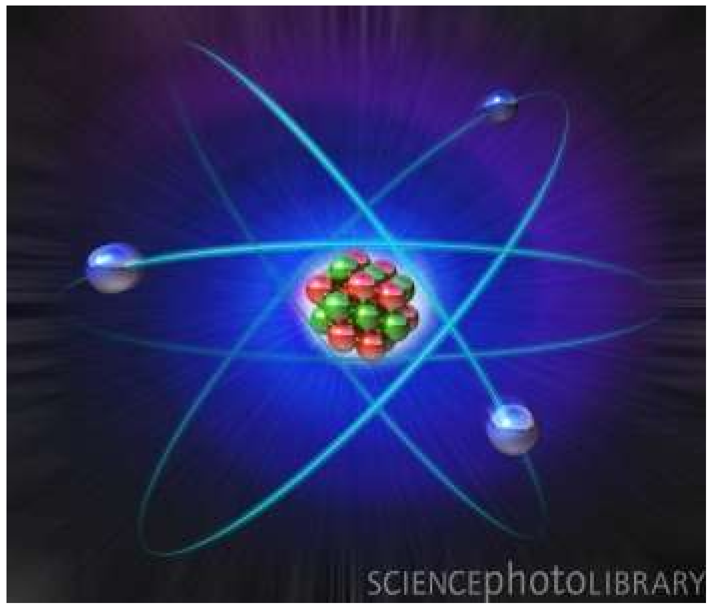
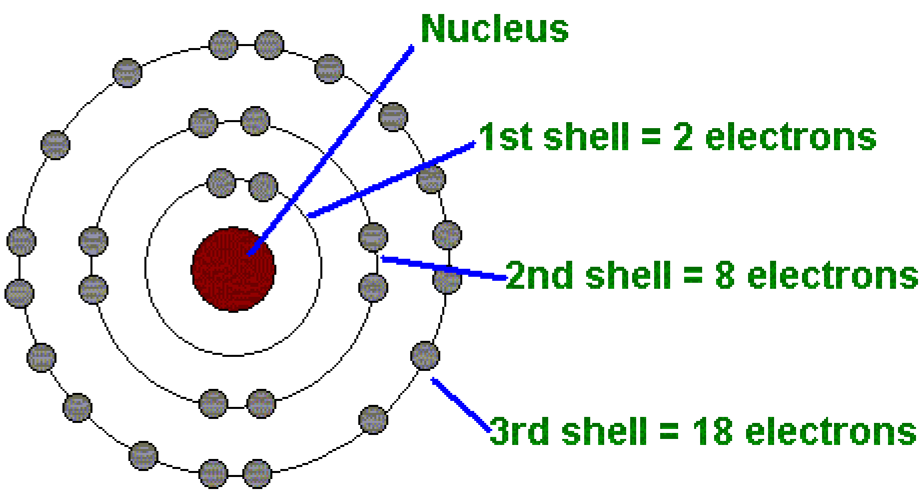
AP Chapter 6 – Electronic Structure of Atoms

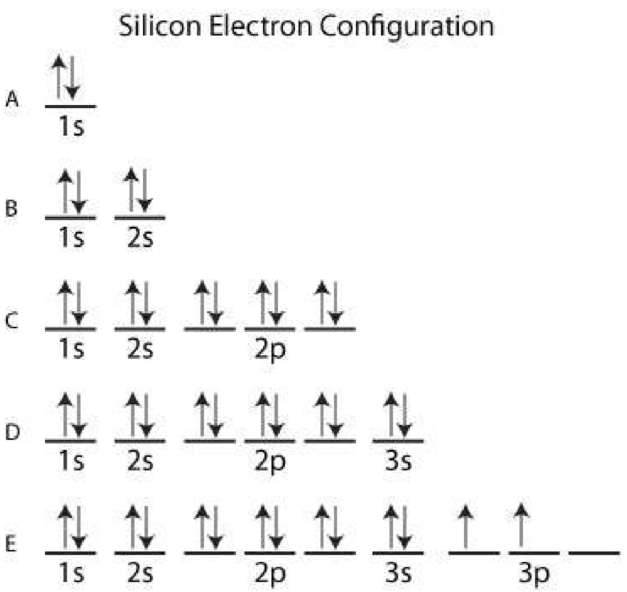


Jennie L. Borders

Section 6.1 – The Wave Nature of Light

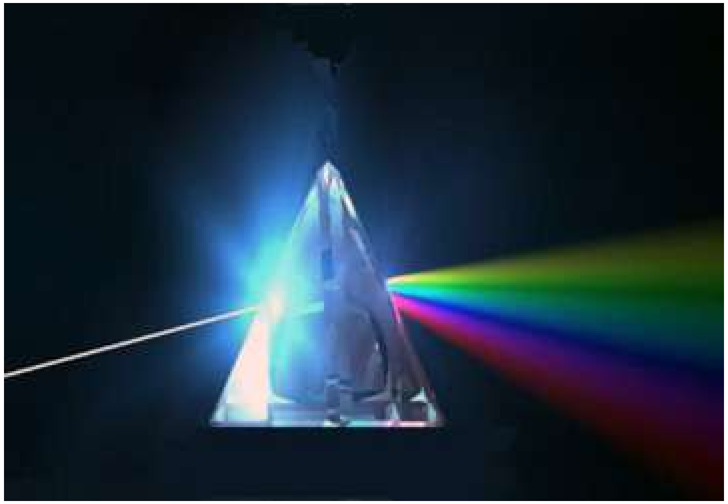
- The electronic structure of an atom refers to the number of electrons in an atom as well as distribution of the electrons around the nucleus and their energies.

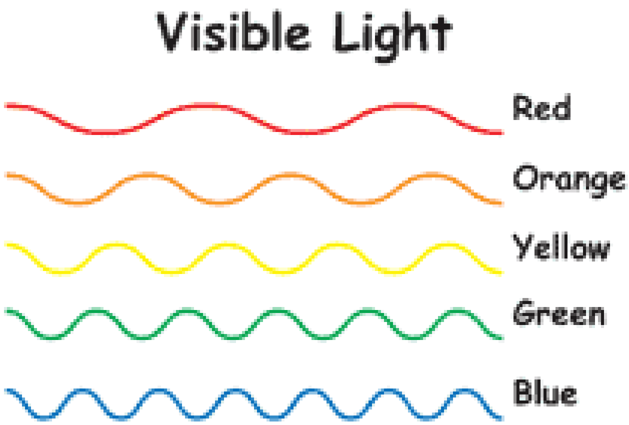




Light

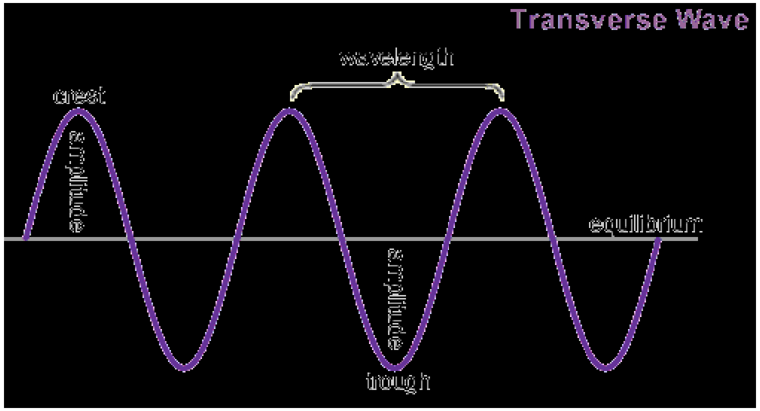
- Visible light is an example of electromagnetic radiation.
- All types of electromagnetic radiation move through a vacuum at a speed of 3.00×10^8 m/s, the speed of light.

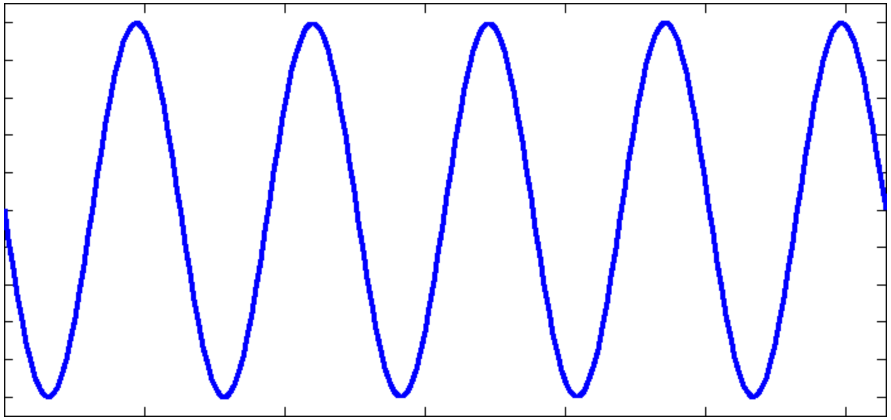




Waves

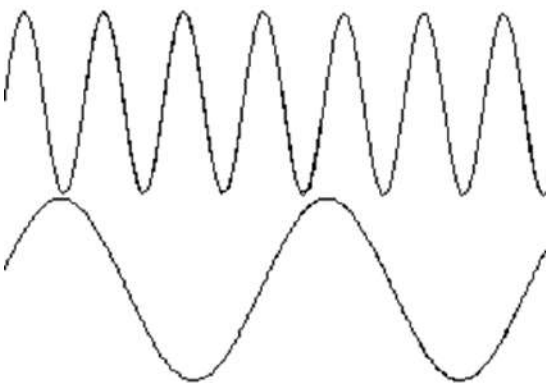
- The distance between two adjacent peaks (or two troughs) is the wavelength.
- The number of complete wavelengths that pass a given point each second is the frequency of the wave.



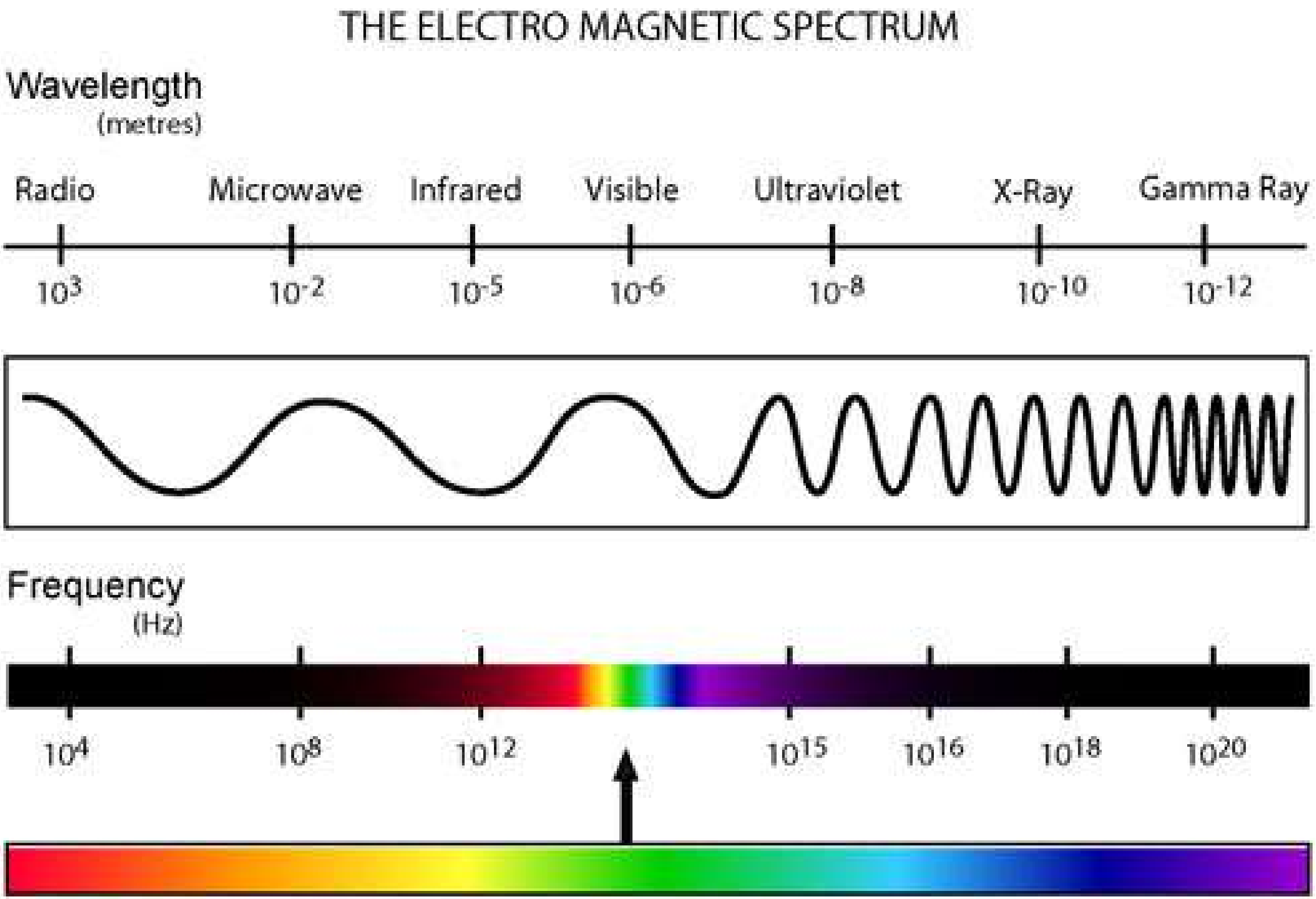


Electromagnetic Waves

- All electromagnetic radiation moves at the same speed, namely the speed of light.
- If the wavelength is long, there will be fewer waves passing a certain point, so frequency will be low.
- If the wavelength is short, then the frequency will be high.



Electromagnetic Spectrum



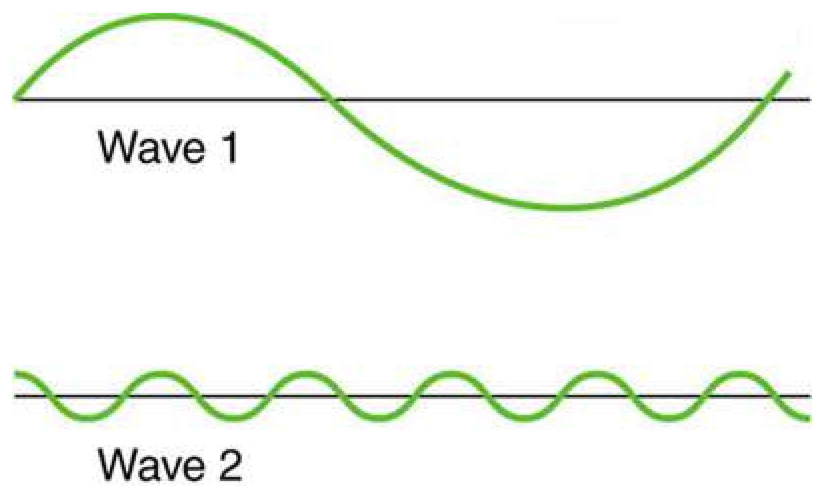
Formula

$$c = \lambda\nu$$

- c = speed of light (3.00×10^8 m/s)
- λ = wavelength (m)
- ν = frequency (Hz or s^{-1})

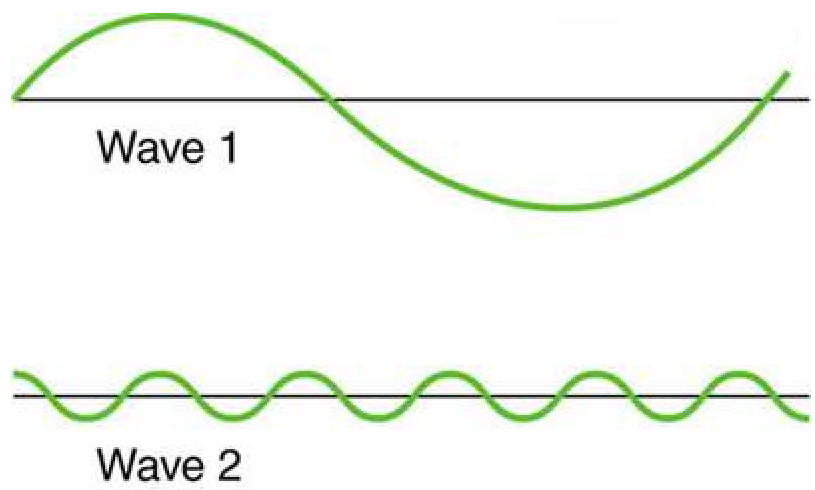
Sample Exercise 6.1

- Which wave has the higher frequency? If one wave represents visible light and the other represents infrared radiation, which wave is which?



Practice Exercise

- If one of the waves represents blue light and the other red light, which is which?



Sample Exercise 6.2

- The yellow light given off by a sodium vapor lamp used for public lighting has a wavelength of 589nm. What is the frequency of this radiation?

Practice Exercise

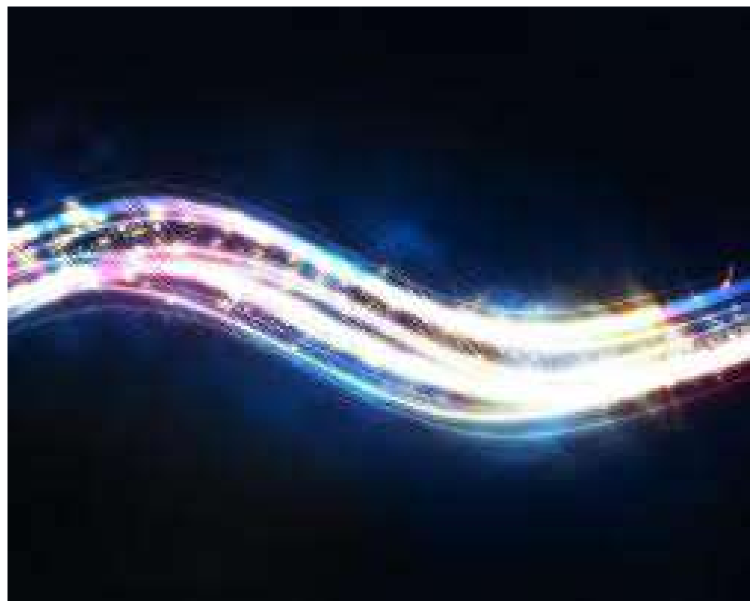
- A laser used in eye surgery to fused detached retinas produces radiation with a wavelength of 640nm. Calculate the frequency of this radiation.

Practice Exercise

- An FM radio station broadcasts electromagnetic radiation at a frequency of 103.4 MHz. Calculate the wavelength of this radiation.

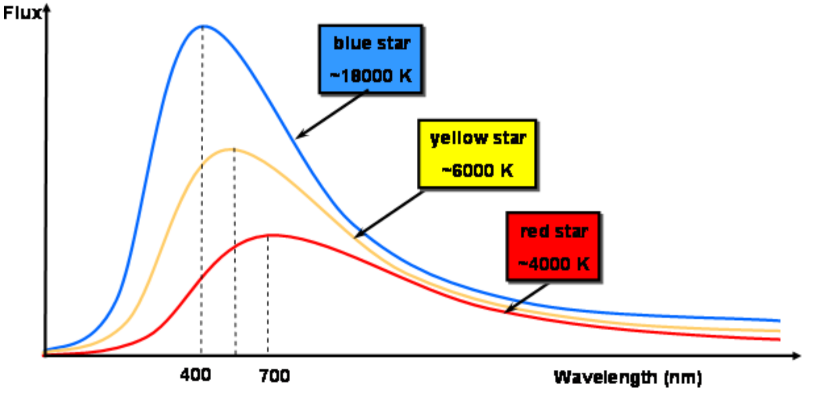
Section 6.2 – Quantized Energy and Photons

- There are three main phenomena that cannot be explained by the wave theory of light:
  1. Blackbody radiation
  2. Photoelectric effect
  3. Emission spectra



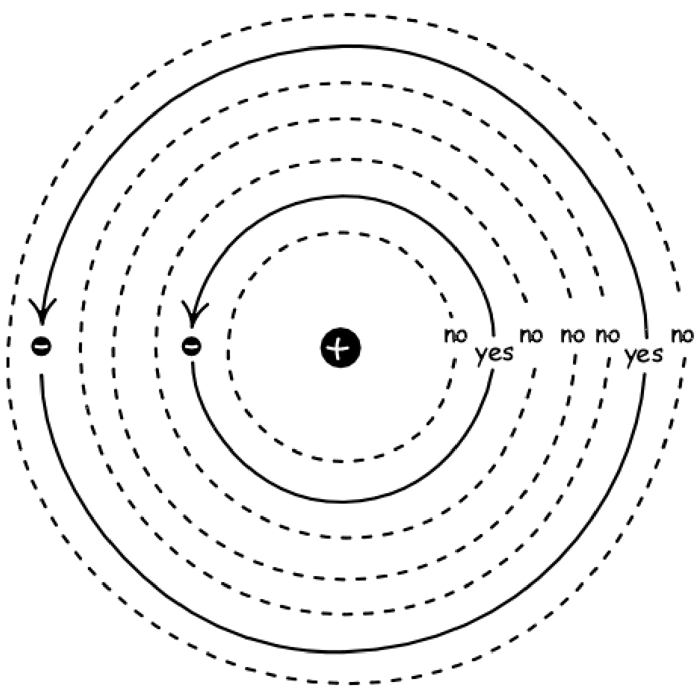
Blackbody Radiation

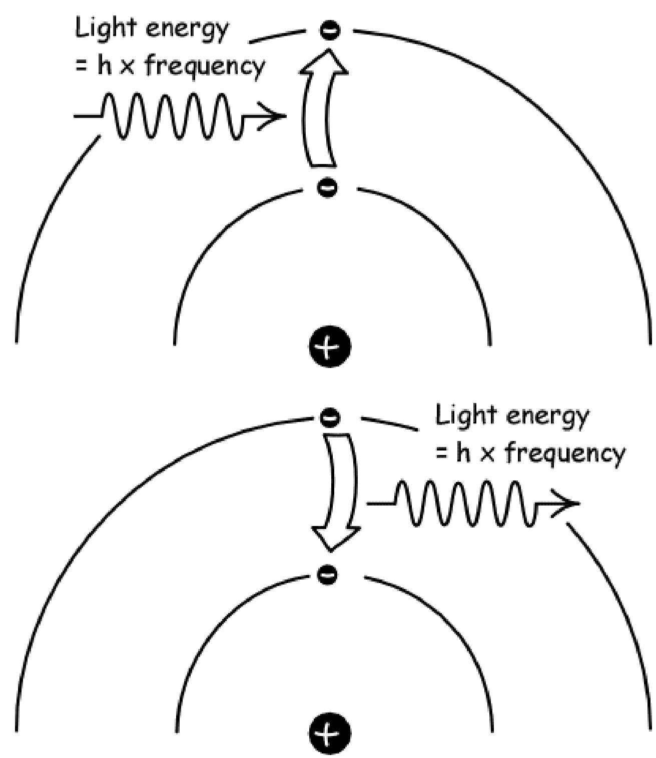
- When solids are heated, they emit radiation. As seen in the red glow of heated metal.
- The wavelength emitted depends on the temperature of the object.



Quantum

- Planck assumed that energy could only be released or absorbed in certain fixed quantities that he called a quantum.





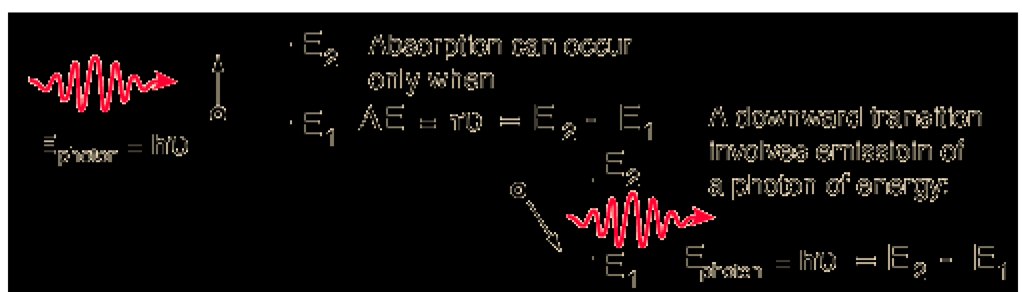
Formula

$$E = hv$$

- E = energy (J)
- h = Planck's constant (6.626×10^{-34} J·s)
- v = frequency (Hz or s^{-1})

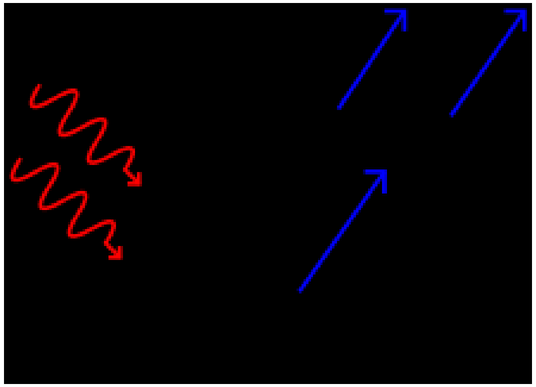
Planck's Theory

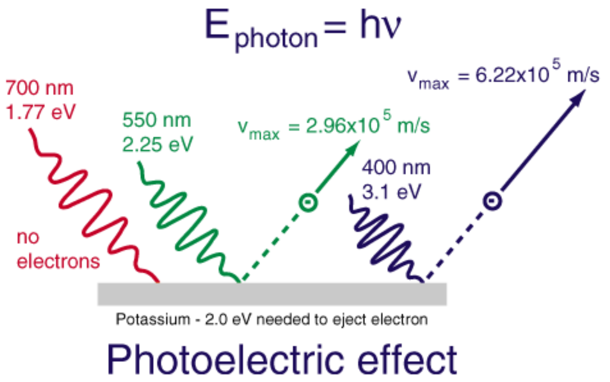
- According to Planck's theory, matter is allowed to emit or absorb energy only in whole-number ratios of hv.
- Since hv is a very small number, a quantum of energy is a very small amount.
- Because energy can be released only in specific amounts, we say that energy is quantized.



Photoelectric Effect

- The photoelectric effect occurs when light of a certain frequency shining on the surface of a metal causes electrons to be emitted from the metal.
- In this case, light behaves like a stream of tiny energy packets called photons.





Sample Exercise 6.3

- Calculate the energy of one photon or yellow light with a wavelength of 589nm.

Practice Exercise

- A laser emits light with a frequency of $4.69 \times 10^{14}\ s^{-1}$. What is the energy of one photon of the radiation from this laser?

Practice Exercise

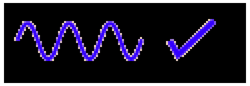
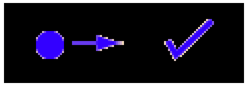
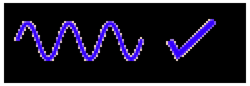
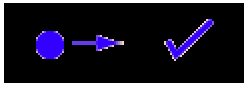
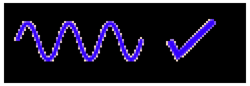
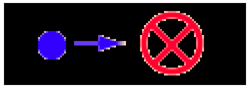
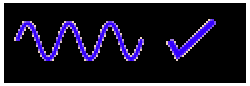
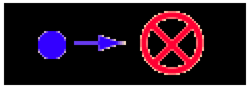
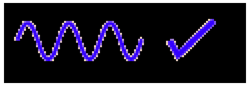
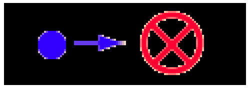
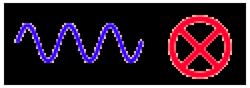
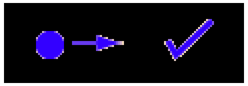
- If the laser emits a pulse of energy containing 5.0×10^{17} photons of this radiation, what is the total energy of that pulse?

Practice Exercise

- If the laser emits 1.3×10^{-2}J of energy during a pulse, how many photons are emitted during this pulse?

Wave vs. Particle

- We must consider that light possesses both wavelike and particle-like characteristics.

Phenomenon	Can be explained by waves.	Can be explained by particles.
Reflection		
Refraction		
Interference		
Diffraction		
Polarization		
Photoelectric effect		

Section 6.3 – Line Spectra and the Bohr Model

- A spectrum is produced when radiation from a source is separated into its different wavelengths.
- Spectrums from light containing all wavelengths is called a continuous spectrum.
- A spectrum containing radiation of only specific wavelengths is called a line spectrum.

(a)

(b)

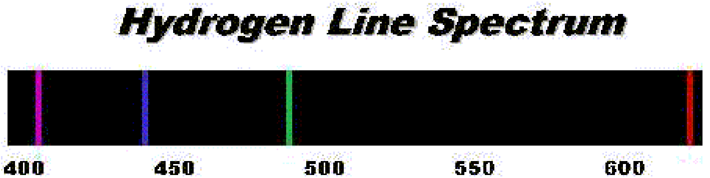
Rydberg Equation

- The Rydberg equation was used to calculate the line spectrum wavelengths for hydrogen.

$$\frac{1}{\lambda} = (R_H)\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

R_H = Rydberg constant ($1.096776 \times 10^7\ m^{-1}$)

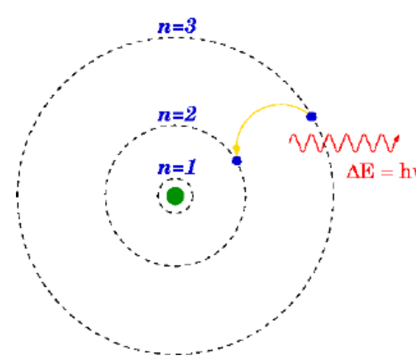
n_1 and n_2 are positive integers that represent energy levels.



Bohr's Model

- Bohr based his model on 3 postulates:

1. Only orbits of certain radii (definite energies) are permitted for the electron.
2. An electron in an orbit has a specific energy and will not spiral into the nucleus.
3. Energy emitted or absorbed by the electron can only change in multiples of $E = hv$.



Energy States of Hydrogen

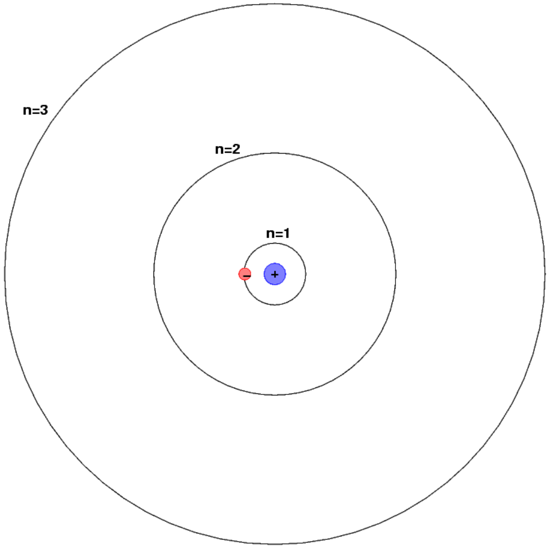
- Bohr calculated the specific energy levels using the equation below.

$$E = (-hcR_H)\left(\frac{1}{n^2}\right)$$

E = energy, h = Planck's constant, c = speed of light, R_H = Rydberg constant, and n = positive integer that represents an energy level.

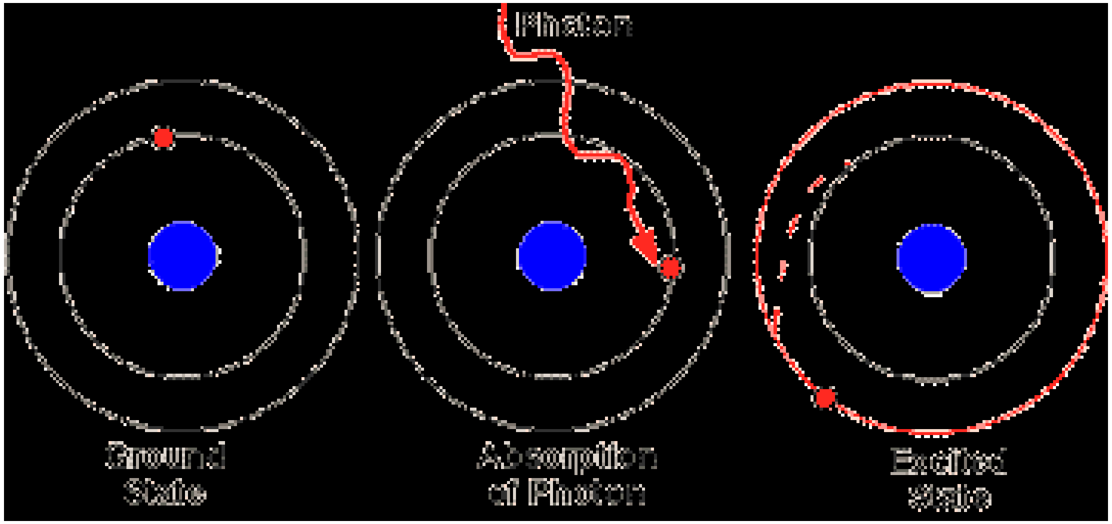
Energy States of Hydrogen

- The energies of the electron of the hydrogen atom are negative for all values of n.
- The more negative the energy is, the more stable the atom will be.
- The energy is most negative for n = 1.
- As n gets larger, the energy becomes less negative.



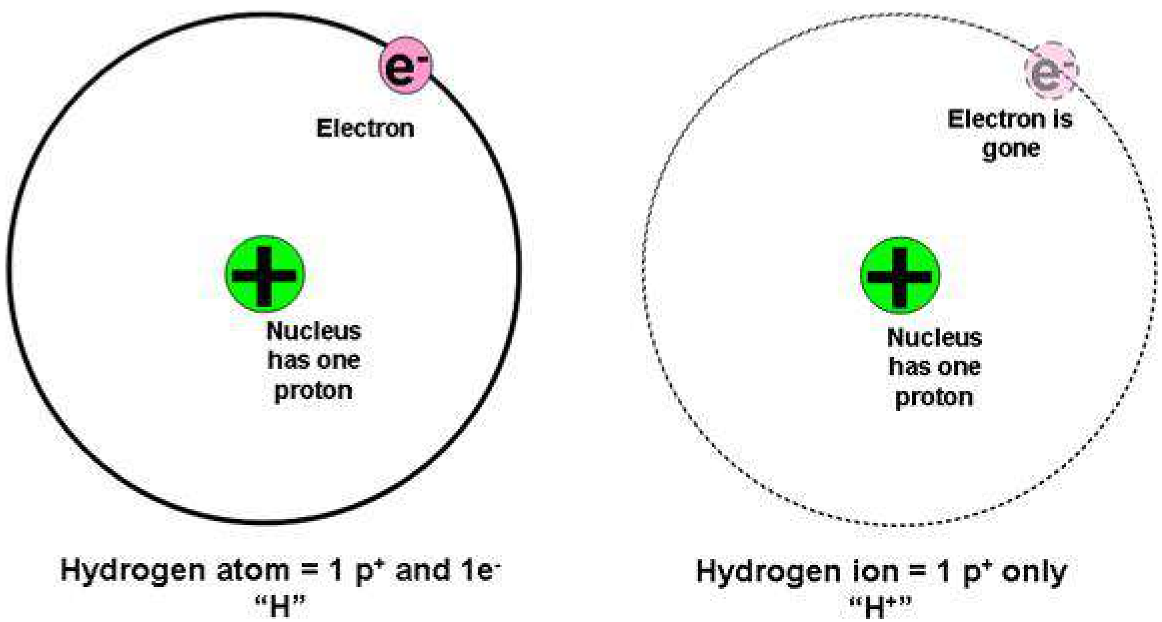
Energy States of Hydrogen

- The lowest energy state (n = 1) is called the ground state of the atom.
- When the electron is in a higher energy level, the atom is said to be in an excited state.



Energy States of Hydrogen

- When n = ∞, the energy approaches zero.
- Thus, the electron is removed from the nucleus of the hydrogen atom.



Changing Energy Levels

- Energy must be absorbed for an electron to move to a higher energy level.
- Radiant energy is emitted when the electron jumps to a lower energy level.

$$E = hv \qquad c = \lambda v$$

so

$$E = \frac{hc}{\lambda}$$

Sample Exercise 6.4

- Using Figure 6.14 in your book (p. 220), predict which of the following electronic transitions produces the spectral line having the longest wavelength: n = 2 to n =1, n = 3 to n = 2, or n = 4 to n = 3.

Practice Exercise

- Indicate whether each of the following electronic transitions emits energy or requires the absorption of energy:

a. $n = 3$ to $n = 1$

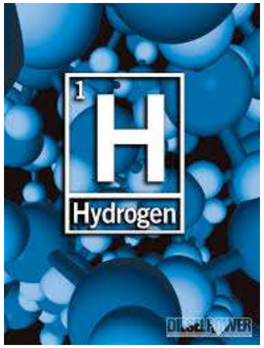
b. $n = 2$ to $n = 4$

Limitations of the Bohr Model

- Bohr's model can only explain the line spectrum of hydrogen, not other elements.

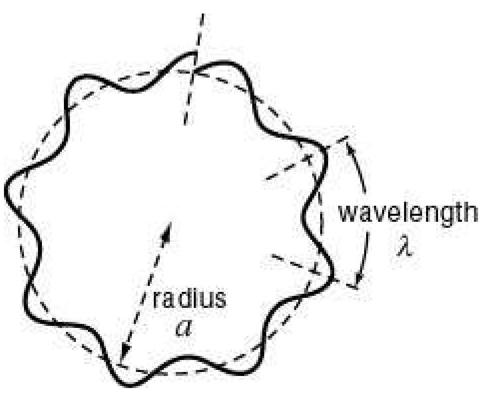
- Important ideas from Bohr's model:

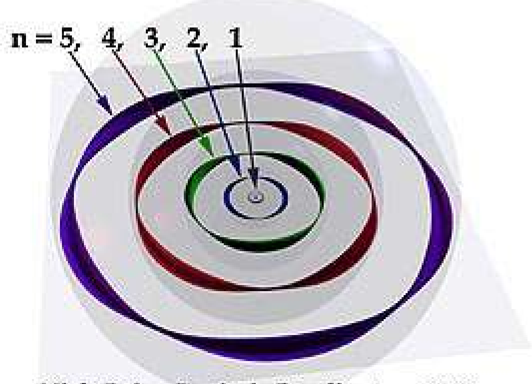
1. Electrons exist in specific energy levels.
2. Energy is involved when electrons jump to different energy levels.



Section 6.4 – The Wave Behavior of Matter

- De Broglie suggested the as the electron moves around the nucleus, it is associated with a particular wavelength.
- De Broglie used the term matter waves to describe the wave characteristics of material particles.





Niels Bohr - Louis de Broglie atom, 1924

De Broglie

- He proposed that the wavelength of any particle depends on its mass and velocity.

$$\lambda = \frac{h}{mv}$$

λ = wavelength, h = Planck's constant, m = mass, and v = velocity.

Based on his formula, the wavelength of an object of normal size would be so tiny that it could not be observed.

Sample Exercise 6.5

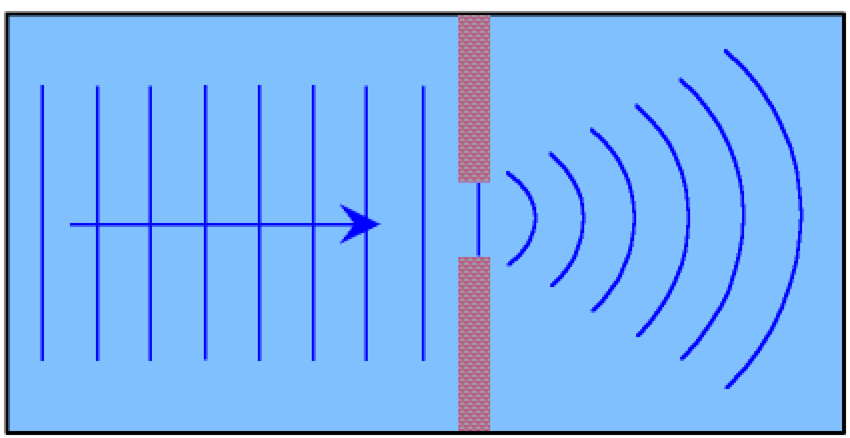
- What is the wavelength of an electron moving with a speed of 5.97×10^{6} m/s? The mass of the electron is 9.11×10^{-31} kg.

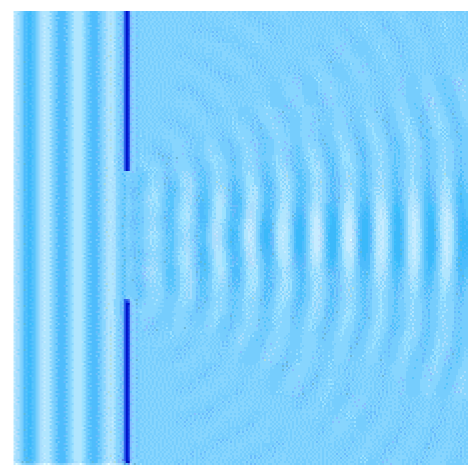
Practice Exercise

- Calculate the velocity of a neutron whose de Broglie wavelength is 500 pm. The mass of a neutron is 1.67×10^{-27} kg.

Diffraction

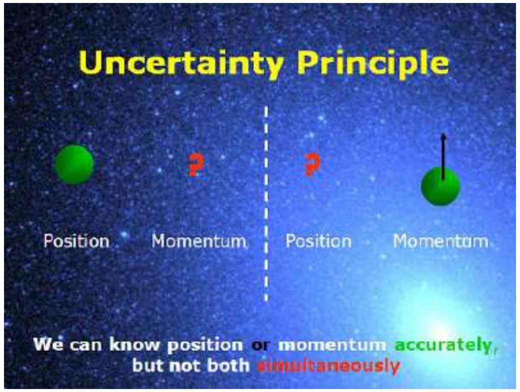
- A few years after de Broglie published his theory, it was found that electrons (like x-rays) will diffract when passed through a crystal.
- Diffraction is normally associated with waves.





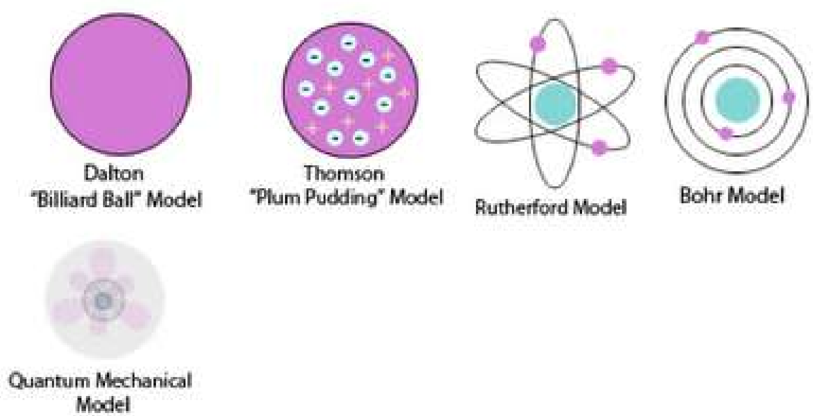
The Uncertainty Principle

- Heisenberg's uncertainty principle states that when an object has an extremely small mass (electron), it is impossible for us to simultaneously know both the exact momentum and location.
- If the object is normal sized, then the uncertainty would be too small to matter.



Section 6.5 – Quantum Mechanics and Atomic Orbitals

- The Bohr model assumes that the electron is in a circular orbit of a particular radius.
- In the quantum mechanical model, the electron's location is based on the probability that the electron will be in a certain region of space.



Quantum Numbers

quantum number	symbol	represents	possible numbers
principle quantum number	n	energy level	1, 2, 3, 4, 5, 6, 7
angular momentum quantum number	l	sublevel	0, 1, 2, 3
magnetic quantum number	m_l	orbital	-3, -2, -1, 0, 1, 2, 3
spin magnetic quantum number	m_s	spin	+1/2 or -1/2

Sample Exercise 6.6

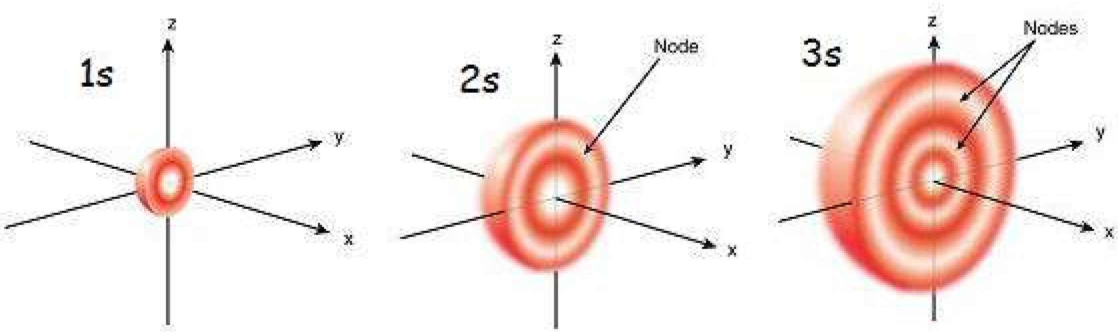
- Predict the number of subshells in the fourth shell, $n = 4$. Give the label for each of these subshells. How many orbitals are in each of these subshells?

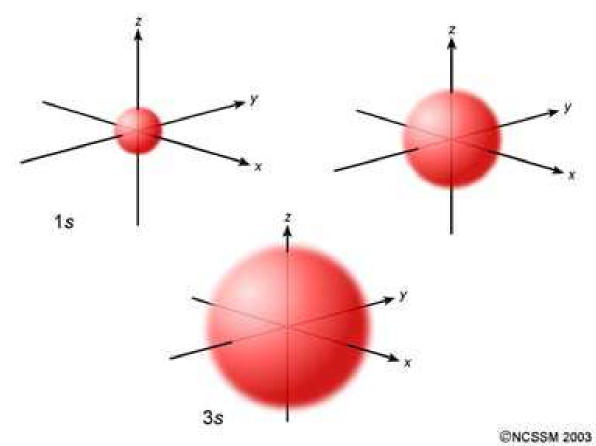
Practice Exercise

- What is the designation for the subshell with n = 5 and l = 1?
- How many orbitals are in this subshell?
- Indicate the values of m_l for each of these orbitals.

Section 6.6 – Representations of Orbitals

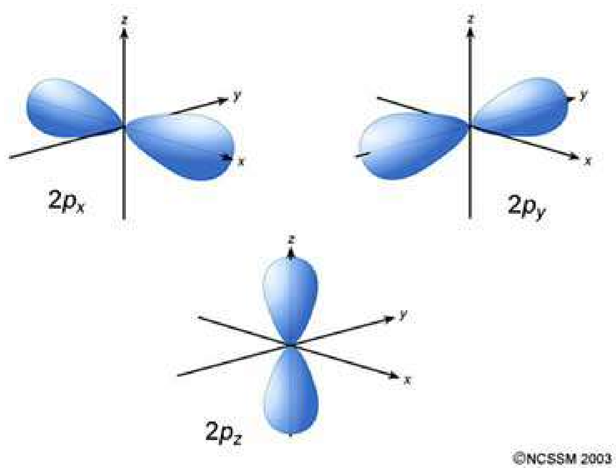
- The s orbital has spherical symmetry. As n increases, the distance from the nucleus increases for s orbitals.
- An intermediate point at which a probability function goes to zero is called a node.

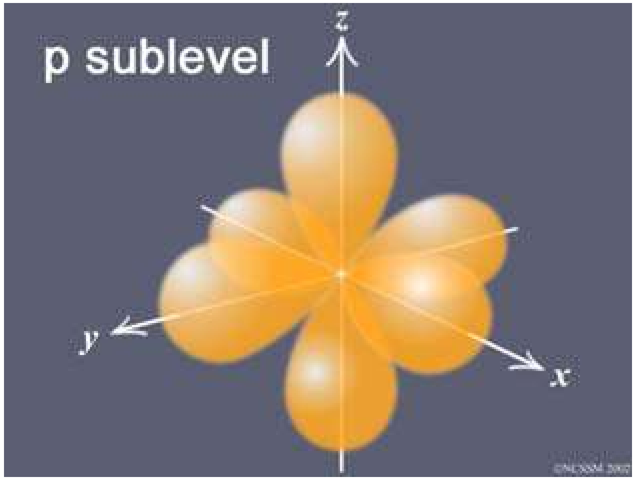




p Orbitals

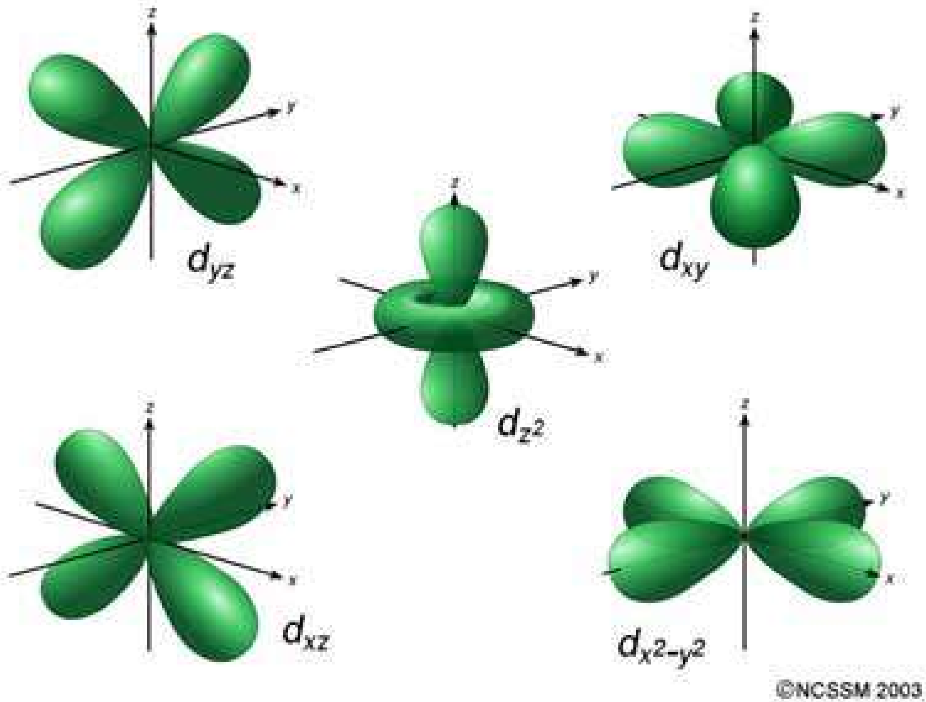
- The electron density for p orbitals occurs in two regions on either side of the nucleus, separated by a node. The two regions are called lobes.
- Each p orbital has the same energy.





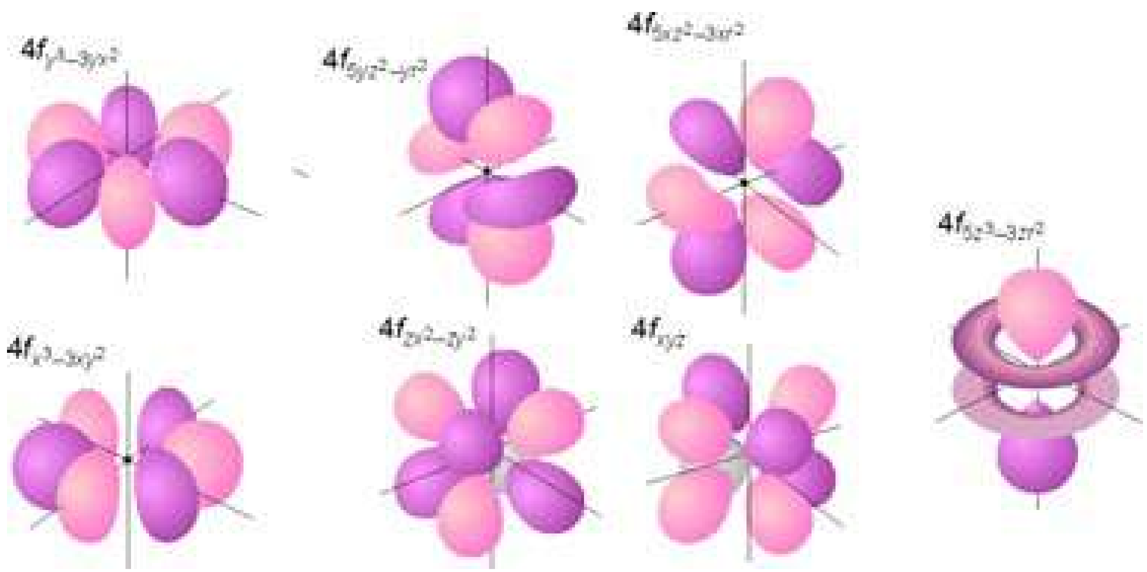
d Orbitals

- There are 5 d orbitals which all have the same energy.



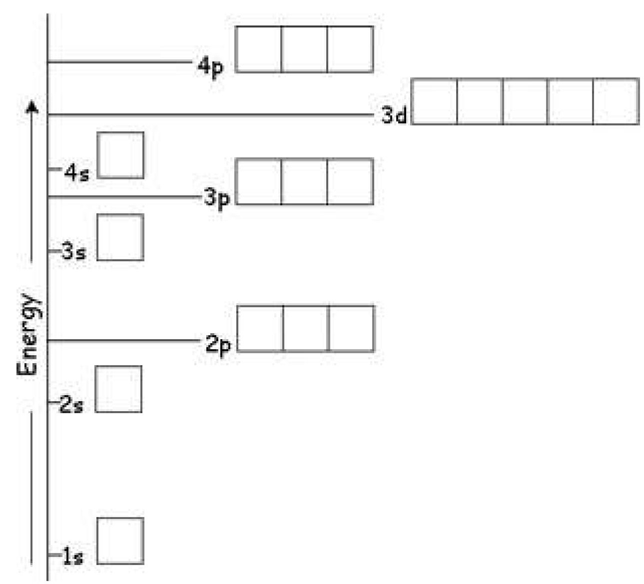
f Orbitals

- There are 7 f orbitals which all have the same energy.

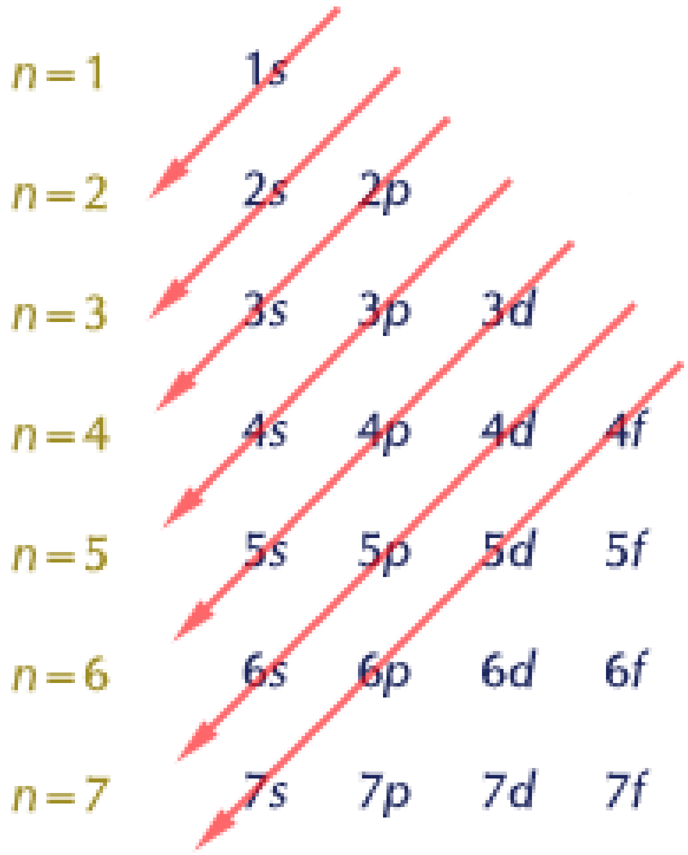


Section 6.7 – Many-Electron Atoms

- In many-electron atoms, the electron-electron repulsions cause the different subshells to have different energies.
- Orbitals with the same energy are degenerate.

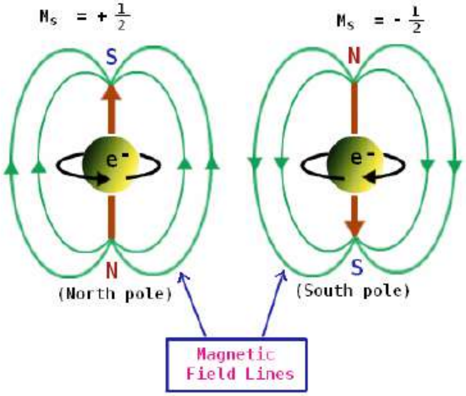


Order of the Sublevels



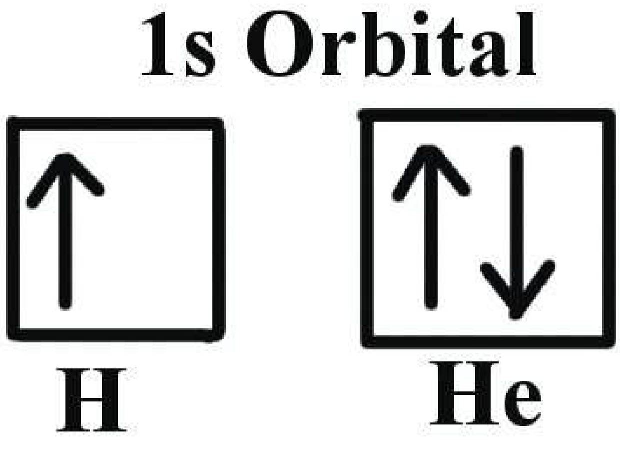
Electron Spin

- Electron spin is a property that causes electrons to behave as if it were a tiny sphere spinning on its own axis.
- A spinning charge produces a magnetic field.
- The two opposite directions of spin therefore produce oppositely charged magnetic fields.



Pauli Exclusion Principle

- The Pauli exclusion principle states that no two electrons in an atom can have the same set of 4 quantum numbers, so an orbital can hold a maximum of 2 electrons and they must have opposite spins.



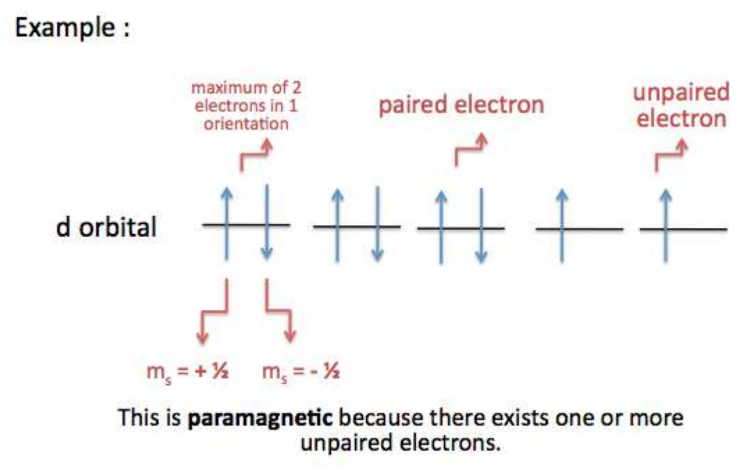
Section 6.8 – Electron Configurations

- The way in which the electrons are distributed among the various orbitals of an atom is called the electron configuration.
- Aufbau's principle states that electrons fill the orbitals in order of increasing energy.

Example	electron notation	orbital notation
hydrogen	$1s^1$	↑
helium	$1s^2$	↑↓
lithium	$1s^2\ 2s^1$	↑↓ ↑ _ _ _
beryllium	$1s^2\ 2s^2$	↑↓ ↑↓ _ _ _
boron	$1s^2\ 2s^2\ 2p^1$	↑↓ ↑↓ ↑ _ _

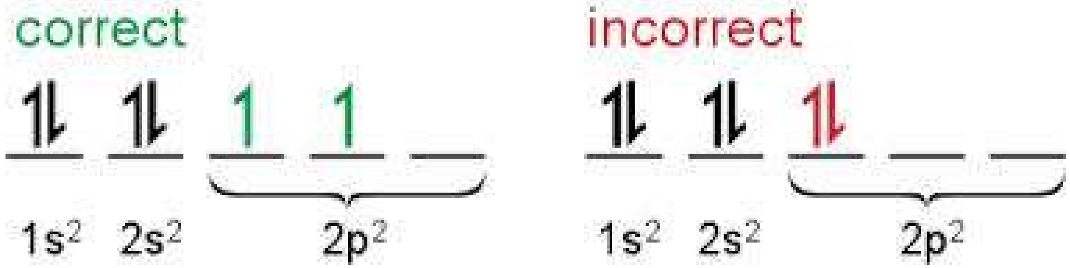
Electron Pairing

- Electrons having opposite spins are said to be paired when they are in the same orbital.
- An unpaired electron is one that is not accompanied by a partner of opposite spin.



Hund's Rule

- Hund's rule states that for degenerate orbitals, the lowest energy is attained when the number of electrons with the same spin is maximized.
- Hund's rule is based on the fact that electrons repel one another.



Sample Exercise 6.7

- Draw the orbital diagram for the electron configuration of oxygen. How many unpaired electrons does an oxygen atom possess?

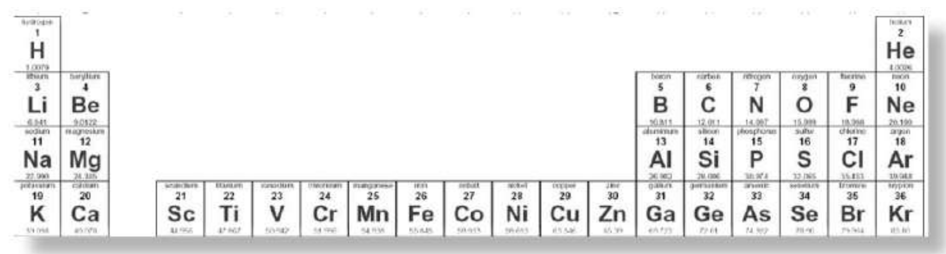
Practice Exercise

- Write the electron configuration for phosphorus. How many unpaired electrons does a phosphorus atom possess?

Noble Gas Configuration

- In writing the condensed/noble gas electron configuration, the nearest noble gas of lower atomic number is represented by its symbol in brackets.
- The inner-shell electrons (core electrons) are left out, only the outer-shell electrons (valence electrons)

are used.



Vanadium

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

Argon

$1s^2 2s^2 2p^6 3s^2 3p^6$

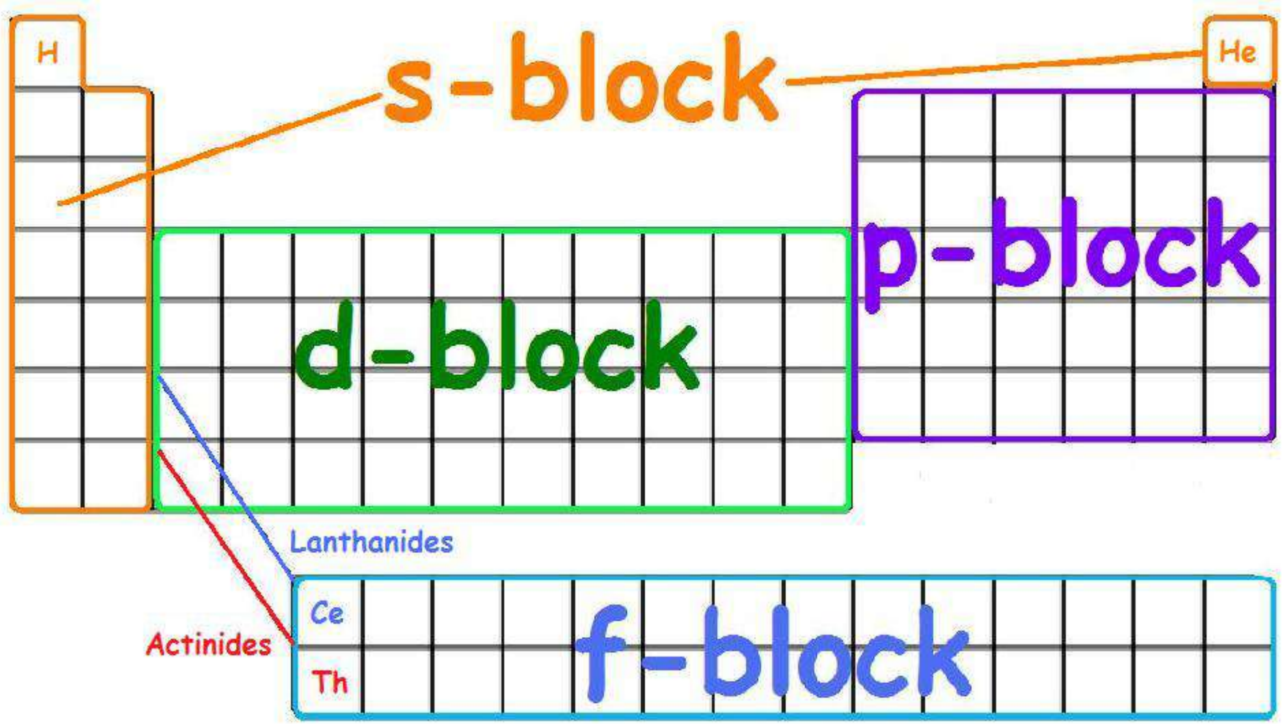
$\longrightarrow$

Vanadium

$[Ar]\,4s^2 3d^3$

Section 6.9 – Electron Configurations and the Periodic Table

- The periodic table is structured so that the elements with the same valence electron configuration are arranged in columns.



Sample Exercise 6.8

- What is the characteristic valence electron configuration of the halogens?

Practice Exercise

- Which group of elements is characterized by an ns^2np^2 electron configuration?

Sample Exercise 6.9

- Write the condensed/noble gas electron configuration for bismuth. How many unpaired electrons does an atom of bismuth possess?

Practice Exercise

- Write the noble gas/condensed electron configurations for the following:
- cobalt
- tellurium

Odd Electron Configurations

- Sublevels are the most stable when ½ full or completely full.
- It is more acceptable for a large sublevel to be stable.
- Most odd electron configurations such as chromium and copper are based on this fact.

21	22	23	24	25	26	27	28	29	30
Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
$3d^1$	$3d^2$	$3d^3$	$3d^5$	$3d^5$	$3d^6$	$3d^7$	$3d^8$	$3d^{10}$	$3d^{10}$

Anomalous configurations

Sample Integrative Exercise

a. Boron occurs naturally as two isotopes, ^{10}B and ^{11}B, with natural abundances of 19.9% and 80.1%. In what ways do the two isotopes differ from one another? Does the electron configuration of ^{10}B differ from that of ^{11}B?

Sample Integrative Exercise

b. Draw the orbital diagram for an atoms of ^{11}B. Which electrons are the valence electrons?

Sample Integrative Exercise

c. Indicate 3 major ways in which 1s electrons in boron differ from its 2s electrons.

Sample Integrative Exercise

d. Elemental boron reacts with fluorine to form BF_3, a gas. Write a balanced chemical equation for the reaction of solid boron with fluorine gas.

Sample Integrative Exercise

e. ΔH_f^o for $BF_{3(g)}$ is -1135.6kJ/mol. Calculate the standard enthalpy change in the reaction of boron with fluorine.

Sample Integrative Exercise

f. When BCl_3, also a gas at room temperature, comes into contact with water, the two react to form hydrochloric acid and boric acid, H_3BO_3, a very weak acid in water. Write a balanced net ionic equation for this reaction.