CHAPTER 6 – Electronic Structure of Atoms

Section 6.1 – The Wave Nature of Light

(a) The speed of light (in a vacuum) is equal to ____________________

(b) The distance between two adjacent peaks (or troughs) is called the _____________________

(c) The number of complete cycles that pass a given point each second is called the

____

(d) Write Equation 6.1 on page 208. Identify what each symbol represents. Indicate the units that are normally used to express each quantity.

(e) Various forms of electromagnetic radiation are listed in the table below. Fill in the missing information.

Form of EMR	wavelength, in units of nanometers	wavelength, in units of meters	frequency, in units of Hz
radio waves	3×10^{9} nm		1×10^{8} Hz
microwaves		3×10^{-2} m	
infrared light		3×10^{-5} m	
red light			4.6×10^{14} Hz
yellow light	580 nm		
violet light	400 nm		
ultraviolet light		2×10^{-7} m	
X-rays	1 nm		3×10^{17} Hz
gamma rays		1×10^{-12} m	

Section 6.2 – Quantized Energy and Photons

(a) A quantum is defined as the ____________________ quantity of energy that can be emitted or absorbed as electromagnetic radiation.

(b) Write Equation 6.2 on page 210. Identify what each symbol represents. Indicate the units that are normally used to express each quantity.

(c) Because energy can be released only in specific amounts, we say that the allowed energies are ____________________. This means that their values are restricted to certain quantities.

(d) Fill in the missing information in the following table.

Form of EMR	frequency, in units of Hz	energy, in units of J
radio waves		6.6×10^{-26} J
microwaves	1×10^{10} Hz	
infrared light	1×10^{13} Hz	
red light		3.0×10^{-19} J
yellow light	5.2×10^{14} Hz	
violet light	7.5×10^{14} Hz	
ultraviolet light		9.9×10^{-19} J
X-rays	3×10^{17} Hz	
gamma rays	3×10^{20} Hz	

(e) When light of a certain minimum frequency is shined on a clean metal surface, the light can cause the atoms at the surface of the metal to lose electrons. This phenomenon is known as the ____

(f) Einstein explained this phenomenon by describing the radiant energy as a stream of tiny energy packets. Each packet, or “particle” of energy is called a __________________.

(g) Our bodies are penetrated by X-rays, but not by visible light. This can be explained because X-rays have ______ than visible light. (Check all that apply.)

_______ a faster speed _______ a shorter wavelength _______ greater energy

_______ a slower speed _______ a higher frequency _______ lower energy

_______ a longer wavelength _______ a lower frequency

(h) A photon of light has a wavelength of 400 nm. Calculate the energy of this photon, in J.

(i) Your answer from part (h) represents the energy of a single photon. Convert this energy into units of kJ/mol.

(j) The ionization energy of sodium is 496 kJ/mol. Would light with a frequency of 400 nm be sufficient to cause sodium to lose its electrons? Justify your answer with a calculation.

(k) Assuming that you said “no” in part (j), calculate the minimum wavelength of light (in units of nm) that would be needed to cause a sodium atom on the surface of a sample of sodium metal to lose an electron.

Section 6.3 – Line Spectra and the Bohr Model

(a) When narrow beam of white light is passed through a prism, a ________________________ spectrum is produced. This is a complete rainbow of colors, from red to violet.

(b) When a high voltage is applied to a glass tube containing hydrogen gas, a pink light is emitted. When this light is passed through a prism, only a few wavelengths are present in the spectrum.

This type of spectrum is called a __________________ spectrum.

(c) When scientists first detected the spectrum for hydrogen in the mid-1800s, it contained only four lines in the visible portion of the electromagnetic radiation spectrum. Calculate the frequency and energy associated with each line.

Color	Wavelength	Frequency (Hz)	Energy (J)
red	656 nm or 6.56×10^{-7} m		
blue-green	486 nm or 4.86×10^{-7} m		
blue	434 nm or 4.34×10^{-7} m		
violet	410 nm or 4.10×10^{-7} m		

(d) To explain the line spectrum of hydrogen, Neils Bohr assumed that electrons in hydrogen

atoms move in ____.

(e) Bohr based his model of the hydrogen atom on three postulates:

- Only orbits of ____

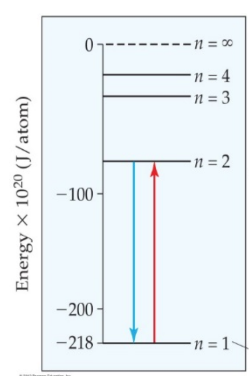
  are permitted for the electron in a hydrogen atom.

- An electron in a permitted orbit is in an ____
  It does not spiral into the nucleus.

- As the electron changes from one energy state to another, energy is either

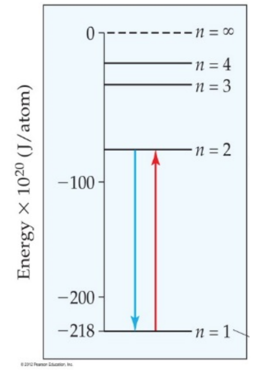
  ________________________ or ________________________.

(f) Based on the diagram at right, should an electron that travels from n = 2 to n = 1 absorb energy or release energy? Explain.



(g) The integer *n*, which can have whole number values of

1, 2, 3, … ∞ is called the ______________________________



(h) As *n* increases, the radius of the orbit gets __________________

(i) The lowest-energy state (*n* = 1) is called the ________________

state. When an electron is in a higher-energy state (*n* = 2 or higher)

the atom is said to be in ______________________________

(j) When an electron moves from n = 3 to n = 2, red light is emitted. When an electron moves from n = 2 to n = 1, a certain form of light is emitted. Is this light in the infrared region or in the ultraviolet region? Justify your answer.

(k) The Bohr model of the atom has its limitations. It cannot predict the line spectra for other atoms besides hydrogen, and it does not account for the wavelike properties of electrons. However, two important ideas that were introduced by the Bohr model are still incorporated into our current model of the atom:

Electrons exist only in ________________________________, which are described by quantum numbers.

Energy is involved in the ____

____

Section 6.4 – The Wave Behavior of Matter

Although the topics described in this section are interesting and relevant for a general chemistry course, they are not emphasized in the AP Chemistry curriculum.

Section 6.5 – Quantum Mechanics and Atomic Orbitals

Schrodinger's wave functions help us to map out a region in space in which an electron can be found. We do not know the exact location of an electron. Instead, we can know the probability of finding an electron in a certain area around the nucleus. The electron density can be mapped, as shown in Figure 6.16 on page 220.

(a) The solution to Schrodinger's equation for the hydrogen atom yields a set of wave functions

called ______________________

(b) The lowest-energy orbital in the hydrogen atom is called the _______ orbital and has a

________________________ shape.

(c) The principal quantum number (*n*) can have positive integral values of 1, 2, 3, … ∞

As the value of *n* increases,

the orbital becomes ____________________ in size, and

the electron has _______________________ energy, and

the electron becomes ___________________ to remove from the atom.

(d) Match the angular momentum quantum number (ℓ) to the letter used to describe the orbital:

ℓ	0	1	2	3
letter used				

(e) Fill in the table below with the name of each orbital as described by the first two quantum numbers. For example: **2p** or **3d**.

n	1	2		3			4			
ℓ	0	0	1	0	1	2	0	1	2	3
orbital										

(f) The collection of orbitals with the same value of *n* is called an electron ________________.

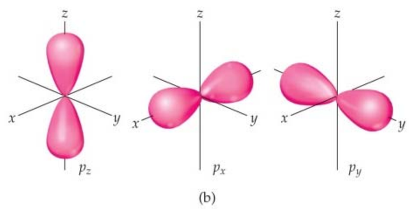
(g) The set of orbitals that have the same *n* and ℓ values is called a _____________________.

(h) The shell with principal quantum number *n* consists of exactly ______ subshells.

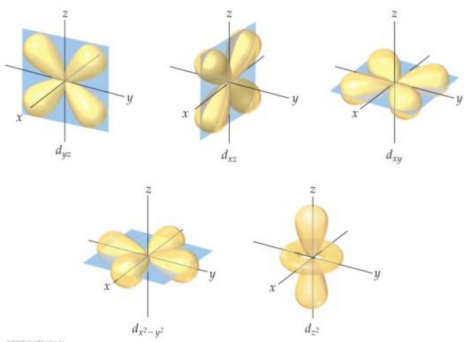
Section 6.6 – Representations of Orbitals

You should know that s-orbitals have a spherical shape.

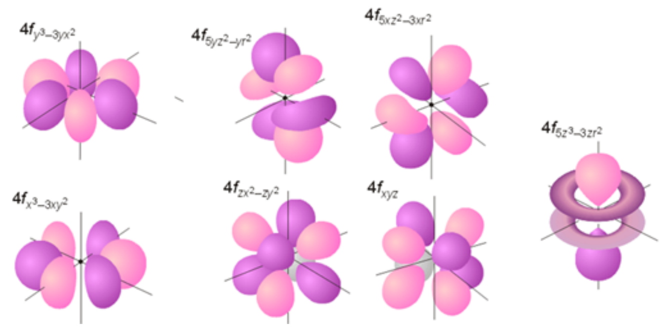
p-orbitals look like this:



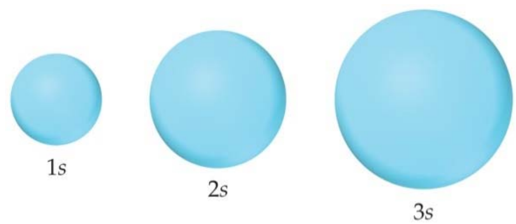
d-orbitals look like this:



f-orbitals look like this:



The only other point about orbital shapes and sizes that is necessary to mention here is that **as the value of n increases, the orbitals get larger in size.**
For example, a 3s orbital is larger than a 2s orbital, which is larger than a 1s orbital.



Section 6.7 – Many-Electron Atoms

In a many-electron atom, the electron–electron repulsions cause the various subshells in a given shell to be at different energies.

(a) In a many-electron atom, for a given value of *n*, the energy of an orbital ________________ with increasing value of ℓ.

(b) All orbitals of a given subshell have _____________ energy. The term for such orbitals is **degenerate.**

(c) According to the Pauli exclusion principle, an orbital can hold a maximum of _______

electrons, and they must have ____________________________________.

Section 6.8 – Electron Configurations

(a) Define the term electron configuration.

(b) The most stable electron configuration is called the ___________________________________

(c) What does it mean when we say that electrons are "paired" in the same orbital?

(d) Write the orbital diagram for nitrogen, showing that the arrangement of the electrons obeys Hund's Rule:

☐	☐	☐☐☐
1s	2s	2p

(e) How many valence electrons does a nitrogen atom have? ____________

How many unpaired electrons does a nitrogen atom have? ____________

(f) You should be able to write the complete electron configurations (and the noble gas abbreviated configurations) for the first 54 elements on the periodic table, from hydrogen to xenon. Refer to Figure 6.31 on page 236 for assistance. You do not have to worry about any exceptions to the aufbau rule (such as Cr or Cu).

Element	Complete Electron Configuration	Noble Gas Abbreviated Configuration	Valence Electrons
O			
Mg			
Si			
Cl			
K			
Ga			
As			

Section 6.9 – Electron Configurations and the Periodic Table

A simple way to figure out the electron configuration of an element is to be familiar with the blocks of the periodic table. Fill in the missing subshells in the table below.