Chapter 13 – Electrons in Atoms
Part 1 – Notes: Revisiting the History of Atomic Theory and Orbitals

Objectives: Summarize the development of the atomic theory from Bohr forward.
Explain the significance of quantized energies of electrons in relation to the quantum mechanical model.
Identify, define, and explain: energy level, quantum, quantum mechanical model (wave mechanical model), atomic orbitals, quantized, sublevel, ground state, excited state, probability, and principal quantum number.
Explain Aufbau Principle, Hund's Rule, and Pauli Exclusion Principle and how they are used to determine the location of an electron in an atom.
List, describe, and differentiate between s, p, d, and f atomic orbitals.
Relate the principal quantum number to the number of orbital types in an atom and to the number of individual orbitals in a given energy level.

Text Reference: Section 13.1 – pages 361-366

We left off with ***Rutherford's*** model of the atom: a small dense nucleus made of protons with electrons outside the nucleus.

The existence of the **neutron** was added to the atom in 1932.

Neils Bohr – 1885 – 1962 – Danish physicist – student of Rutherford

- ***The Planetary Motion Model*** – proposed in 1913
    - Electrons are arranged in concentric circular paths, or orbits, around the nucleus.
- The question that needed to be answered was "**why** don't the electrons crash into the nucleus?"
- Bohr's answer:

The energies of electrons are ***QUANTIZED*** – meaning only certain values are allowed.

Energy level –

- The rungs of a ladder are somewhat analogous to the fixed energy levels of the electrons. The lowest rung corresponds to the lowest energy level. A person can stand on step one or step two – but not on step one-and-a-half.

Quantum – a quantum of energy is the amount of energy required to move an electron from its present level to the next higher one

But the Bohr model of the atom was incorrect. Two young physicists, Louis Victor de Broglie and Erwin Schrödinger, suggested that just as light exhibits characteristics of waves ***and*** particles, the electron may also exhibit both those characteristics. The idea that an electron, a particle, could exhibit properties of a wave was revolutionary.

Erwin Schrödinger (1887 – 1961) and ***Louis Victor de Broglie***
Model: ***The Wave Mechanical Model*** (Also the ***Quantum Mechanical Model***)

- Schrödinger used the new quantum theory to write and solve **mathematical equations** to describe the location and energy of an electron in an atom.
- The model is derived from mathematical solutions to the Schrödinger equation.
- Schrödinger's model is primarily mathematical; there are few (if any) analogies in real life.

The Quantum Mechanical Model – a modern description of electrons in atoms

- It incorporates both *wave-particle* theory and *probability* into its description of electron behavior.
- Energies of electrons are restricted to certain values – similar to Bohr's model.

- The path of an electron around a nucleus is not exact. Instead there is a region of probability where an electron is most likely to be found. The idea is based around the likelihood of finding an electron in a certain region or position.
- This probability can be portrayed as a blurry cloud of negative charge. The **CLOUD** is most dense where the probability of finding the electron is large. The cloud is least dense where the probability of finding the electron is small. Hence, it is difficult to say where the electron cloud ends. There is no edge. There is even a probability of finding the electron a great distance away from the nucleus.
- The path of an electron is not precisely predictable; however the work of deBroglie and Schrödinger led to orbitals.

Atomic Orbital: An *orbital* is a probability map, a region in which the electron has a 90% (or 95%) probability of being found.

Principal energy levels –

Principal Quantum Number: refers to the energy level in which an electron is located – designated "***n***"

- Principal quantum numbers are positive integers, beginning with 1
- 1 is the lowest energy level, and electrons in this level have a principal quantum number of ***n*** = 1
- The higher the ***n***. the higher the average energy of the electrons in the energy level.
- The higher the ***n***, the greater the average distance of the electron from the nucleus.

S***ublevels*** –

- The number of orbitals inside the energy level depends upon the energy level itself.
- For ***n*** = 1, there is 1 type of sublevel. For ***n*** = 2, there are two types of sublevels/atomic orbitals.
- The number of atomic orbitals in a given energy level is equal to the principal quantum number.

There are four basic types of orbitals: s, p, d, and f. There are additional orbitals designated g, h, i, j, k, …

	Shape	***Groups of…***	***Begin where?***	***Relative energy***
s-orbital				
p-orbital				
d-orbital				
f-orbital				

Within a given energy level, the energy increases accordingly: s < p < d < f

- but this order does not hold true when you are talking about different energy levels.

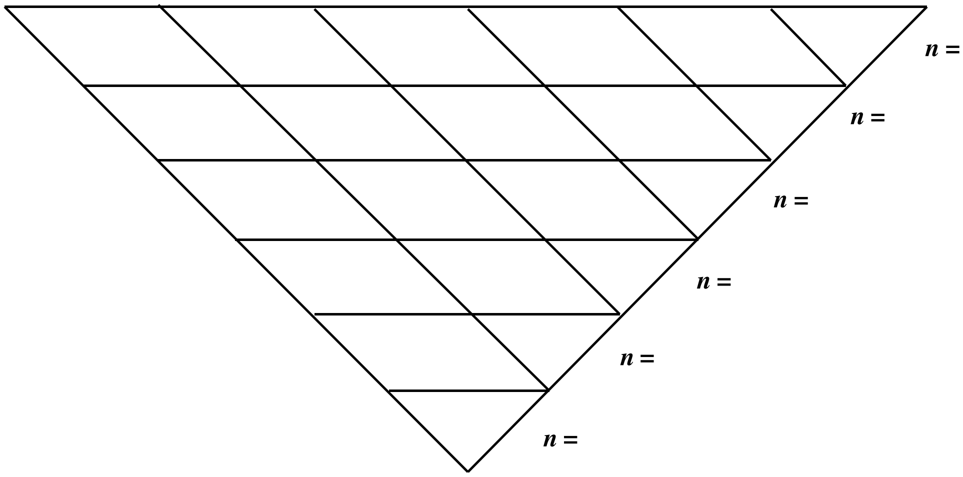
The size of the orbital increases with each energy level – but the shape remains the same.

- A 1s orbital and a 2s orbital are both spherical, but the 1s orbital is smaller than the 2s orbital.

Relationship between ***n*** and number of different types of orbitals: ***n* = number of sublevels = # of orbital types**

Relationship between ***n*** and the number of orbitals: **# of orbitals = n^2**

If each orbital can hold 2 electrons, relationship between ***n*** and number of orbitals: **# of possible electrons = 2 n^2**



sublevels	orbitals

Rules to Remember

1. Energy levels, ***n***, increase by integer values.
2. The number of orbitals per energy level is equal to n^2. the first energy level will have one orbital. The second energy level, ***n*** = 2, will have 4 orbitals. The third energy level, ***n*** = 3, will have nine orbitals. And so on.
3. Every energy level has an ***s*** orbital. The s-orbital is sphere shaped.
4. Energy levels ***n*** = 2 and higher contain a set of three p-orbitals: p_x, p_y, and p_z. The p-orbitals have two lobes.
5. Energy levels ***n*** = 3 and higher contain a set of five d-orbitals: d_{xy}, d_{xz}, d_{yz}, d_{z2}, and d_{x2-y2} that have four lobes.
6. Energy levels ***n*** = 4 and higher contain a set of seven f-orbitals. The subscripts are not important for our purposes; just know the total number. They also have shapes more complex than can be simply described.
7. Aufbau Principle (Building Up) –

8. Pauli Exclusion Principle –

9. Hund's Rule –

Electron Configuration Notation: a b^c

a = b = c =

Chapter 13 – Electrons in Atoms
Part 1 – Assignment: Revisiting the History of Atomic Theory and Orbitals

Provide complete answers to the following questions.

1. Why was Bohr’s theory for the hydrogen atom initially accepted, and why was it ultimately discarded?

2. What major assumptions (that was analogous to what had already been demonstrated for electromagnetic radiation) did de Broglie and Schrödinger make about the motion of electrons?

3. Discuss briefly the difference between an orbit (as described by Bohr for hydrogen) and an orbit (as described by the more modern picture of the atom).

4. In general terms, explain how the quantum mechanical model of the atom describes the electron structure of an atom.

5. How many orbitals are in the following sublevels:

   3p = ____________________ *2s* = ____________________ *4f* = ____________________

   4p = ____________________ *3d* = ____________________ *6g* = ____________________

6. How did Bohr answer the objection that an electron traveling in a circular orbit would radiate energy and fall into the nucleus?

7. What is the significance of the border of the electron cloud?

8. What is an atomic orbital?

Chapter 13 – Electrons in Atoms
Part 2 – Notes: Electron Configurations and Orbital Diagrams

Objectives: Apply the Aufbau principle, the Pauli Exclusion Principle, and Hund's Rule in writing electron configurations of elements.
Identify, define, and, explain: electron configuration, Aufbau principle, Hund's Rule, Pauli Exclusion Principle, and orbital diagram.
Explain why all of the 3^{rd} energy level does not fill before the 4^{th} energy level starts to fill.

Text Reference: Section 13.2 (Part) – pages 367-369

Key Items to Keep in Mind:

1. Energy levels, ***n***, increase by integer values.
2. The number of orbitals per energy level is equal to n^2. the first energy level will have one orbital. The second energy level, ***n*** = 2, will have 4 orbitals. The third energy level, ***n*** = 3, will have nine orbitals. And so on.
3. Every energy level has an ***s*** orbital. The s-orbital is sphere shaped.
4. Energy levels ***n*** = 2 and higher contain a set of three p-orbitals: p_x, p_y, and p_z. The p-orbitals have two lobes.
5. Energy levels ***n*** = 3 and higher contain a set of five d-orbitals: d_{xy}, d_{xz}, d_{yz}, d_{z2}, and d_{x2-y2} that have four lobes.
6. Energy levels ***n*** = 4 and higher contain a set of seven f-orbitals. The subscripts are not important for our purposes; just know the total number. They also have shapes more complex than can be simply described.
7. Aufbau Principle – Electrons fill orbitals from lowest to highest energy.
8. Pauli Exclusion Principle – No two electrons in the same atom can have the same four quantum numbers. In other words, an orbital can hold at most two electrons and if there are two electrons, they have paired spins.
9. Hund's Rule – When electrons occupy a set of degenerate orbitals, one electron enters each orbital until all orbitals contain one electron with their parallel spins.

Electron Configurations utilize the (first) principal quantum number, ***n***, and the second quantum number, ***l***, along with an integer superscript to represent the organization of an atom's electron from lowest to highest energy.

Example Electron Configuration: $1\,s^2$ (read "one s two" NOT "one s-squared")

The ***1*** represents

The ***s*** represents

The superscript ***2*** indicates

Another way to represent the electrons of an atom is through the use of an orbital diagram, also known as a box diagram. Orbital diagrams are drawn with the first two numbers above the box (or series of boxes). One or two arrows are placed in the box to indicate the number of electrons present. If there are two arrows in the box, representing two electrons in the orbital, they must have opposite spins (Pauli Exclusion Principle). The arrows are pointed in opposite directions to represent the opposite spins.

<u>Electron Configuration and Orbital Diagrams for the First 18 Elements</u>

At this time, all elements are considered to be in their ground state, the state of least energy.

Hydrogen: The neutral hydrogen contains 1 electron.

Helium: Recall the Pauli Exclusion Principle as you draw in the second electron.

Lithium: The first energy level fills and then the second energy level starts to fill; recall Aufbau Principle.

Beryllium: The fourth electron completes the 2s orbital.

Boron: Boron has 5 electrons, four of which fill the 1s and the 2s orbitals. The fifth electron must fill the orbital of the next highest energy, the 2p. Since there is a group of three p-orbitals, it is uncertain which fills first, by convention, it is assumed that the 2 p_x orbital fills first.

Carbon: Carbon has six electrons. Two occupy the 2p orbital. Recall Hund's Rule. For reasons not considered here, in the separate 2p orbitals, the electrons have the same spin.

When the electron configuration is written with a single term to indicate the two p-orbitals, it is understood that the electrons are in separate orbitals, even though the configuration does not indicate it; it may be written this way because we know that Hund's Rule must be fulfilled.

Nitrogen: Nitrogen has 7 electrons. Each of the electrons in the 2p orbitals will have the same spin.

Oxygen: Oxygen has 8 electrons. The eighth electron is paired with the 5th in the $2p_x$ electron; each has opposite spin.

Fluorine: Fluorine has 9 electrons. The 9th electron is paired with the 6th electron in the $2p_y$ orbital; each has opposite spin.

Neon: Neon has 10 electrons and they completely fill the orbitals of the first and second energy levels.

Sodium: Sodium has 11 electrons. The 11th electron must enter the 3rd energy level, since the 2nd level is filled.

Magnesium: Magnesium has 12 electrons.

Aluminum: Aluminum has 13 electrons. The 13th electron starts to fill the 3p orbitals.

Silicon: Silicon has 14 electrons. When filling the 3p orbitals, remember Hund's Rule.

Phosphorus: Phosphorus has 15 electrons.

Sulfur: Sulfur has 16 electrons.

Chlorine: Chlorine has 17 electrons.

Argon: Argon has 18 electrons.

What happens after the 18th electron fills into the 3p sublevel? What orbital do you think will fill next?

According to the Aufbau principle, we know orbitals fill in order from low to high energy. Examine the energy diagram to determine which orbital fills next.

Chapter 13 – Electrons in Atoms
Part 2 – Assignment: Electron Configurations and Orbital Diagrams

Answer the following questions. Be complete and neat.

1. Arrange the following sublevels in order of ***decreasing*** energy: *2p, 3s, 3p, 3d,* and *4s*

   __________ __________ __________ __________ __________

   Explain why the above trend is true.

2. Why does one electron in a potassium atom go into the fourth energy level instead of squeezing into the third energy level along with the eight electrons already present?

3. What is meant by *3p3*? Draw an orbital diagram to ***help*** explain your answer.

4. What is the maximum number of electrons that can go into each of the following sublevels?

a. *2s*	__________	b. *3p*	__________	c. *4s*	__________
d. *3d*	__________	e. *4p*	__________	f. *5s*	__________
g. *4f*	__________	h. *5p*	__________	i. *7g*	__________

5. Which of these orbital designations are invalid: *4s* *3f* *2d* *3d* *4g* *6g* *1p*

6. How many paired electrons are there in an atom of each of the following:

   a. helium b. boron

   c. sodium d. oxygen

7. An atom of an element has two electrons in the first energy level and five electrons in the second energy level. Write the electron configuration for the atom. State the name of the element. How many unpaired electrons does an atom of this element have?

8. An atom of a specific element has two electrons in the first energy level, 8 electrons in the second energy level, and five electrons in the third energy level. Write the electron configuration for the atom. State the name of the element. How many unpaired electrons does an atom of this element have?

Chapter 13 – Electrons in Atoms
Part 3 – Notes: More Electron Configurations & Valence Electrons

Objectives: Apply the Aufbau principle, Hund's Rule, and Pauli Exclusion Principle in writing electron configurations of elements.
Identify, define, and, explain: electron configuration, Aufbau principle, Hund's Rule, Pauli Exclusion Principle, and valence electron.
Explain why all of the 3^{rd} energy level does not fill before the 4^{th} energy level starts to fill.

Text Reference: Section 13.2 (Part) – pages 367-370

You know the electron configuration for argon is ____.

You also know that the third energy level, ***n*** = 3, has three different sublevels: s, p, and d.

Because *s* and *p* orbitals of the third energy level are filled in argon, you might expect that the 19^{th} electron of element 19, potassium, to be placed into a *3d* orbital; however, experiments show that the chemical properties of potassium are very similar to lithium and sodium. Chemists associate similar chemical properties with similar valence-electron arrangements; hence it is predicted that the outer electrons of potassium be s^1, resembling sodium ($3s^1$) and lithium ($2s^1$). We expect the last electron in potassium to occupy the 4s orbital instead of one of the 3d orbitals. This means that the fourth principal energy level begins to fill before the third energy level has been completed.

Examine the diagram of energy levels and orbitals. How is it that the 4^{th} energy level begins to fill before the third energy level has been completed?

Potassium:

Calcium:

Now, scandium has 21 electrons. Where does the 21^{st} electron go?
Careful studies of energies of ground state atoms have given a filling order that is represented in the energy diagram.

Scandium:

Note the remaining electron configurations for the transition metals in the fourth period will follow a similar pattern.

After 3p fills, _______________ fills, followed by _______________, and _______________.

Let's put the filling order to use as we write the ground state electron configurations for the following atoms.

Titanium:

Nickel:

Copper:

Zinc:

Arsenic:

Krypton:

Molybdenum:

Lead:

Tungsten:

<u>Valence Electrons</u>

Valence electrons - electrons in the outermost orbitals of an atom

Examples 1 and 2

Electron configuration of:	Nitrogen	Phosphorus
Highest energy level		
Electrons in the highest energy level		
Total number of valence electrons		

Examples 3 and 4

Electron configuration of:	Sodium	Lithium
Highest energy level		
Electrons in the highest energy level		
Total number of valence electrons		

The valence electrons are the most significant to chemists because those are the electrons involved when atoms attach to each other (form bonds). The inner electrons are known as ***core electrons***. The core electrons are not involved in bonding atoms to one another.

If you study the configurations of the elements in the same groups you will find that, with the exception of helium, elements of the same group…

EXAMPLES 5 – 10:

Element	***Symbol***	***Group***	***Configuration of Valence Electron***
Neon			
Argon			
Magnesium			
Calcium			
Sulfur			
Selenium			

ELEMENTS WITH THE SAME NUMBER OF VALENCE ELECTRONS (elements in the same family) EXHIBIT SIMILAR CHEMICAL BEHAVIOR!!!!!

Chapter 13 – Electrons in Atoms
Part 3 – Assignment: More Electron Configurations & Valence Electrons

Write the electron configurations for the following elements. Also, write the number of valence electrons in each.

1. zirconium

2. hafnium

3. rutherfordium

4. silver

5. tin

6. ytterbium

7. uranium

8. antimony

9. How many electrons are in the second energy level of an atom of each of the following elements?

   a. chlorine b. phosphorous c. potassium

10. Write the electron configuration for an arsenic atom.

    Calculate the total number of electrons in each energy level.

    State which energy level(s) is/are not full.

11. Ms. Anderson's favorite equation is that of an active metal + water. You react sodium with water and you get sodium hydroxide and hydrogen gas. You react potassium with water and you get potassium hydroxide and hydrogen gas. What do you get if you react lithium with water? Why is there similarity between the products and also the equations (including the coefficients) when you react an alkali metal with water?

Chapter 13 – Electrons in Atoms
Part 4 – Notes: Anomalous and Abbreviated Electron Configurations

Objectives: See objectives from Chapter 13 Part 2 and 3.
Write abbreviated configurations for elements.
Explain why the electron configurations for the chromium and copper columns do not fit the standard pattern.
Correctly write the electron configurations for chromium molybdenum, tungsten, copper, silver, and gold, and explain why it is written in such a manner.

Text Reference: Section 13.2 (Part) – pages 367-370

<u>**Anomalies of Electron Configurations**</u>

Based on your current knowledge, write the electron configuration of chromium and copper.

Chromium:

Copper:

These electron configurations are actually **INCORRECT**!!! Take a moment to cross them out.

Atoms in the ground state have the lowest possible energy. Atoms want to have the lowest possible energy. That is why some of them form ions and some form covalent compounds, to obtain a state of minimum energy.

When electrons fill into the orbitals, they do so in such a manner as to minimize the energy of the atom. ***There is a special stability that comes from a half-full or a filled energy level.*** It is this special stability that will give use the correct electron configuration for chromium and copper.

Chromium's correct electron configuration is:

Note, this happens because

Copper's correct electron configuration is:

Note, this happens because

Note, the 3d orbital is completely filled and the 4s orbital is half-filled. It is more stable to have the 3rd energy level completed than it is to have the 3rd energy level missing one and the 4th energy level just started. So, for copper, the 3rd energy level is completed which leads to greater stability (low energy) to the atom.

There are a number of other anomalies that occur throughout the periodic table. You will be responsible for the anomalies of chromium, copper, molybdenum, silver, tungsten, and gold. Mo, W, Ag, and Au follow the same valance configuration as chromium and copper, based upon the column in the periodic table.

There are other anomalies for which you are not held responsible.

Note the following rules for writing the electron configurations:

1. In a principal energy level that has *d* orbitals, the s-orbitals from the next level before the d-orbital in the current level.
2. After lanthanum, which has the electron configuration [Xe] $6s^2$ $5d^1$, a group of fourteen elements called the lanthanide series occurs. This series of elements corresponds to the filling of the seven 4f orbitals.
3. After actinium, which has the electron configuration [Rn] $7s^2$ $6d^1$, a group of fourteen elements known as the actinide series occurs. This series corresponds to the filling of the seven 5f orbitals.

Abbreviated electron configurations

Sometimes writing electron configurations is a long and tedious process. Is there a shorter way???

The knowledge of valence electrons, electron configurations, and noble gases may be used to write abbreviated electron configurations. Abbreviated electron configurations allow us to write electron configurations quickly.

Electron configuration for **NEON:**

Electron configuration for **MAGNESIUM:**

Abbreviated configuration for **MAGNESIUM:**

Write the abbreviated electron configuration for the following elements:
Sodium:

Aluminum:

Silicon:

Phosphorus:

Sulfur:

Argon:

Note that for all of the elements except the noble gases, the electrons represented after the noble gas are also the valence electrons. The noble gas core used is actually abbreviating the core electrons.

Ground State versus Excited State
Hydrogen has only one single electron and this electron fills into the lowest energy level. Although only the first energy level is used, hydrogen still has a full set of atomic orbitals that are not used in the **GROUND STATE** of hydrogen. However, in the **EXCITED STATE**, the single electron will go into a higher energy level. So, even though in the ground state, only the first energy level is used, hydrogen has the other atomic orbitals so they may be used when necessary – in the excited state or during bonding.

Chapter 13 – Electrons in Atoms
Part 4 – Assignment: Anomalous, and Abbreviated Electron Configurations

For each of the following elements, write the complete electron configuration and list the number of valence electrons.

1. thorium

2. tellurium

3. technetium

4. mercury

For each of the following, write the abbreviated configuration and list the number of valence electrons.

5. phosphorous

6. arsenic

7. barium

8. Give the symbol and name of the element to which theses electron configurations correspond.

   a. $1s^2\ 2s^2\ 2p^6\ 3s^1$

   b. $1s^2\ 2s^2\ 2p^3$

   c. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$

Chapter 13: Electrons in Atoms
Part 5 – Notes: Electromagnetic Radiation and Waves

Objectives: Calculate the wavelength, frequency, or speed of a wave.
Identify, define, and explain: electromagnetic radiation, amplitude, wavelength, frequency, hertz, spectrum, wave, crest, trough, electromagnetic spectrum, visible spectrum, sound waves, and speed of light.
Differentiate between sound waves and radio waves.
List the parts of the electromagnetic spectrum and their relative energies.
List the components of the visible spectrum in order of frequency or wavelength.
Calculate the wavelength or frequency of a radio station signal.
Differentiate between electromagnetic waves and non-electromagnetic waves in terms of speed and energy.
State and explain the relationship between wavelength and frequency.

Text Reference: Section 13.3 (Part) – pages 372-375

<u>WAVES</u>

Waves transmit energy through a medium. If you give a stone kinetic energy (you throw it) and it lands in the middle of a pond with a smooth surface, "ripples" will form on the surface of the water. Those "ripples" are waves. The energy from the stone is transferred through the medium, water, in the form of waves. If a twig is floating on the surface of the water, the waves will move the twigs vertically (up and down) but will not carry the twig horizontally.

Waves can be represented using drawings and mathematical equations.

The ***crest*** of a wave is the top peak of the wave. The wave's ***trough*** is the bottom point of the wave.

An imaginary line may be drawn horizontally at an equal distance from both the crest and the trough of the wave. The ***amplitude*** of a wave is the distance from the imaginary line to the crest or trough of a wave.

The ***frequency*** of a wave is the number of waves that pass a given point in a specified unit of time.

The symbol for frequency is the Greek letter *nu*. *Nu* = ____________.

The unit for frequency is ***hertz***, which is abbreviated ***Hz***. One hertz is equal to one cycle per second.

Wavelength is the distance between similar points in a set of waves, such as from crest to crest or trough to trough.

The symbol for wavelength is the Greek letter *lambda*. *Lambda* = ____________.

The most common unit used when expressing wavelength is meter; however, the ***Angstrom*** is also used.

Angstrom = ____________ 1 Angstrom = 1 x 10^{-8} cm (exactly) = 1 x 10^{-10} m (exactly)

The ***speed*** of any wave equals wavelength times frequency. Examine how the units define your relationship.

Formula for the speed of a wave:

Example 1: A water wave has a frequency of 4.75 x 10^{-2} Hz and a wavelength of 1.50 x 10^{1} m. Calculate the wave's speed.

Example 2: The speed of a wave is 4.75 m/s and its frequency is 8.35 Hz. Calculate the wavelength.

ELECTROMAGNETIC RADIATION

Electromagnetic radiation is energy that can travel through a vacuum, in the form of waves and at the speed of light. Electromagnetic radiation has no mass.

NOTE: *For the time being, whenever you are solving problems that involve electromagnetic radiation, assume that all electromagnetic waves travel at the speed of light even though they are not in a vacuum.*

The speed of light, c, is 3.00×10^8 meters/second. The speed of any wave is equal to the product of its wavelength and frequency. From this information it is easy to derive that when electromagnetic waves are concerned: speed of light = wavelength times frequency.

FORMULA:

NOTE: *Whenever you are solving problems using the formula give above, make certain that all measurements for wavelength are expressed in meters. If the wavelength is given in Angstroms, convert Angstroms to meters then apply the formula.*

Relationship between ***frequency*** *and* ***wavelength:***

For a complete diagram of the electromagnetic spectrum, see figure 13.10 on page 373 of your textbook. Note that visible light makes up a very small portion of the whole electromagnetic spectrum.

You should recall from previous science classes that the visible light consists of seven different colors. These colors listed in order of ***increasing frequency*** are:

Of the parts of the visible spectrum, which has the *lowest wavelength*?_______________ *lowest frequency*? _______________

Of the parts of the visible spectrum, which has the *highest wavelength*?_______________ *highest frequency*? _______________

There are no precise boundaries between the various types of waves that compose the electromagnetic spectrum; and there are not precise boundaries between the various colors of visible light. *However*, the following frequencies are associated with the colors indicated.

WAVE	*FREQUENCY*
Red light	4.3×10^{14} Hz
Yellow light	5.2×10^{14} Hz
Blue light	6.4×10^{14} Hz
Violet light	7.5×10^{14} Hz

Radio Waves

Radio stations send out radio waves on a specific frequency. Depending on the strength of the broadcasting antenna, the listening area may be large or small. No two broadcasting signals may be the same in overlapping areas.

For AM radio stations, the station call number is the frequency of the radio wave in *kilohertz*, kHz (x 10^3Hz). The frequencies range from 53 to 170 on the AM band. For FM radio stations, the station call number is the frequency of the radio wave in *megahertz*, MHz (x 10^6 Hz). The frequencies go from 88 to 108 on the FM band. These individual frequencies have associated wavelengths that may be determined through calculations.

Example 1: A gamma ray has a frequency of 3.75×10^{23} Hz. What is the wavelength?

Example 2: What radio station sends out a signal with a wavelength of 3.25 m?

Chapter 13: Electrons in Atoms
Part 5 – Assignment: Electromagnetic Radiation and Waves

Answer the following questions, neatly and completely. Show all set-ups, work, units, etc.

1. The speed of sound at $15.0^{o}C$ is 340. m/s. What is the wavelength of a sound wave with a frequency of 27.5 Hz?

2. A stone is tossed into a pond. The kinetic energy of the stone is transferred to the water and produces a wave. A cork floating in the water floats up and down at a rate of 15.0 times in 5.00 seconds. What is the frequency of the wave produced in hertz?

3. Calculate the wavelength of the waves produced by the stone from question 2 in Angstroms. (Use the answer from #2.)

4. List the colors of the visible spectrum in order of increasing frequency?

5. What radio station sends a signal with a wavelength of 3.007 m?

6. What is the wavelength of the signal sent from the antenna of WDHA (105.5 FM)?

7. What is the wavelength of the signal sent from the antenna of 770. AM?

8. A beam of infrared light has a frequency of 5.469×10^{13} Hz. What is the wavelength?

9. A hydrogen lamp emits several lines in the visible region of the spectrum. One of these lines has a wavelength of 6.56×10^{-5} cm. What are the color and frequency of this radiation?

10. A mercury lamp emits radiation with a wavelength of 4.36×10^{-7} m.
    a. What is the wavelength of this radiation in centimeters? in Angstroms?

    b. In what region of the electromagnetic spectrum is this radiation?

    c. Calculate the frequency of this radiation.

Chapter 13 – Electrons in Atoms
Part 6 – Notes: Energy, Photons, Bright-Line Spectrum, and the Photoelectric Effect

Objectives: Calculate the wavelength, frequency, or energy of light.
Explain the origins of the atomic emission spectrum of an element.
State and explain the relationship between energy and wavelength and between energy and frequency.
Differentiate between ground and excited states and explain how it is easier to remove an electron from a higher energy level.
Identify, define, and explain: spectrum, atomic emission spectrum, Planck's constant, photons, photoelectric effect, ground state, de Broglie equation, and Heisenberg uncertainty principle.

Text Reference: Section 13.3 – pages 374-382

The Bright-Line Spectrum

- You are aware that every frequency in the electromagnetic spectrum is associated with a specific quantum energy.
- When electromagnetic energy is put through a light bulb, the light bulb releases photons with energies associated with all of the frequencies from red to violet light. The combination of all the frequencies from red to violet produces white light. When the white light is observed through a prism, a continuous "rainbow" of colors appears to the observer.
- Atoms of a gas can be excited by passing electricity through a gas contained inside a glass tube.

Expectation:

Reality:

Why:

Ground State:

Excited States:

You are familiar with the fact that electromagnetic radiation displays characteristics of both waves and particles: the ***wave-particle duality***. Electromagnetic radiation is transferred to matter in *units* or ***quanta of energy*** called ***PHOTONS***.

The energy of a photon is directly proportional to the frequency of electromagnetic radiation. So, as the frequency of an electromagnetic waves increases, the energy of the photons from that wave will also increase.

Each frequency has a specific energy. The relationship between energy of a photon and frequency can be expressed by the following mathematical relationship.

Energy of a photon = *Planck's constant* x frequency

or

Formula:

The symbol for **energy** is **E**.

You should already know the unit for energy is ***joule*** and is abbreviated _______. A *joule* is __________________________.

You know that the symbol for frequency is __________ with units of ____.

The symbol for ***Planck's Constant*** is _______________. ***Planck's Constant*** is equal to ______________________________.

The speed and energy equations may be combined into ***de Broglie's Equation***:

Relationships:

What type of relationship exists between frequency and wavelength?

What type of relationship exists between frequency and energy of a photon?

What type of relationship exists between wavelength and energy of a photon?

Laws of classical physics – there is no limit to how large or how small the energy gained or lost by an object may be.
So according to this, the bright-line spectrum should be a continuous rainbow. But it is NOT.
We cannot always apply macroscopic laws to subatomic events.

Max Planck – German physicist – 1858-1947
Question: Why does an object change color when heated? Heating iron causes it to change from black to yellow to red to white to blue as its temperature is increased.
Answer: The energy of a body changes only in small discrete units – ***quanta*** – small packages of energy.

Page 376 – It appears that thermal energy may continuously supplied to heat liquid water to any temperature between 0^o and 100^oC. Actually the water temperature increases by infinitesimally small steps, which occurs as individual water molecules absorb quanta of energy. An ordinary thermometer is unable to detect small changes in temperature. Thus your everyday experiences give you no clue to the fact that energy is quantized.

Remember, the energy of a photon of light, heat, or other radiation is calculated by $E = hv$.

<u>The Photoelectric Effect</u>

- When light shines on metals they emit electrons (more specifically called ***photoelectrons***). The alkali metals are particularly sensitive to the effect.
- The light has to have a high enough frequency – in other words have enough energy.
- Potassium
  - Red light – $v = 4.3\text{x}10^{14}$ Hz to $4.6\text{x}10^{14}$Hz – will not cause an electron to eject from potassium – no matter how much of the light gets used.
  - Yellow light – $v = 5.1\text{x}10^{14}$Hz to $5.2\text{x}10^{14}$Hz – will start the process of ejecting electrons from potassium – even if it is very weak in intensity.
- Not enough energy – no photoelectrons ejected
- Just above threshold energy – electron ejected with minimal energy/speed
- Well above threshold energy – electron ejected with significant energy/speed
- Greater intensity of light – more waves of a given frequency – more electrons ejected
- The ***number of electrons ejected*** depends on the number of light waves – the intensity.
- The ***energy of the ejected electrons*** depends on the frequency (or energy) of the light waves.

<u>Heisenberg Uncertainty Principle</u>

It is ***impossible*** to know both the position and the velocity of a particle at the same time. The more you find out about one of the pieces of information, the less you are able to determine about the other. You need to use information about the position at an instant to find information about its speed – but by the time you look at that position, it is already gone and you know nothing about its new position – unless you use information about its speed – but then you will not know anything about its new velocity.

This uncertainty is more obvious and significant with small objects than with large objects. Comparing an electron and a baseball – there is the uncertainty with the baseball – but it is so small it is nearly immeasurable. But, with regard to an electron, the uncertainty is much more significant due to the extremely small size of an electron.

Chapter 13 – Electrons in Atoms
Part 6 – Assignment: Energy, Photons, Bright-Line Spectrum, and the Photoelectric Effect

Solve the following problems. Show all set-ups, work, units, etc. Solve problems 1 – 6 on a separate sheet of paper.

1. A photon from a source of electromagnetic radiation transmits 4.58×10^{-19} J of energy to matter. What is the frequency of this electromagnetic radiation? What is the wavelength of the electromagnetic radiation?

2. What is the wavelength of electromagnetic radiation if a photon transmits 2.04×10^{-18} J of energy to matter?

3. Calculate the quantum energy from the electromagnetic radiation with a frequency of 3.20×10^{21} Hz.

4. Calculate the quantum energy from electromagnetic radiation with a wavelength of 4.20×10^{-7} m.

5. What is the energy of the radio wave from station WKTU – 92.3 FM?

6. A radio wave has a wavelength of 2.921 m. What is the energy of the radio station signal?

7. Electromagnetic radiation exhibits properties of both __________________ & __________________.

8. A "packet" or "unit" of electromagnetic radiation is called a ____________________.

9. Different __________________ of light carry different amounts of energy per photon.

10. An atom that possess excess energy is said to be in a(n) ____________________ state.

11. An atom may release its excess energy by emitting a(n) ____________________ of electromagnetic energy.

12. A beam of red light has ***higher*** or ***lower*** energy photons than blue light. (Choose one.)

13. Because a given element's atom emit only certain photons of light, we know only certain _______________ _______________ are occurring in those particular atoms.

14. The energy of an emitted photon corresponds to the difference in energy between the different ____________ _______________ of the atom.

15. How do we know that the energy levels of the hydrogen are not continuous, as physicists originally assumed?

16. Explain the origin of the atomic emission spectrum of an element.

17. Can classical physics explain the photoelectric effect? Explain.

18. Compare the ground state and the excited state of an electron.

19. What will happen if the following occur?
    a. Monochromatic light shining on the alkali metal cesium is just above the threshold frequency.
    b. The intensity of light increases but the frequency remains the same.
    c. Monochromatic light of a shorter wavelength is used.

20. What will happen when a hydrogen atom absorbs a quantum of energy?

Chapter 13 – Electrons in Atoms
Part 7 - Notes: Quantum Numbers

Objectives: Identify, define, and explain: quantum number, principal quantum number, angular momentum quantum number, magnetic quantum number, and spin quantum number.
Use three quantum numbers to identify a given atomic orbital and four quantum numbers to identify an electron.
Explain why no two electrons in a given atom can have the same four quantum numbers.
Give a brief description of what each of the four quantum numbers does and how it functions to denote a portion of an electron's address in an atom.

Text Reference: Page 364

Scientists have been unable to determine the exact position of an electron in an atom. Electrons in an atom may move to higher or lower energy levels by absorbing or releasing energy, in units or quanta of energy. Also, electrons at different energy levels are believed to travel in certain regions called orbitals. These orbitals are a probability map for the location of an electron.

The exact location of an electron cannot be pinpointed; however, the electron may be described using ***quantum numbers***. There are four quantum numbers. The numbers represent an electron's address within an atom. ***No two electrons in the same atom can have the same address at the same time***.

The Principal Quantum Number = n

The first quantum number is called the principal quantum number. It describes the **energy level** of an electron. The size of orbitals increases as the energy level increases; so it may be said that the first quantum number is related to the size of the orbital. The letter ***n*** is the abbreviation for the principal quantum number.

Energy levels range from n = 1 to n = infinity. The value of ***n*** must be a whole number from one to infinity.

The Angular Momentum Quantum Number = l

The second quantum number is called the angular momentum quantum number and it describes the **shape** of the orbital in which the electron has the highest probability of being found. The letter ***l*** is the abbreviation for the angular momentum quantum number.

Orbitals may have the shape of a sphere. Sphere-shaped orbitals are called s-orbitals and are designated by a numerical value of ***l*** = 0. Orbitals may have two-lobe shapes and be called p-orbitals. P-orbitals are designated by a numerical value of ***l*** = 1. There are also "four-lobed" orbitals called d-orbitals. These orbitals are designated numerically by ***l*** = 2. There are more complex orbitals with unusual shaped called f-orbitals. F-orbitals have a numerical designation of ***l*** = 3. Other orbitals theoretically exist and are designated by numbers ***l*** = 4, 5, 6, …, to a maximum value of ***n*** – 1.

So, when an electron has a principal quantum number of ***n***, it may have an angular momentum quantum number that is any whole number from 0 through ***n*** – 1.

The Magnetic Quantum Number = m

The third quantum number is called the magnetic quantum number and it describes the **orientation of the orbitals** around the x, y, or z-axis. The abbreviation for the magnetic quantum number is ***m***. Values of ***m*** range from –***l*** to +***l*** and all whole number integers in between.

The first three quantum numbers make up 3/4 of an electron's address but they also define the orbital in which the electron has the greatest probability of being found. **No two orbitals can have the same first three quantum numbers** otherwise it allows for two orbitals to be in the same space and that can't happen.

Spin Quantum Number = m_s

The fourth quantum number is called the spin quantum number and it indicates the **spin** of the electron being described. The fourth quantum number is abbreviated m_s.

Electrons are believed to spin about their own axes. When two electrons occupy the same orbital they are said to spin in opposite directions. By convention, one electron is arbitrarily assigned a spin of +1/2 and the other is assigned a spin of –1/2.

NOTE: The energy of an electron is defined by its four quantum numbers. **No two electrons in the same atom may have the same set of four quantum numbers at the same time.** This is the *Pauli Exclusion Principle*. It basically says that no two electrons can be in the same space at the same time.

Chapter 13 – Electrons in Atoms
Part 7 - Assignment: Quantum Numbers

1. TRUE or FALSE: Scientists are able to determine the exact position of electrons in atoms.
2. An electron moves to a higher energy level by _____________ energy and to a lower level by _____________ energy.
3. Energy released by electrons is released in the form of __________________.
4. An electron's "address" is described by using ________________ _______________.
5. The energy level of an electron is denoted by the ____________________ quantum number.
6. The spin of an electron is denoted by the ____________________ quantum number.
7. The orientation of the orbital is denoted by the ____________________ quantum number.
8. The shape of the orbital is denoted by the ____________________ quantum number.
9. State what the following letters represent:
   a. p ______________________________
   b. ***m*** ______________________________
   c. s ______________________________
   d. $\boldsymbol{m_s}$ ______________________________
   e. y ______________________________
   f. ***n*** ______________________________
   g. d ______________________________
   h. ***l*** ______________________________
   i. f ______________________________
10. An electron may spin two ways. Each direction is arbitrarily assigned a spin of __________ or __________.
11. An orbital may be defined using __________ of the hour quantum numbers. They are _________________.
12. TRUE or FALSE: Two electrons in an atom may have the same four quantum numbers.
13. You know the answer to number 12 because of the _______________ _______________ _______________.
14. Two electrons in an atom can have the same __________, __________, and __________ values. It means that they are in the same orbital.