT- hybridization and MO Theory.

Valence bond theory: atoms bond covalently via orbital overlaps; sigma (σ) bonds involve end-to-end overlapping of orbitals; pi (π) bonds form when unhybridized p orbitals overlap sideways; sigma bonds have more overlap and are stronger than pi bonds

* Draw and use molecular orbital diagrams to predict bond order/relative molecule stability and magnetism (paragmagnetic or diagmagnetic) in diatomic molecules or ions. (e.g. Using molecular orbitals for second row elements explain why N2 is much more stable than O2 in terms of bond order……..)
* Recognize the correct molecular electron configuration for N2 and O2 when give an energy diagram showing the electron distribution in their molecular orbitals i.e. (σ2s 2  σ2s 2 \* σ2p 2 etc….