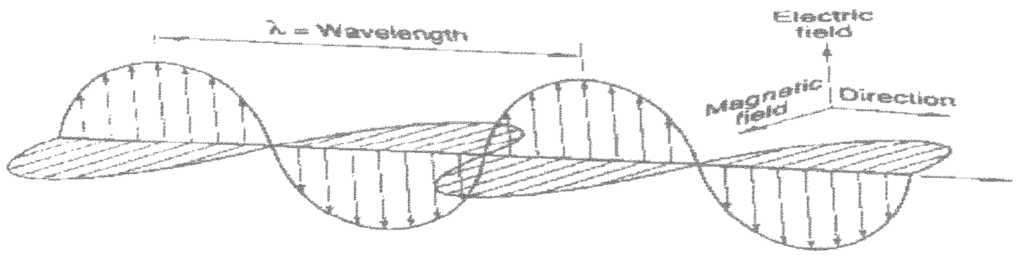
Chapter 4 – Electron Configuration

Introduction

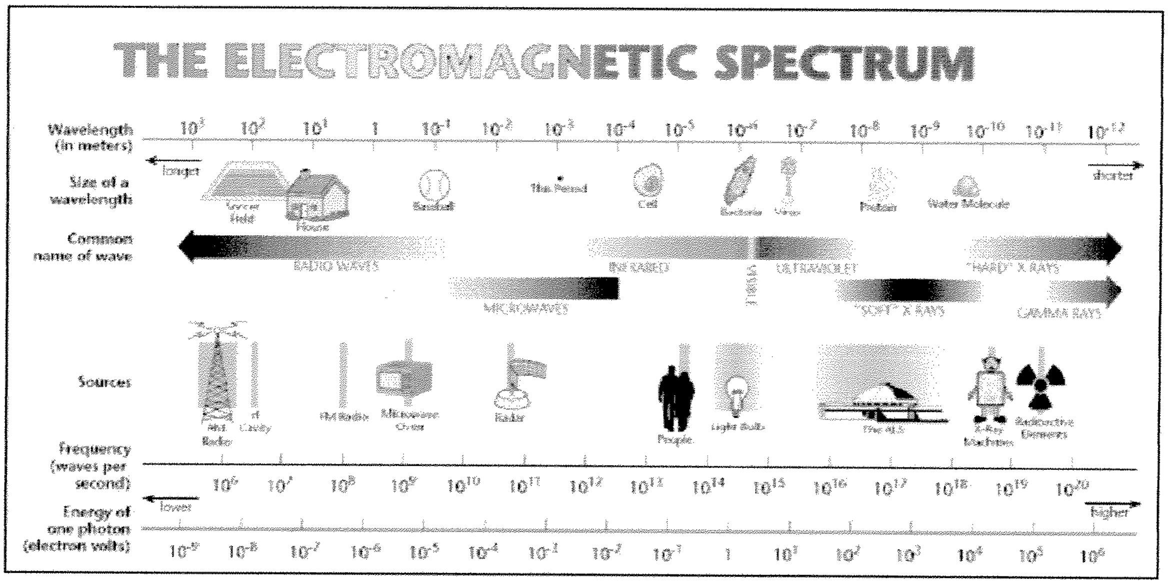
In this chapter we're going to pick up where we left off with Bohr's model of the atom, examining the energy levels of electrons, and how the electromagnetic spectrum helped to place electrons in distinct energy levels. Thus, leading to the quantum-mechanical model that describes electrons in terms of both waves and quantized energy.

Electromagnetic radiation

Electromagnetic radiation or radiant energy is energy traveling in waves. This energy consists of waves made up of oscillating electric and magnetic fields at right angles together.



Types of electromagnetic energy include light, x-rays, gamma rays, TV and radio waves, and microwaves. (You can also refer to figure 4-7 in your text)



To better understand electromagnetic radiation you must understand waves. There are four basic characteristics of waves, amplitude, wavelength (λ, lambda), frequency (ν), and speed. **Amplitude** is the height of a wave from its origin to its peak. **Wavelength** is the distance between waves, or the distance from peak to peak. **Frequency** is the number of cycles per second, s^{-1} or Hertz (Hz). Light energy travels at a constant speed (distance/time). **Speed of light = 3.00 X 10^8 m/s (constant, *c*)**

The relationship between wavelength and frequency can be expressed: **$\lambda = c/v$** (the shorter the wavelength, the greater the frequency).

- For example: Determine the wavelength of a wave if the frequency is 102 s^{-1}.

  $\lambda = c/v$
  $\lambda = (3.00 \times 10^8 \text{ m/s})/102 \text{ s}^{-1}$
  $\lambda = 2.94 \times 10^6 \text{m}$

- Determine the frequency of a wave if the wavelength is 4.5×10^{-9}m.

  $\lambda = c/v$
  $4.5 \times 10^{-9} \text{ m} = (3.00 \times 10^8 \text{ m/s})/\nu$
  $\nu = (3.00 \times 10^8 \text{ m/s}) / (4.5 \times 10^{-9} \text{ m})$
  $\nu = 6.7 \times 10^{16} \text{ s}^{-1}$

Practice: Determine the missing piece of information about the following waves.

1. What is the frequency of a wave if the wavelength is 12m?

2. What is the wavelength of a wave if the frequency is 1.2s^{-1}?

3. What is the frequency of a wave that has a wavelength of 3.22m?

4. What is the wavelength of a wave that has a frequency of $4.3 \times 10^{14}\text{s}^{-1}$?

Quantum Theory

The quantum theory or **quantum-mechanical model** is the atomic model that explains an atoms properties by considering electrons both waves and packets of energy. This model is a culmination of theories of many different scientists. We will focus on six: Planck, Einstein, Compton, DeBroglie, Bohr and Heisenberg.

Max Planck suggested that energy is emitted and absorbed in discrete amounts or quanta – quantization of energy. Planck related this "quanta" of energy to the frequency of the radiation, $E=hv$. h is a constant referred to as Planck's constant.

Albert Einstein proposed that light consists of quanta of energy which behave like tiny *particles* of light – **photons.** This theory is referred to as the **photoelectric effect**. The energy of the photon depends on the frequency of the light. Einstein proposed that the amount of energy carried in each photon could be found by using Planck's equation. Based on Einstein's work we now know that violet light carries more energy than red light because violet light has a greater frequency.

Arthur Compton proved Einstein's theory that light waves are made up of photons. Compton was able to demonstrate that a photon could collide with an electron. If light was simply a wave there could be no collision, but by showing that a photon could collide with an electron (much like two marbles hitting) it was confirmed that a photon was a particle. Photons are particles that travel at the speed of light and also have wavelength and frequency.

The electrons of an atom will either absorb the photon or not, the energy carried by the photon of light must match the specific requirement of the electron in order to excite the electron. The photon of light must contain the exact *quanta* of energy. When an atom gives off energy, it emits that energy as a photon of light. The spectrum of emitted light is unique for each element, and can be viewed as a "fingerprint" for the element. Line spectra revealed that electrons exist in discrete energy levels for the atom. **Bohr** used line spectra of the elements to explain quantized energy levels. Bohr labeled each energy level by its quantum number, n.

DeBroglie was able to put to two theories of light being waves and light being particles together. (Dual nature) DeBroglie also stated all objects have wavelike behavior, but for most objects their mass is too great for the wave to be observable. Electrons, however, has such a small mass that their wavelike behavior can be observed.

Heisenberg believed that because of the dual nature of matter and energy it would be virtually impossible to simultaneously measure the momentum and the position of an electron. This theory is called Heisenburg's Uncertainty Principle.

Name: __________________________ Period: ___ Date: _____________________

Homework: Electromagnetic Radiation and Quantum Theory

Complete the following wavelength problems.

1. What is the frequency of blue light if its wavelength is 595nm?

2. What is the wavelength of a wave if the frequency is 5.56×10^{13}hz?

3. What is the frequency of a wave if the wavelength is 2.13m?

4. What is the wavelength of if the frequency is 3.87×10^{12} Hz?

Answer the following short answer questions.

5. What characteristic of a light wave determines whether or not the light will cause a positive photoelectric effect?

6. What is the speed of light?

7. The number of complete waves passing a fixed point in a given time is called?

8. The wavelength of microwave radiation is (shorter/longer) than the wavelength of visible light.

9. What type of radiation is produced by a heat lamp?

10. A wave with a high frequency has a (short/long) wavelength.

Match the following scientists with their contributions. (a scientist may be used more than once)

a. Planck b. Einstein c. Compton d. DeBroglie e. Bohr f. Heinsenberg

____ 11. Used the atoms line spectrum to explain quantized energy levels.

____ 12. Wrote the equation $E=hv$.

____ 13. Stated that energy is emitted or absorbed in discrete pieces called quanta.

____ 14. Proposed the photoelectric effect.

____ 15. Demonstrated that an photon could collide with an electron.

____ 16. Stated all objects have wavelike behavior, but for most objects their mass is too great for the wave to be observable.

____ 17. Proposed that light consists of quanta of energy.

____ 18. Proposed that it is virtually impossible to simultaneously measure the momentum and the position of an electron.

____ 19. Proved that light consists of tiny particles, or photons.

____ 20. Labeled each energy level by its quantum number, *n*.

Choose the best answer for the following multiple choice.

____ 21. DeBroglie derived a mathematical relationship between the mass and velocity of a moving particle and the ________________ that it would exhibit.

a. motion
b. amplitude
c. wavelength
d. charge

____ 22. When radiation is absorbed by a hydrogen electron, the hydrogen atom changes its ground state to ____________.

a. an excited state
b. a lower state
c. another atom
d. the nucleus of the atom

QUANTUM NUMBERS AND ELECTRON CONFIGURATION

The **quantum-mechanical model** brings together all the ideas that had been expressed about electrons. The model states that electrons are treated as waves the have quantized energy. (Electrons behave like waves and exist in quantized energy levels.) Quantum mechanics is the study of the laws of motion that govern behavior of small particles. In this unit we will discuss the behavior of electrons, and where they are located. Electrons can be found in energy levels that are composed of orbitals, regions in space around the nucleus. These orbitals have various shapes and are labeled with the letters *s, p, d,* and *f*. Each energy level consists of a different combination of these orbitals. Electron configurations (locations) are determined by the distribution of the atom's electrons within the principal energy levels.

As Heisenberg stated in his Uncertainty Principle it is impossible to know the exact location and velocity of an electron at any given point, however, through the use of **quantum numbers** we try to give an approximate location of an electron. Through quantum numbers we can tell what energy level, sublevel, and even the spin the electron has. There are four quantum numbers that are used together to give the probability of where an electron can be found, they are the "address" of the electron.

The first quantum number is labeled ***n*** and is the **principle energy level**. The second quantum number is called **azimuthal**, and labeled ***l***. Azimuthal refers to the sublevel that the electron is in. The third quantum number is **magnetic**, labeled ***m***. Magnetic refers to which orbital (within the sublevel) the electron is in. Each orbital can hold 2 electrons. The *s* sublevel has one orbital, *p* has three, *d* has five and *f* has 7. The final quantum number tells the **spin** of the electron, labeled *s*. The **Pauli Principle** states that no two electrons may have the same spin in the same orbital. The spin is often represented by up and down arrows or +1/2 and −1/2.

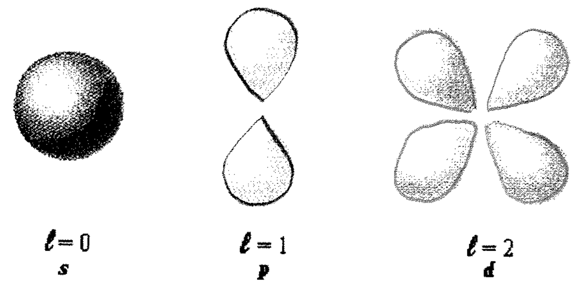
Principal Energy Levels – quantum number, ***n***

Sublevels: ***n***=1, one sublevel 1*s*
n=2, two sublevels 2*s*, 2*p*
n=3, three sublevels 3*s*, 3*p*, 3*d* and so on.

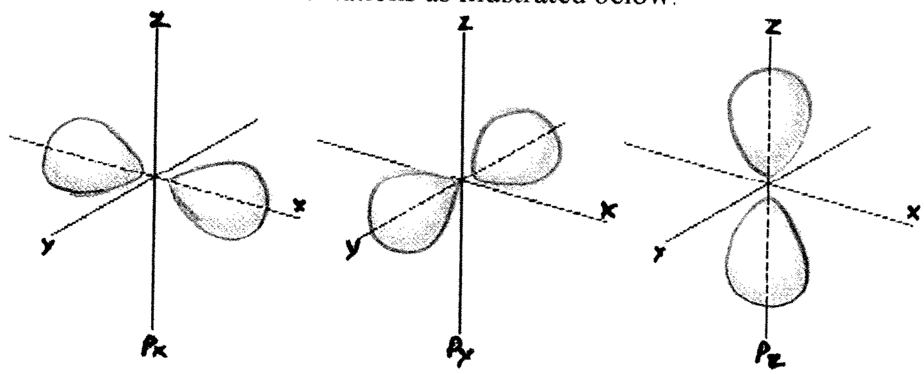
Practice: Complete the following table.

Description	**Quantum #**	**Label**	**Name**
The number of orbitals	____________	________	________________
The energy level	____________	________	________________
The spin of the electrons	____________	________	________________
The sublevel of the electron	____________	________	________________

s is spherical shape *p* is dumbbell shaped *d* is two sets of **dumbbells**



The *p* orbitals can have three orientations as illustrated below:



* as electrons are distributed throughout the levels overlap, we will discuss this in detail.

Summary of Allowed Combinations of Quantum Numbers

n	*l*	*m*	Subshell Notation	Number of Orbitals in the Subshell	Number of Electrons Needed to Fill Subshell	Total Number of Electrons in Subshell
1	0	0	1s	1	2	2
2	0	0	2s	1	2	
2	1	1,0,-1	2p	3	6	8
3	0	0	3s	1	2	
3	1	1,0,-1	3p	3	6	
3	2	2,1,0,-1,-2	3d	5	10	18
4	0	0	4s	1	2	
4	1	1,0,-1	4p	3	6	
4	2	2,1,0,-1,-2	4d	5	10	
4	3	3,2,1,0,-1,-2,-3	4f	7	14	32

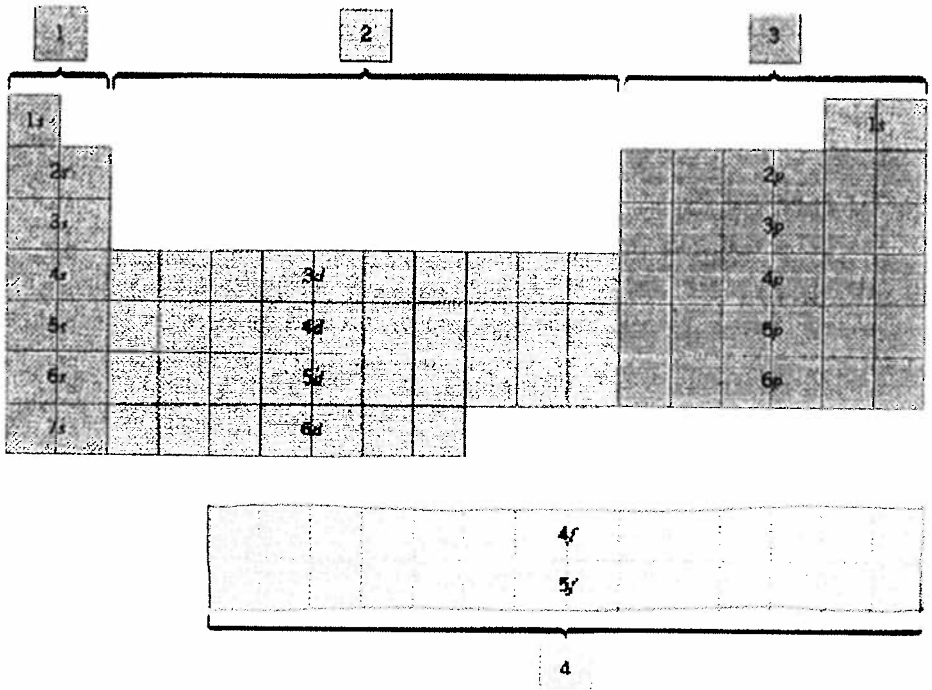
Practice: Answer the following questions based on quantum numbers.

1. If n=2, what is the maximum # of electrons that can be present? ________

2. If n=4, what is the total # of orbitals present? ________

3. If n=1, what is the name of the sublevel present? ________

4. If n=3, what sublevels are present? ________



- For example: Write the electron configuration for Titanium.
  Ti = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$
  * A short hand method maybe used by refering to the most recent noble gas and then completing the electron configuration: [Ar] $4s^2\ 3d^2$

- Write the electron configuration for Chlorine.
  Cl = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
  The short hand method would be: [Ne] $3s^2\ 3p^5$

Practice: Write both the full electron configuration for the following and the short hand electron configuration.

1. Zinc

2. Bromine

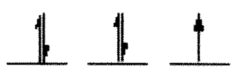
3. Scandium

4. Argon

When this "address" is written the information is written as $1s^2$, this would indicate that the electron spoken of is in the first energy level and the second electron in the *s* orbital. If the spin is desired, boxes or lines will be used to represent the orbitals.

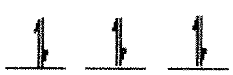
The orbitals drawn would represent the *p* orbital for Oxygen ($1s^2\ 2s^2\ 2p^4$). Each orbital will receive one electron until they each have one and then a second will be placed in each orbital.

- For example: What would the quantum numbers be for Fluorine?
  Fluorine has 9 electrons to place; the first two would be placed in the 1*s* orbital, the next two in the 2*s* orbital, and the last five would be placed in the *p* orbitals. It would look like this:
  F = $1s^2\ 2s^2\ 2p^5$, the *p* orbital would be drawn as:



- Write the quantum numbers for Neon.

Neon has 10 electrons and they would be written as: Ne = $1s^2\ 2s^2\ 2p^6$
The *p* orbital would be drawn as:



Practice: Write the quantum numbers and draw the outer orbitals for the following elements.

1. Lithium ____________________
2. Magnesium ________________
3. Phosphorus ____________________
4. Argon _________________________

The within a given energy level the *s* orbital always has the lowest energy, therefore, the orbitals of the different energy levels begin to overlap. As a result, the size and shape of the orbital control the energy of an orbital for most atoms. The resultant filling pattern is illustrated diagrams illustrated below.

Name: ______________________________ Period: ____ Date: ____________________

Homework: Quantum Numbers and Electron Configuration

Answer the following questions based on your knowledge of quantum numbers.

1. What is the name of the set of numbers given to electrons in an atom, (the electrons temporary address).

2. Describe an orbital pair.

3. What does Heisenberg's Principle state?

4. Which quantum level signifies the number of sublevels?

5. What is the pathway of an electron called?

6. What is the maximum number of electrons that the third energy level can have?

7. What does the magnetic quantum number refer to?

8. What are the labels for the four principle quantum numbers?

9. What does the Pauli Principle state?

10. Which quantum number refers to the size of an electron cloud?

Answer the following questions pertaining to quantum numbers.

11. If n=1, what is the total number of orbitals present? _________

12. If n=4, what is the maximum number of electrons that can be present? _________

13. If n=3, what is the maximum number of electrons that can be present? _________

14. If n=2, what is the total number of orbitals present? _________

15. If n-=2, what are the names of the sublevels that exist? _________

Write the full electron configuration, short hand configuration and orbital notation for the following.

16. sodium

17. Chromium

18. Krypton

19. Barium

20. Silver