Honors Chemistry Chapter 4 Arrangement of Elements in Atoms

Section 1: Development of a New Atomic Model pgs. 96-103.

Objectives:

1. **Explain the mathematical relationship among the speed, wavelength, and frequency of electromagnetic radiation.**
2. **Discuss the dual wave-particle nature of light.**
3. **Discuss the significance of the photoelectric effect and the line-emission spectrum of hydrogen to the development of the atomic model.**
4. **Describe the Bohr model of the hydrogen atom.**

Vocabulary: Define the following.

1. **electromagnetic radiation--**
2. **electromagnetic spectrum--**
3. **wavelength--**
4. **frequency--**
5. **photoelectric effect--**
6. **quantum--**
7. **photon--**
8. **ground state--**
9. **excited state--**
10. **line-emission spectrum**

11. continuous spectrum--

Rutherford's solar system model of the atom did not explain where electrons are located around a nucleus and what prevented them from being sucked into the positive nucleus. New theories were needed to explain electron position. It was the study of the properties of light and the close relationship of light and the electron that led to modern theories of electron position around the nucleus of the atom.

<u>Properties of Light</u>

Light is closely related to electrons. Light has always been thought of a behaving as a wave, but with the discovery of light's relationship with electrons, which are particles, a new understanding emerged that light also possesses properties of particles. This has led to our understanding of the dual nature of light. Depending on the experiment the behavior of light can be described in terms of waves, or in terms of particles.

<u>Light as a Wave</u>

Light travels through space as electromagnetic radiation in the form of a wave. Visible light is only one part of many wavelengths that together make up the electromagnetic spectrum.

What are other electromagnetic radiation<u>s</u> that you know of from reading this section?

<u>Answers</u>:

Do radio waves have a higher or lower frequency than visible light?

<u>Answer:</u>

Which end of the visible spectrum of light has the highest frequency and therefore the most energy, the red end or the violet end?

<u>Answer</u>:

All forms of electromagnetic radiation travel in a vacuum at the same speed of 3.0 x 10^8 m/s which is the speed of light.

<u>Wavelength</u>

The wavelength is the <u>distance</u> between successive waves. The Greek symbol for wavelength is λ, (lambda).

Frequency

The frequency is the number of complete waves passing a fixed point in a given time. Frequency is measured in hertz (Hz). The Greek symbol for frequency is υ (nu).
A hertz equals 1 cycle per second, or 1 Hz = 1 s^{-1}, in which the $^{-1}$ is the reciprocal, or 1 Hz = 1/1s.
The relationship between frequency and wavelength is described by the equation $\lambda = c/\upsilon$.
These relationships between properties of electromagnetic waves demonstrate that**the shorter the wavelength of a wave the greater its frequency.**

The Photoelectric Effect Light as a Particle

The photoelectric effect occurs when electrons are emitted from the surface of a metal when light strikes the metal. There is a certain minimum frequency of light that causes this effect.
Max Planck proposed the term quantum of energy or the minimum energy that can be lost or gained by an atom in an attempt to explain the photoelectric effect. In this respect the quantum acts as a particle. The energy of the quantum is equal to the frequency of the radiation multiplied by a constant labeled h and equal to 6.626×10^{-34}J.s. The entire equation is E = hv.
Einstein expanded upon Planck's work and stated that each particle of light carries a quantum of energy he called a photon. This photon must carry a minimum amount of energy to knock loose an electron from a metal.

The Hydrogen-Atom Line-Emission Spectrum

Much of today's understanding of the electron came from the work of Niels Bohr in his studies of the hydrogen atom and its single electron. This led to today's modern atomic theory called the quantum theory.
Summary:
The hydrogen atom's electron travels about the nucleus in defined orbits or energy levels. The lowest energy level or the path closest to the nucleus is called the ground state. When energy is absorbed by this electron called absorption, the electron moves to a higher energy level farther from the nucleus called the excited state.
A certain amount of energy must be applied to move the electron from one level to another. This movement is all or nothing, it moves from one level to the next not in between. The electron's movement to a defined level can be represented by whole number integers.
When an electron loses energy and returns to the ground state, the absorbed energy is emitted as photons of light. When this emitted light is passed through a prism, lines appeared at certain wavelengths. This is called a line-emission spec-

trum. These lines fell at certain points of the spectrum of light. If the energy was emitted in all frequencies it would have produced a continuous-line spectrum that we see when light is passed through a prism.

Bohr concluded that the photon's energy emitted as it fell from one level to the lower level was equal to the frequency of the photon or $E_{photon} = hv$.

The spectral lines corresponded to the frequency of the emitted photons which confirmed Planck's and Einstein's work. It also supplied evidence for orbitals or regions where electrons traveled around a nucleus.

This is the reason that a simple understanding of light is needed to understand the relationship between electron's and the light they emit as a means of finding an electron's relative location around a nucleus.

Bohr's work however did not explain more than one electron or chemical activity.

Answer the following questions.

1. **What was the major shortcoming of the Rutherford model?**

2. **Write the equation for:**
   a. **the relationship between frequency and wavelength.**

      $\lambda =$

   b. **the energy of a photon**

      $E_{photon} =$

3. **What part of the electromagnetic spectrum are humans able to see?**

4. **What is the relationship between wavelength and frequency?**

5. **What is the name of the type of spectrum which we see in visible light?**

Section 2: The Quantum Model of the Electron pgs. 104-110.

Objectives:

1. **Discuss Louis de Broglie's role in the development of the quantum model of the atom.**
2. **Compare and contrast the Bohr model and the quantum model of the atom.**

3. Explain how the Heisenberg uncertainty principle and the Schrodinger wave equation led to the idea of atomic orbitals.
4. List the four quantum numbers and describe their significance.
5. Relate the number of sublevels corresponding to each of an atom's main energy levels, the number of orbitals per sublevel, and the number of orbitals per main energy level.

Vocabulary: Define the following.

1. Heisenberg uncertainty principle--

2. quantum theory--

3. orbital--

4. quantum number--

5. principal quantum number--

6. angular momentum quantum number--

7. magnetic quantum number--

8. spin quantum number--

<u>Electrons as Waves</u>

In 1924 Louis de Broglie stated that electrons acted as waves confined to a space around a nucleus. These waves existed at specific frequencies. This was confirmed by experiments showing that electrons like waves could be bent (diffracted) or show interference (waves can either combine showing larger energies or cancel each other out exhibiting low energy).

<u>Heisenberg Uncertainty Principle</u>

Werner Heisenberg stated that it is impossible to determine simultaneously both the position and velocity of an electron. This is the result of the fact that photons and electrons have about the same energies. The illumination of an electron by photon so the electron can be seen would deflect the orbit of the electron with the result that we can never be sure of its exact position. This is reflected in today's understanding of electrons existing as an <u>electron cloud</u> rather than as a defined orbit as Rutherford assumed.

<u>Schrodinger Wave Equation</u>

Schrodinger reasoned that if electrons acted as waves then electrons must have specific frequencies, however as Heisenberg pointed out the positions of these electrons can not exactly be known but would best be described in mathematical terms that would show the probability of finding an electron in some region around a nucleus.

In contrast to the Bohr model of the atom that states that electrons exist in defined region, the quantum theory states that only the probable location of an electron can be known.

<u>Atomic Orbitals and Quantum Numbers</u>

Scientists use 4 quantum numbers to specify the properties of the orbits and the electrons within them. No two electrons can have the same quantum numbers.

<u>The Principle Quantum Number, *n*</u>

The principle quantum number or *n* is the main energy level that an electron occupies and are integers from 1 to 7. The greater the number the greater the energy and the farther an electron is from the nucleus.

More than one electron can be in an energy level but they are found in different regions of the level or orbitals. The largest number of orbitals in any given level is equal to n^2. For example in the third level there are found up to 3^2 or 9 orbitals.

<u>Angular Momentum Quantum Number, *l*</u>

The quantum number *l* indicates the shape of the orbital and these shapes are called sublevels. The values of *l* or the different shapes are equal to the *n* value minus 1. The sublevels are given a further designation of s, p, d, and f. For instance the s sublevels looks like a sphere, while the p sublevels appear as a figure 8. The d and f sublevels have more complex shapes.

n = 1 has 1 sublevel an s	**The *l* value = 0**
n = 2 has 2 sublevels s and p	**The *l* value = 1**
n = 3 has 3 sublevels s, p, and d	**The *l* value = 2**

n = 4 has 4 sublevels s, p, d, and f. The *l* value = 3

Each atomic orbital is listed by the principle quantum number followed by the letter of the sublevel. For instance an electron in the second level found in a p sublevel would be written as 2p. An electron in level 4 in a d sublevel is written as 4d. How is an electron in p sublevel of the 3 energy level written?
Answer:

In this level what are the other sublevels?
Answer:

Magnetic Quantum Number, *m*

The angular momentum number describes the different shapes of the sublevels while the magnetic quantum number describes the 3 dimensional orientation of each sublevel in space. The *m* values are whole numbers from −*l* to 0 to +*l*.
For instance the f sublevel has an *l* value of 3. The magnetic quantum numbers would be −3, −2, −1, 0, +1, +2, +3. Each of these 7 numbers would be tied by dimensional diagrams to a particular organization in space. This means that in total there are 7 orbitals in an f sublevel.

Spin Quantum Number

Electrons repel each other, unless they spin in opposite directions on their axes. Therefore any orbital can hold only a maximum of two electrons.
Spin does not have a letter value as the other quantum numbers but is indicated by a +1/2 or −1/2.

***A note to the student about quantum numbers*.**

Quantum numbers have become the darling child to the nuclear physicists and chemists in research. To introductory students they are perplexing and indeed they are seldom referred to again throughout the textbooks written at this level. It has become traditional to teach these numbers but only a simple understanding of these quantum numbers is necessary to continue on with the course. It will not be necessary to use the quantum numbers in this course, however, any students going onto AP chemistry will encounter quantum numbers again.

What you need to know

Students will need to know the following:

The principle quantum number *n* is also called the main energy level. This corresponds to the major regions (distances) around the nucleus 1 through 7. Within the major levels there are differently shaped sublevels (angular momentum). The sublevels shapes are s, p, d, and f. Within these sublevels are the actual electrons moving in a particular 3 dimensional way (magnetic quantum number). The dimensional alignment is called the orbital. The orbitals have the same name as the sublevel. An s sublevel has 1 s orbital, a p sublevel has 3 p orbitals, a d sublevel has 5 d orbitals, and an f sublevel has 7 f orbitals. Each of the orbitals can hold a maximum of 2 electrons. Small atoms have limited space and can only have a few number of sublevels within a level while larger atoms can have more main levels with more sublevels. As the atoms in the periodic table become larger the number of levels increases meaning there is more room to fit more sublevels with electrons arranged in more possible locations. A research physicist or chemist will use quantum numbers to describe the location. We will use a more simplified version using only level number, sublevel name, and total number of electrons in the orbitals of the sublevel.

Why does this need to be studied?

In the coming chapters the formation of chemical compounds will be studied. Compounds are formed when electrons from different atoms interact with each other. The only electrons that interact are the outermost electrons. The understanding of electrons based on quantum numbers will tell us which electrons are the outermost electrons. Knowing the outermost electrons called valence electrons will allow us to determine the chemical formulas for these compounds.

Answer the following questions.

1. **What letter is used as the principle quantum number?**

2. **What does the principle quantum number tell us?**

3. **What is the shape of a p sublevel?**

4. **How many orbitals are found in a d sublevel?**

5. **What is the maximum number of electrons in an orbital?**

6. How does the sublevel and orbital names compare?

7. For any two electrons in the same orbital how do their spins compare?

8. What effect does a photon have on the position of an electron when it strikes an electron?

9. What is another name for a main energy level?

<u>Section 3 Electron Configurations pgs. 111-122.</u>

<u>Objectives:</u>

1. List the total number of electrons needed to fully occupy each main energy level.

2. State the Aufbau principle, the Pauli exclusion principle, and Hund's rule.

3. Describe the electron configurations for the atoms of any element using orbital notation, electron configuration, and when appropriate, noble-gas notation.

<u>Vocabulary:</u> Define the following.

1. electron configuration--

2. Aufbau principle--

3. Pauli exclusion principle--

4. Hund's rule--

5. noble gas--

6. noble-gas configuration--

Rules for Electron Configuration

The ground state is the lowest energy arrangement of electrons. The arrangement of electrons in an atom is known as the atom's electron configuration.
There are 3 rules that govern how electrons fill in around the nucleus of the atom.

Aufbau Principle

The Aufbau principle states that electrons fill the lowest energy orbital first. Electrons fill in the following manner. The lowest energy electrons fill the 1s then the 2s followed by the 2p. In the third level the energies of the sublevels begin to overlap so that the first the 3s fills then the 3p fills but instead of the 3d level filling the 4s fills because it is of a lower energy than the 3d sublevel. You will be given a handout on this.
After the 3d level fills the 4p sublevel begins to fill. Though this seems quite a bit to memorize there are some aids to assist you in writing the configurations which will be given to you to use when you begin to write electron configurations.
Before further electron configurations are discussed the second rule of electron configuration must be examined. The is the Pauli exclusion principle which states that no two electrons can have the same 4 quantum numbers. This means that in dealing with two electrons in the same orbital, three of the 4 quantum numbers will be the same but they must spin in opposite directions. Thus the 4th quantum number will be different. To show this fact up and down arrows are used to indicate opposite spins.
The third rule is called Hund's rule which states that orbitals of equal energy are occupied by one electron before the second electron is added to any similar orbital. The first electrons added in multiple orbitals of a sublevel must have the same spin direction.
The following diagram called orbital notation shows how both the Aufbau principle, Hund's rule and the Pauli exclusion principle are depicted.
Recall that in the p sublevel there are 3 p orbitals. Each orbital can hold either one or two electrons. If all the p orbitals of any p sublevel are completely filled there would be 6 electrons.
Before the p sublevel orbitals fill the s sublevel of that energy level must fill first. (Aufbau principle).
This is how the p orbitals would fill after the s has been filled. The first electron would enter one of the three p orbitals and would be written as an arrow up or down. (Remember that whatever direction you choose the next electron to go into a p orbital must also be in the same direction).

<u>Orbital Notation Diagrams</u>

↑↓	↑		
s	**p**	**p**	**p**

The second electron available would enter the next p orbital.

↑↓	↑	↑	
s	**p**	**p**	**p**

The third electron would enter the third p orbital.

↑↓	↑	↑	↑
s	**p**	**p**	**p**

Only after each p orbital has a single electron each with the same spin will the orbitals begin to accept the second electron. (<u>Hund's rule).</u>
The fourth electron would begin to occupy a p orbital having a single electron but it must have an opposite spin to give it a different quantum number. (<u>Pauli exclusion principle).</u>

↑↓	↑↓	↑	↑
s	**p**	**p**	**p**

The remaining 5th and 6th electrons would then enter the remaining two p orbitals.

<u>Electron Configuration Notation</u>
Before writing electron configuration notation read the quick review below.
The principal energy levels are labeled *n*, and are called the principal quantum numbers. Within these principal energy levels are one or more sublevels.
<u>The 4 types of sublevels are labeled *s,p,d*, and *f*</u>.
<u>The number of sublevels is equal to that particular quantum number</u>. For instance quantum level 1 has only 1 sublevel. Quantum level 2 has 2 sublevels etc.
<u>Orbitals</u> are found within sublevels. Remember orbitals have certain shapes, sizes, and energies. Each sublevel has a certain number of available orbitals. *s* = 1 orbital, *p* = 3 orbitals, *d* = 5 orbitals, *f* = 7 orbitals. The orbital name is the same as the sublevel name.
Example: Principal levels and sublevels.
When n = 1 there can only be 1 sublevel called *s*.
When n = 2 there are two sublevels called *s* and *p*.
When n = 3 there are three sublevels called *s, p*, and *d*.
When n = 4 there are four sublevels called *s, p, d*, and *f*.
The Pauli exclusion principle states that electrons have spins; either clockwise or counterclockwise. Each orbital can hold at most 2 electrons with opposite spins.
Adding the Pauli exclusion principle to the above example results in the following:

<u>sublevel</u>	<u>orbitals</u>	<u>maximum electrons</u>
s	1	2
p	3	6
d	5	10

f	**7**	**14**

SUMMARY TABLE

Principal Energy Levels	**Sublevels**	**Orbitals**
n **= 1**	**1*s***	**1*s* (one)**
n **= 2**	**2s, 2p**	**2s (one) + 2p (three)**
n **= 3**	**3s, 3p,3d**	**3s (one) + 3p (three) + 3d (five)**
n **= 4**	**4s,4p,4d,4f**	**4s (one) + 4p (three) + 4d (five) + 4F (seven)**

Electron Configuration

Electron configuration notation is a type of abbreviated way of writing electron location instead of using orbital notation.
The electron configuration is the distribution of electrons among the levels, sublevels, and orbitals of an atom. The sublevel with the least energy is the s sublevel. The energies continue to increase through the p and d sublevels with the f sublevel having the highest energy.
The ground state is the lowest energy level. This is the normal and most stable level for the atom.
Above the 3p sublevel the energies of the principal energy levels begin to overlap. (See handout on the Diagonal Rule and p. 116.
The electron configuration also closely follows the periodic table and an additional handout will be given to you to help write electron configurations.

In writing the electron configuration notation method the number of electrons in a sublevel is shown as a superscript.
Hydrogen - Atomic number 1 $1s^2$
Helium - Atomic number 2 $1s^2$
Lithium - Atomic number 3 $1s^2 2s^1$
Carbon - Atomic number 6 $1s^2 2s^2 2p^2$
Sodium - Atomic number 11 $1s^2 2s^2 2p^6 3s^1$
Calcium -Atomic number 20 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
With elements higher atomic numbers than calcium the 3d sublevels which are of a lower energy begin to fill after the 4s but before the 4p sublevels.
Use the diagonal rule to write the electron configuration of nickel, atomic number 28, and rubidium, atomic number 37.

Nickel Answer:

Rubidium Answer:

Noble Gas Notation

Once the student has become familiar with writing electron configurations an abbreviated method can now be used called the noble gas notation. Elements of Group 18 are called the noble gases. To simplify electron notations the symbol for the noble gas in the lowered number row of the periodic table is enclosed in brackets signifying the complete notation up to that gas. Only the remaining notation need be written for the element that the notation is describing.
For instance magnesium in row 3 of the periodic table has the electron configuration of $1s^22s2^22p^63s^2$. The noble gas in the lower row number 2 is neon number 10. The noble gas abbreviated form for magnesium would therefore be $[Ne]3s^2$.
Scandium atomic number 21 has the configuration $1s^22s^22p^63s^23p^64s^23d^1$. The noble gas notation is $[Ar]4s^23d^1$

Write the noble gas configuration for the following.
Chlorine atomic number 17

Bromine atomic number 35

Tin (Sn) atomic number 50

Exceptions to the Aufbau Principle

Chromium (Cr) #24 and Copper (Cu) #29 are exceptions to the Aufbau Principle. These exceptions are due to the fact that half filled orbitals are more stable than configurations without any particular arrangement. It is not necessary to study these, rather it is important that one knows that they do exist.
Examples:
Chromium
Expected: $1s^22s^22p^63s^23p^6\underline{4s^2}\underline{3d^4}$
Actual: $1s^22s^22p^63s^23p^6\underline{4s^1}\underline{3d^5}$
Write the expected and actual configuration for copper.
Expected:

Actual: