AP Chemistry

Chapter 6 Outline

a) The Wave Nature of Light
   i) Electromagnetic radiation = radiant energy
      (1) Many forms, including visible light
      (2) Speed of light in a vacuum is a constant $c = 3.00 \times 10^8$ m/s (on green sheet)
      (3) wavelength =
         (a) distance between two adjacent points
         (b) λ
      (4) frequency
         (a) the number of wave fronts that pass a given point in a second ν
         (b) common unit Hertz, (s^{-1})
   ii) $c = \lambda\nu$ CRITICAL EQUATION; be able to use this in calculations

b) Quantized Energy and Photons
   i) Energy is quantized: energy can only be released in specific amounts
      (1) Quantum or photon
   ii) $E = h\nu$ CRITICAL EQUATION; be able to use this in calculations
      (equation and value for h is on green sheet)
   iii) Radiant energy is quantized   photons vs. wave behavior
      (1) Light possesses both wave-like and particle-like behavior

c) Line Spectra and the Bohr Model
   i) Monochromatic light = light of a single wavelength
   ii) Spectrum = when radiation from a source is separated into its different wavelengths
      (1) Continuous spectrum = rainbow of colors, containing light of all wavelengths
      (2) Some radiation sources give off light with only a few, specific wavelengths—line spectra
         (a) Emission spectra
         (b) Absorption spectra
      (3) Rydberg equation: $\frac{1}{\lambda} = R_h\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$ where $n_2 > n_1$
      (4) Not essential, but might be useful…not on the formula sheet
   iii) Bohr's Model…greatly emphasized in redesign of AP Chem
      (1) Only orbits of certain radii are allowed, corresponding to certain definite energies.
      (2) An electron in a permitted orbit has an "allowed" energy and will not spiral into the nucleus.
      (3) Energy is emitted or absorbed by the electron only as the electron changes from one allowed state to another. This energy is emitted or absorbed as a photon.
      (4) Allowed energy levels vary with n
         (a) $E = -2.18 \times 10^{-18}\ J\left(\frac{1}{n^2}\right)$ very useful formula; not on green sheet anymore!

(i) Note relationship between E and n
(b) Ground state = lowest energy state of atom
(c) Excited state = when the atom is in a higher energy orbit
(d) The radius increases as n^2
(e) $\Delta E = E_f - E_i = E_{photon} = h\nu$

d) The Wave Behavior of Matter
i) Matter waves --De Broglie
(1) Any object with mass and velocity acts as a wave. However, for ordinary matter, the wavelength is so tiny as to be unobservable. For electrons, the wave properties are very significant!
(2) $\lambda = \frac{h}{mv}$ Useful formula; no longer on green sheet
ii) Heisenberg Uncertainty Principle = It is inherently impossible to simultaneously know both the momentum and location of an electron with any precision.
iii) Modern model: the electron is a particle, whose behavior is described in terms appropriate to waves. We can precisely describe the energy of the electron while discussing its location in terms of probabilities.

e) Quantum Mechanics and Atomic Orbitals
i) Wave Mechanics = Quantum Mechanics = incorporates both wave and particle behaviors of electrons
(1) Erwin Schrödinger's equation
(2) Solutions lead to mathematical functions, called wave functions Ψ (psi)
(a) Ψ^2 = probability density = electron density = probability of finding the electron in a certain region of space at a given instant
(b) Often represented as "electron clouds"= orbitals = specific distribution of electron density
ii) Only certain orbitals, with certain energies, are allowed
(1) 3 quantum numbers to describe an orbital
(a) Principal quantum number n
(b) n = 1, 2, 3, …
(i) as n increases, energy increases & electron is less tightly bound to nucleus
(ii) as n increases, the orbital becomes larger
(iii) $E_n = -2.18 \times 10^{-18}\ J \left(\frac{1}{n^2}\right)$ just as in Bohr model
(2) Azimuthal or angular momentum quantum number ℓ
(a) $m_\ell = 0, \ldots, n-1$
(b) ℓ determines shape of orbital

Value of ℓ	0	1	2	3
Letter used	s	p	d	f

(1) magnetic quantum number m_ℓ (or sometimes just m)
(a) $m_\ell = -\ell, \ldots, 0, \ldots, +\ell$
(b) m_ℓ determines orientation of orbital in space (aka, how many orbitals in a sublevel)

ii) electron shell = collection of orbitals with the same value of n
   (1) The total number of orbitals in a shell is n^2
   (2) subshell = the set of orbitals that have the same n and ℓ values
   (3) The shell with principal quantum number n will consist of n subshells
   (4) Each subshell has specific number of orbitals
      (a) s orbitals are singlets;
      (b) p orbitals come in sets of 3;
      (c) d orbitals come in sets of 5;
      (d) f orbitals come in sets of 7
iii) ground state = when electrons are all in their lowest energy orbital
   (1) excited state = when an electron is occupying a higher energy orbital than normal

a) Representations of orbitals
   i) S orbitals
      (1) Spherically symmetric; often represented as spherical boundary surface
      (2) Radial probability functions: maximum of function gives most probable distance from nucleus
      (3) Node = intermediate point at which probability of finding an electron is zero
   ii) P orbitals first appear in 2nd shell
      (1) Dumbbell shaped orbitals with two lobes
      (2) The three orbitals in the set are at 90^o angles to each other
   iii) D orbitals first appear in 3rd shell
      (1) 4 have "cloverleaf" shape; 5th has two lobes, with a donut (torus)
   iv) F orbitals first appear in 4th shell
      (1) 8 lobes! Not even shown in our text!

f) Many-Electron Atoms
   i) In a many-electron atom, for a given value of n, the energy of an orbital increases with increasing value of ℓ.
      (1) The precise energies of the orbitals depends on the atom
      (2) All orbitals of the same subshell are "degenerate" –they have the same energy as one another
   ii) Electrons have the property of "spin," which is quantized.
      (1) Spin magnetic quantum number m_s (or sometimes just s) = +½, -½
   iii) Pauli Exclusion Principle
      (1) No two electrons in an atom can have the same set of four quantum numbers.
      (2) An orbital can hold a maximum of two electrons, and they must have opposite spins.

g) Electron Configuration = the way in which electrons are distributed among the various orbitals of an atom
   i) Aufbau principle
      (1) Orbitals are filled in order of increasing energy.
      (2) Be able to use the periodic table to predict filling sequence
   ii) Hund's Rule

(1) For degenerate orbitals, the lowest energy is obtained when the number of electrons with the same spin is maximized.
(2) For degenerate orbitals, place 1 electron in each orbital, all with parallel spin, before pairing any electrons. This minimizes electron-electron repulsions.

iii) Orbital notation: arrows and boxes

iv) Noble gas notation: use symbol of nearest noble-gas of lower atomic number in brackets to represent "noble gas core"
(1) Core electrons = inner shell electrons
(2) Valence electrons = outer shell electrons, involved in chemical bonding

v) Transition elements
(1) d block elements = elements in groups 3-12
(1) Often exhibit exceptional electron configurations
(2) It is more stable to have fully filled or half-filled subshells of degenerate orbitals! Cu, Cr, etc.
(a) Spherical, symmetrical electron clouds minimize repulsions

ii) Lanthanides and Actinides
(1) Fill the 4f and 5f orbitals, respectively
(2) Sometimes exhibit exceptional electron configurations, involving the d electrons

h) Electron Configurations & The Periodic Table
i) Know the locations of the s, p, d and f blocks
ii) The periodic table is your best guide to predicting the filling sequence!
iii) Representative elements
(1) main block elements
(2) elements in s and p blocks
iv) Valence Electrons
(1) For main block elements, group number or (group number -10) gives number of valence electrons
(a) For main block elements, we do not consider completely full d or f subshells to be among the valence electrons.
(b) For transition elements, we do not consider completely full f subshells to be among the valence electrons.