

Electromagnetic Radiation

- Know what the wavelength, frequency, and energy are, and how they relate to each other.
- Know about the electromagnetic spectrum.

Line Spectra

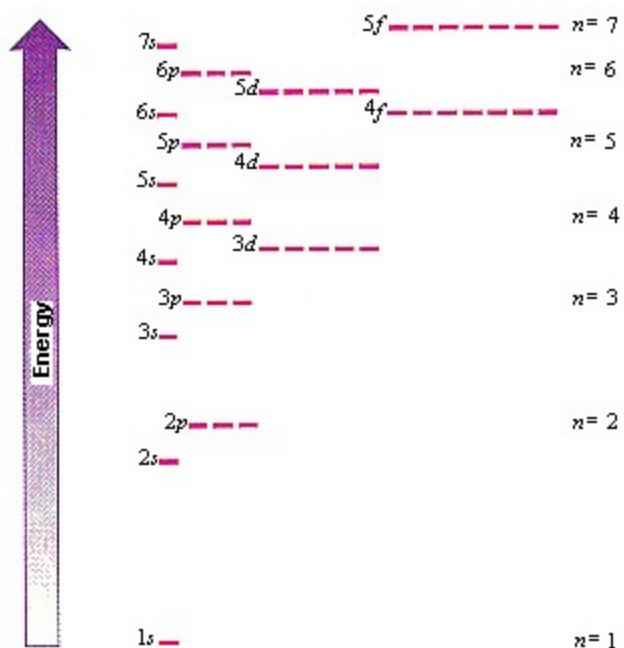
- Why do emission lines occur? What about the reverse spectra (absorption spectra), where you have a continuous spectrum with some lines of black?

Atomic spectra and energy levels - Bohr's Atom

- Energy levels in atoms. Electrons transfer from one energy level to the other.
- When they move up to a higher level, energy is absorbed. (absorption of light)
- When electrons fall back to lower levels, energy is released. (emission of light)
- When the energy levels are spread apart more, the energy released or absorbed is greater, and the frequency is greater (and the wavelength is shorter).

Sublevels and orbitals - Quantum mechanics

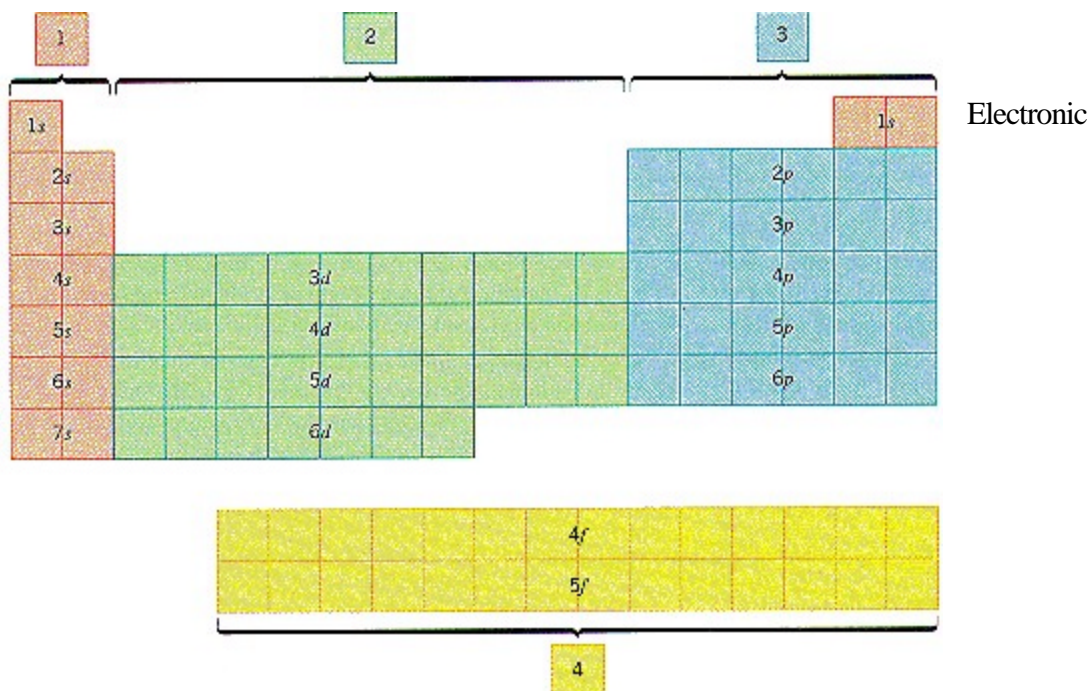
- Electrons exist in orbitals, two electrons per orbital. The shape of an orbital describes the probability that an electron will be found at a certain location. Two electrons can fit in each orbital, and they must have opposite spin.
- The principal energy levels ($n=1,2,3,\dots$ from the Bohr atom) are split into sublevels (s, p, d, f). Each sublevel has a certain number of orbitals. Know the shapes of orbitals in the s, p, and d sublevels. Know the number of orbitals in each of the four types of sublevels, and the number of electrons that can go in each sublevel. Know how you can remember this stuff by looking at the periodic table.



The diagram to the left shows the sequence and relative energy of sublevels for each and every element. The number indicates the principal energy level; the letter indicates the sublevel; and the dashes represent each orbital. Notice that the "s" sublevels have only one orbital, so each "s" sublevel can hold two electrons. The "p" sublevels have three orbitals each, so they can hold 6 electrons. The "d" orbitals have five orbitals and can hold 10 electrons. The "f" orbitals have seven orbitals and can hold 14 electrons each.

Remember that all atoms have the same energy levels available to be filled. The ground state of an atom is the situation when the electrons fill the lowest available orbitals and sublevels.

As you look across the periodic table, for instance, from hydrogen to helium, these elements have one and two electrons, respectively. Both electrons can fit into the first energy level (1). The elements lithium to neon have a full 1st energy level, and some electrons in the second energy level. First, the 2s sublevel gets used, so Li and Be would have their highest energy electron(s) in the 2s sublevel. After the 2s sublevel is full, the next six electrons go into the 2p sublevel. For example, carbon would have its two highest energy electrons in the 2p sublevel (it also has two electrons in the 2s sublevel and two in the 1s sublevel). Note: if electrons have a choice of pairing up in a "p" orbital or spreading out to other "p" orbitals in the same sublevel, which is the case for carbon, then they always spread out first; then, if all three "p" orbitals are occupied by an electron, they start pairing up (e.g., for the elements oxygen, fluorine, and neon). The electronic configuration for carbon is written like this: $1s^2 2s^2 2p^2$.



configurations - Know how to write these for every atom. Use the periodic table as guidance. The sections marked on the table on this page indicate the sublevel of the last (highest energy) electron for atoms of each element. In the example above, carbon's highest energy electrons are "2p" electrons. You can see that C falls in the blue, 2p, section. Notice that the alkali metals and alkaline earth metals have their last one or two electrons in an "s" sublevel. These are the valence electrons, as all other electrons in these elements fall into the next lower principal energy level.

The best way to remember this is to remember the locations for each sublevel (letter) and the number of the principal energy level that starts each section (1s, 2p, 3d, 4f).