Read Sections 3.6 and 3.7 before viewing the slide show.
Unit 12
Electrons and their Arrangement

• Line Spectra and Their Implications (3.6)
• Determining the Maximum Number of Electrons (3.6)
• Quantum Mechanical Interpretation of Electrons (3.7)
Exploring the Electron(3.6)

- Armed with a nuclear picture of the atom, it is now time to explore the electrons.
- In the nuclear view, the nucleus is incredibly tiny and the bulk of the space taken up by the atom is due to the electrons.
- As atoms approach each other, the electrons will be the initial contact for any further reactions.
- Thus, it is important to understand as much as possible about the electron.
- Much of our understanding comes from looking at the interaction of matter with energy.
• For our purposes here, it is best to think of light as having wavelike properties.
• Isaac Newton first observed that white light could be broken into component colors (commonly referred to as a spectrum: ROY G BIV – red, orange, yellow, green, blue, indigo, and violet, just like a rainbow)

![Image from http://files.chem.vt.edu](http://files.chem.vt.edu)
Light (3.6)

• Different colors correspond to light that has a different wavelength and, thus, a different energy.
• Observation of the interaction of light with matter can lead to a significant amount of information about the matter.
Some of our early information about electrons came from line spectra (spectra is plural, spectrum is singular).

If a gas is placed in a glass tube and a high voltage applied, the gas will emit light containing various colors (consider neon lights).

If the light coming from the tube is separated into colors of the spectrum, it is possible to learn something about the energy of the processes causing the light emission.
Line Spectra (3.6)

• The figure contains line spectra for a variety of elements.

Two important things to notice:

• Only some wavelengths (energies) are observed
• The spectrum for each of the elements shown is unique – it could be used to identify the element.
The gas discharge tubes contain a huge number of atoms yet only a few energies are observed based on the few number of spectral lines on the previous page.

This suggests that only certain energy changes can occur when the gas tube is energized – energy is transferred in only discrete packets.

Niels Bohr (1913) provided an explanation for the observed spectra.
Bohr’s Interpretation of Line Spectra (3.6)

• Electrons cannot have just any amount of energy, but can only have certain amounts. This is called the quantitization of energy. (Similar to an ATM machine, you can only extract money in multiples given by the machine – the money is quantized.)

• Though the picture on the right is known to not be totally accurate, it is a reasonable visualization tool.

• The nucleus you are already an expert at.

• The electrons are located in discrete energy shells outside of the nucleus. They really aren’t like planets, but it is an image that may be helpful early on.
Bohr’s Interpretation of Line Spectra (3.6)

• The n number in the diagram to the right indicates the shell number starting with n=1 close to the nucleus.
• When an electron falls from shell n=2 to n=1, it will emit the energy indicated between the levels.
• The n=3 to n=2 transition is less, and so on.
• Since color is related to energy, this approach describes the appearance of the few lines seen in the line spectra.

Image from http://wiki.chemeddl.org/
Bohr’s model worked well for hydrogen, but not other elements. The vision of the electron has been changed vastly since the 1920’s with the advent of quantum mechanics. Without going into great detail, the electron has both wavelike and particle-like properties. Its energy is still quantized, but we can only talk about probabilities of finding it in regions around the nucleus – no more planetary system appearance.
• Rather than talking about orbits (as in planets) we talk about orbitals – regions in space in which there is a high probability of finding the electrons.

• There are four primary orbital types discussed:
  • s orbital – spherical in shape
  • p orbital – dumbbell shaped
  • d orbital – dumbbell shaped with some variation
  • f orbital – dumbbell shaped with some variation
Orbital Shapes (3.7)

- s orbitals
- p orbitals
- d orbitals
- f orbitals

Image from www.chemistryexplained.com
Image from www.sites.google.com

Image from www.chemistryexplained.com
Important Points about Orbitals (3.7)

- Not all shells have all orbitals. The presence of orbitals in numbered shells occurs as follows:
  - s orbital: Start in the n=1 shell and continue as n increases
  - p orbital: Start in the n=2 shell and continue as n increases
  - d orbital: Start in the n=3 shell and continue as n increases
  - f orbital: Start in the n=4 shell and continue as n increases
- Each shell can have at most 1 s orbital, 3 p orbitals, 5 d orbitals, and 7 f orbitals
- The set of all of the same orbitals combined together in a shell is called a subshell
- Since each orbital can have at most two electrons, the maximum number of electrons in an s subshell is 2, in a p is 6, in a d is 10, and in an f is 14.
Writing Electron Configurations (3.7)

• The diagram at the right may be used to determine how electrons are filled into the shells and subshells of an atom.

• Recall that each s subshell can hold at most 2 electrons, each p 6 electrons, each d 10 electrons, and each f 14 electrons.

• To determine the distribution of electrons, start filling at the upper arrow and follow from one arrow to the next, filling each subshell with its maximum number of electrons until you run out.

• Example on the next page
• Suppose we wanted to write the electron configuration for Be with 4 electrons
  
  • The 1s takes 2 electrons and the 2s takes 2 more electrons. A method of indicating this distribution is called writing the electron configuration. The general format is: 
    \[ nx^m \]
    where \( n \) is the shell number, \( x \) is the subshell letter, and \( m \) is the number of electrons in the subshell.
  
  • For our Be example, the electron configuration would be \( 1s^22s^2 \).

• More examples on the next page
• Electron configuration for phosphorous with 15 electrons.
  • $1s^22s^22p^63s^23p^3$

• Electron configuration for oxygen with 8 electrons
  • $1s^22s^22p^4$

• Electron configuration for sodium with 11 electrons
  • $1s^22s^22p^63s^1$