Given below are rough outlines of chapters 7 - 10. These are the major concepts with which you should be well-acquainted. As this is just an outline, it is not meant to be a substitute for your notes or for the text. You are responsible for all material covered in lecture and in the textbook!

### Chapter 7: Quantum Theory of the Atom

**Electromagnetic Radiation**
- energy in the form of a wave
- relationships exist between energy \( E \), wavelength \( \lambda \), frequency \( \nu \), speed \( c \), and classification according to the electromagnetic spectrum

**Planck, Einstein, and de Broglie**
- What contributions did these scientists make to our understanding of wave-particle duality?
- wave-particle duality: all matter and energy has a dual nature - particle and wave characteristics
- Planck: quantized energy, relationship between \( E \) and \( \nu \) (and \( \lambda \))
- Einstein: photoelectric effect, photons
- de Broglie: \( \lambda \) associated with object of known mass and velocity

**Atomic (Line) Spectrum of Hydrogen**
- How did the understanding and interpretation of the line spectrum of hydrogen evolve?
- contributions of Balmer, Rydberg, Bohr, and Schrodinger

### Quantum Mechanics
- What are the contributions of Schrodinger, Heisenberg, and Born?
- Schrodinger: wave equation to determine values of the wave function, \( \psi \); each \( \psi \) represents an atomic orbital
- difference between \( \psi \) and \( \psi^2 \) and what they represent
- probability of finding an electron relative to distance from nucleus; probability distributions
- Heisenberg Uncertainty Principal: an electron’s position and momentum cannot be known simultaneously

**Quantum Numbers and Atomic Orbitals**
- assignment and interpretation of quantum numbers \( n, l, m_l, m_s \)
- relationship between quantum numbers and specific atomic orbitals
- shells (defined by \( n \)), subshells (defined by \( l \)), and spatial orientation (defined by \( m_l \)) of atomic orbitals
- What do the individual atomic orbitals look like?
- relative energies of atomic orbitals

### Chapter 8: Electron Configurations and Periodicity

**Electron Configurations**
- How do electrons populate atomic orbitals? Pauli Exclusion Principle, Hund’s Rule
- writing and interpreting electron configurations; relationship to position in periodic table

**Periodic Trends**
- What are the trends in the following properties based and position in the periodic table? Can you explain why?
- atomic size (radius)
- effective nuclear charge
- ionization energy
- electron affinity
- electronegativity

### Chapter 9: Ionic and Covalent Bonding

**Bonding in Ionic Compounds**
- Lewis symbols and the octet rule related to ions and ionic compounds
- Lattice energy: chemical equation that defines lattice energy; relationship to ion size and charge
- electron configurations of ions
- trends in ion size: size of ion relative to parent atom; relative sizes of isoelectric ions

**Covalent Bonding**
- shared electron pair bonds; octet rule related to covalent compounds
- single bonds vs. multiple bonds: number of electron pairs shared; bond length; bond strength (or energy)
- bond polarity and dipole moments: based on relative electronegativity of elements; periodic trend in electronegativity
- bond dissociation energy and calculation of \( \Delta H \) of reactions
Lewis Structures
- drawing Lewis structures for compounds or ions that both follow and violate the octet rule
- resonance structures
- formal charge determination and determination of “best” Lewis structure

Chapter 10: Molecular Geometry and Chemical Bonding Theory
VSEPR Theory, Molecular Geometry, and DipoleMoments
- use to predict electron region and molecular geometries and bond angles
- use to predict preferred lone pair arrangement in trigonal bipyramidal and octahedral geometries
- effect of lone pairs or multiple bonds on bond angles in compounds
- use to predict compound polarity based on geometry (as opposed to bond polarity)

Some review problems:
The purpose of these review problems is simply to provide you with extra practice. In no way should you interpret these problems to be a representation of the exam. In other words . . . **DO NOT EXPECT THE EXAM TO LOOK LIKE THIS!!**

1. a. Calculate the frequency of infrared radiation with a wavelength of 800 nm.
   b. Calculate the wavelength of microwave radiation with frequency of $1.0 \times 10^{11}$ Hz.

2. Use the Bohr model of the hydrogen atom and calculate the wavelength corresponding to the transition of an electron from $n = 6$ to $n = 2$. What region of the electromagnetic spectrum does this correspond to? What color? (see fig 7.5)

3. Calculate the wavelength of an electron with velocity equal to $1.5 \times 10^8$ m$\cdot$s$^{-1}$, and for a neutron with the same velocity. Refer to the inside back cover of your text for masses of an electron and neutron.

4. Determine the number of . . .
   a. orbitals with $n = 4$
   b. subshells with $n = 1$
   c. orbitals in 5p subshell
   d. electrons in a full 3d subshell

5. Which of the following sets of quantum numbers are not allowed? Explain why.
   a. $n = 4$, $l = 2$, $m_l = -1$, $m_s = +1/2$
   c. $n = 4$, $l = 4$, $m_l = -1$, $m_s = -1/2$
   b. $n = 5$, $l = 0$, $m_l = -1$, $m_s = +1/2$
   d. $n = 2$, $l = 1$, $m_l = 1$, $m_s = -1/2$

6. Write the ground state electron configuration for the following elements:
   a. calcium
   b. nitrogen
   c. bromine
   d. zirconium
   e. gold
   g. uranium

7. Place the following orbital subshells in order of increasing energy: 2s, 4d, 2p, 3p, 1s.

8. Calculate the energy and wavelength of a photon whose frequency is $1.145 \times 10^{13}$ s$^{-1}$.

9. a. The principal quantum number for an atomic orbital is 4. List all possible values of $l$ for this orbital.
   b. The angular momentum quantum number ($l$) for an atomic orbital is 2. List all possible spatial orientations for the orbitals in this subshell.

10. Sketch and then compare/contrast (with respect to size, shape and orientation) the following pairs of orbitals
   a. 2s vs. 4s
   b. 3p$_x$ vs. 3p$_y$
   c. 3p$_z$ vs. 3d$_{z^2}$
   d. 3d$_x$ vs. 3d$_{x^2-y^2}$

11. The ground state electron configuration for an atom is: [Ar]4s$^2$3d$^3$.
   a. Identify this element.
   b. Is this element a metal, nonmetal or metalloid?
   c. Is this element best classified as a main group metal, a main group nonmetal, a transition metal, or an f-block element?

12. Indicate which element in each of the following pairs you would expect to have the larger atomic radius.
   a. Na or K
   b. K or Ca
   c. Si or Ga
13. Indicate which element in each of the following pairs you would expect to have the higher 1st ionization energy.
   a. Bi or Xe       b. Se or Te       c. Rb or K       d. Na or Mg       e. P or S

14. Indicate which element in each of the following pairs you would expect to have the more favorable electron affinity.
   a. Rb or Sr       b. I or Se       c. Ba or Te       d. Cl or Br

15. Predict the formulas of the ionic compounds that will form between Na and S, and Al and Cl. Use Lewis symbols of the elements to show the transfer of electrons and rationalize the formulas you predicted.

16. Write the electron configurations for each of the following:

17. Arrange the following sets of species in order of increasing radius.
   a. Rb, Rb⁺      b. Se, Se²⁻      c. Ni, Ni²⁺, Ni³⁺

18. Arrange the following in order of increasing radius and explain the order based on effective nuclear charge: P³⁻, Ar, Ca²⁺, Cl⁻, K⁺, S²⁻.

19. Based on Lewis symbols, predict the chemical formula for the simplest compound that forms between Si and Cl.

20. Arrange the following in order of increasing electronegativity.
   a. Sr, Cs, Br      b. Ca, Ge, Ga      c. P, As, S

21. Use trends in electronegativity to predict the polarities of the following bonds. Then use electronegativity values (see figure 9.15) to calculate Δχ to arrange the same bonds in order of increasing polarity. H–Se, P–Cl, N–Cl, C–C

22. Draw the Lewis structures for the following:
   a. P₂           e. SnCl₅⁻     i. TeF₆
   b. COBr₂        f. S₂⁻      j. XeF₅⁺
c. HNO₂ (N bonded to one O and one OH group)  g. XeF₂      k. BeBr₂
   d. ClO⁻         h. SeF₄

23. Write resonance structures for each of the following: FNO₂, N₃⁻.

24. Draw one Lewis structure that follows the octet rule for each of the following, and then calculate the formal charge on each atom in each structure: SO₂, SO₃, SO₄²⁻, SO₅²⁻.

25. Which of the following has the shorter carbon – oxygen bond? Lower carbon – oxygen bond dissociation energy? methanol, CH₃OH or formaldehyde, CH₂O

26. Use VSEPR theory to predict the shape of the following species:
   a. CCl₄           e. PF₅
   b. H₂Se          f. SCl₂
   c. AsF₃          g. SnCl₅⁻
d. NO₃⁻         h. ClF₂

27. Which of the following molecules are nonpolar? H₂S, PF₃, TeF₆, BeF₂, AsF₃, SCl₄

28. Which of the following are trigonal planar? SnCl₅⁻, BCl₃, GaF₅, PH₃

29. Write balanced chemical equations (complete with physical states) that correspond to each of the following:
   a. 2nd ionization energy for aluminum    c. lattice energy of calcium chloride
   b. formation of calcium chloride        d. electron affinity for oxygen

30. Use bond dissociation energies from your text (table 9.5) to calculate ΔH for the following reaction:

   2 NH₃ (g) + Cl₂ (g) → N₂H₄ (g) + 2 HCl (g).
Chem 172: Answers to Review Problems for Exam 1

1. a. \(3.75 \times 10^{14} \text{ s}^{-1}\)
   b. \(3.0 \times 10^{-3} \text{ m}\)

2. 411 nm, visible region, violet light

3. for electron: \(4.85 \times 10^{-12} \text{ m}\)
   for neutron: \(2.65 \times 10^{-15} \text{ m}\)

4. a. 16
   b. 1
   c. 3
   d. 10

5. a. allowed
   b. not allowed; if \(l=0\), must equal 0
   c. not allowed; if \(n=4 l\) can equal 0, 1, 2, or 3 (0 \(\rightarrow\) n=1)
   d. allowed

6. a. \([\text{Ar}]4s^2\)
   b. \([\text{He}]2s^22p^3\)
   c. \([\text{Ar}]4s^33d^{10}4p^5\)
   d. \([\text{Kr}]5s^24d^{10}\)
   e. \([\text{Xe}]6s^14f^{12}5d^{10}\)
   f. \([\text{Rn}]7s^25p^66d\)

7. 1s < 2s < 2p < 3p < 4d

8. \(E = 7.587 \times 10^{-21} \text{ J; } \lambda = 2.62 \times 10^{-5} \text{ m}\)

9. a. \(l = 0, 1, 2, 3\)
   b. \(m_l = -2, -1, 0, 1, 2\) or five spatial orientations; \(d_{xy}, d_{xz}, d_{yz}, d_{3z^2-r^2}, d_{z^2}\)

10. a. 2s and 4s are both spherical, but the 4s orbital is larger.
    b. 3p, and 3p, both have 2 lobes and one nodal plane but have different spatial orientations.
    c. 3p, vs. 3d: 3p, has 2 lobes and 1 nodal plane; 3d has 2 lobes + a donut of electron density around the middle and 2 nodal planes
    d. 3d and 3d both have 4 lobes and 2 nodal planes but different spatial orientations.

11. a. vanadium
    b. metal
    c. transition metal

See also Fig. 7.25 and 7.26 in text
12. a. K  
   b. K  
   c. Ga

13. a. Xe  
   b. Se  
   c. K  
   d. Mg  
   e. P (remember there is a glitch in the trend between group 5 and 6)

14. a. Rb (alkaline earth metals have unfavorable electron affinity values)  
   b. I  
   c. Te  
   d. Cl

15. Na₂S and AlCl₃

16. a. [Ne]3s²  
   b. [Ne]  
   c. [Ar]4s¹3d⁶4p⁶  
   d. [Ar]4s¹3d⁶4p⁶  
   e. [Ar]4s¹3d⁶  
   f. [Ar]3d⁶  
   g. [Ar]3d⁶

17. a. Rb⁺ < Rb  
   b. Se < Se⁻  
   c. Ni²⁺ < Ni²⁺ < Ni

18. Ca²⁺ < K⁺ < Ar < Cl⁻ < S²⁻ < P⁻

19. SiCl₄

20. a. Cs < Sr < Br  
   b. Ca < Ga < Ge  
   c. As < P < S

21. C–C and N–Cl are both nonpolar (Δχ = 0) < H–Se (Δχ = 0.3) < P–Cl (Δχ = 0.9)

22. (a) P₂ : P ≡ P  
   (b) COBr₂ : Br–C–Br  
   (c) HNO₂  
   (d) ClO⁻  
   (e) SnCl₃ 
   (f) S₂⁻  
   (g) XeF₂ : F–Xe–F  
   (h) SeF₄ : F–Se–F
(i) TeF₆

(j) XeF₅⁺

(k) BeBr₂

23) FNO₂

\[ \begin{array}{c}
\text{F} \\
\text{N} \\
\text{O} \\
\text{O}
\end{array} \]

\[ \rightarrow \]

\[ \begin{array}{c}
\text{F} \\
\text{N} \\
\text{O} \\
\text{O}
\end{array} \]

\[ \text{N₃}^− \]

\[ \begin{array}{c}
\text{N} \\
\text{N} \\
\text{N}
\end{array} \]

\[ \rightarrow \]

\[ \begin{array}{c}
\text{N} \\
\text{N} \\
\text{N}
\end{array} \]

\[ \rightarrow \]

\[ \begin{array}{c}
\text{N} \\
\text{N} \\
\text{N}
\end{array} \]

24. SO₂

\[ \begin{array}{c}
\text{O} \\
\text{S} \\
\text{O}
\end{array} \]

\[ \text{S} + 1 \]

\[ \text{O} \]

\[ -1 \]

\[ \text{S} \]

\[ +2 \]

\[ \text{O} \]

\[ -1 \]

\[ \text{S} \]

\[ -1 \]

\[ \text{O} \]

\[ -1 \]

*Note: There are other structures you can draw that are also acceptable, especially since S can have greater than an octet.*

SO₃²⁻

\[ \begin{array}{c}
\text{O} \\
\text{S} \\
\text{O}
\end{array} \]

\[ \begin{array}{c}
\text{O} \\
\text{S} \\
\text{O}
\end{array} \]

\[ \begin{array}{c}
\text{O} \\
\text{S} \\
\text{O}
\end{array} \]

SO₄²⁻

\[ \begin{array}{c}
\text{O} \\
\text{S} \\
\text{O}
\end{array} \]

\[ \begin{array}{c}
\text{O} \\
\text{S} \\
\text{O}
\end{array} \]

\[ \begin{array}{c}
\text{O} \\
\text{S} \\
\text{O}
\end{array} \]

\[ \begin{array}{c}
\text{O} \\
\text{S} \\
\text{O}
\end{array} \]

\[ \text{S} + 2 \]

\[ \text{each } \text{O} = 1 \]
25. CH₃O has the shorter C-O bond; CH₃OH has the lower C-O bond dissociation energy

\[ \text{H}_3\text{C} - \text{O} - \text{H} \quad \text{vs.} \quad \text{H} - \text{C} - \text{H} \]

26. a. tetrahedral
   b. bent or v-shaped
   c. trigonal pyramidal
   d. bent or v-shaped
   e. trigonal bipyramidal
   f. see-saw
   g. trigonal bipyramidal
   h. linear

27. TeF₅ and BeF₂ are nonpolar

28. BCl₃ and GaF₃ are trigonal planar

29. a. Al⁺ (g) → Al²⁺ (g) + e⁻
   b. Ca (s) + Cl₂ (g) → CaCl₂ (s)
   c. CaCl₂ (s) → Ca²⁺ (g) + 2 Cl⁻ (g)
   d. O (g) + e⁻ → O⁻ (g)

30. –28.1 kJ/mol