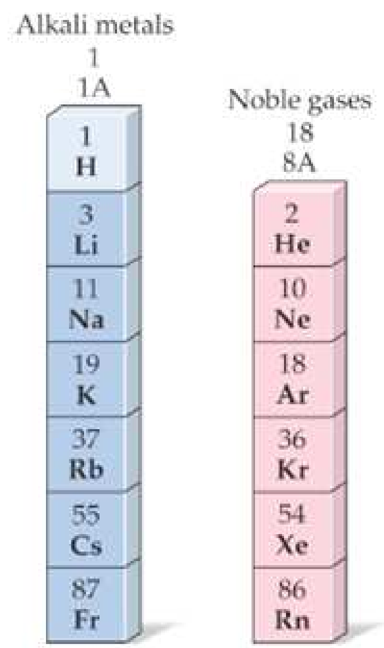
Chapter 9: Electrons in Atoms and the Periodic Table

Why is hydrogen reactive and flammable but helium is not?

- Not only is hydrogen reactive but so are ALL the Group IA elements.
- Helium and the other Noble Gases are ALL unreactive.



Recall Mendeleev's Periodic Law: When the elements are arranged in order of increasing atomic number, **certain sets of properties reoccur periodically.**

WHY???

Elemental properties reoccur periodically because of **the behavior of the electrons** in those atoms.....

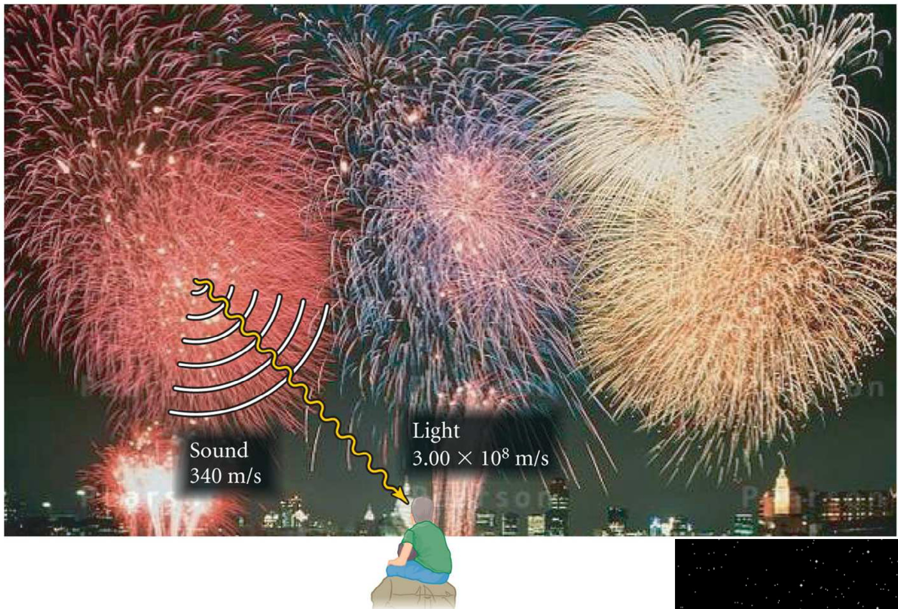
To understand the behavior of electrons in atoms, we first have to understand the behavior of light.

The Nature of Light - It's A Wave!

Light is one of the forms of energy

- Visible light - the form of ***<u>electromagnetic radiation</u>*** that we can "see"
- Microwaves, radiowaves, x-rays, ultraviolet radiation, infrared radiation - the forms of electromagnetic radiation that we can't see.
- Electromagnetic radiation travels in waves, like waves move across the surface of a pond

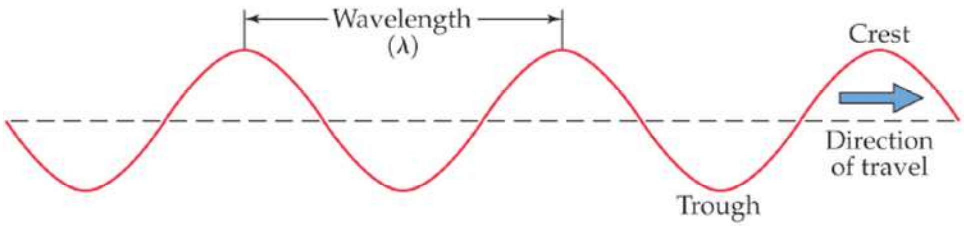
All electromagnetic radiation travels in waves at 3.00×10^8 m/sec (the speed of "light" in a vacuum) and obeys the physics of waves.



How fast is 3.00×10^8 m/sec?

The Electromagnetic Spectrum

All types of electromagnetic radiation are characterized by their wavelength and frequency



__**Wavelength (λ)**__: a measure of the distance covered by a wave (the distance from one crest to the next)

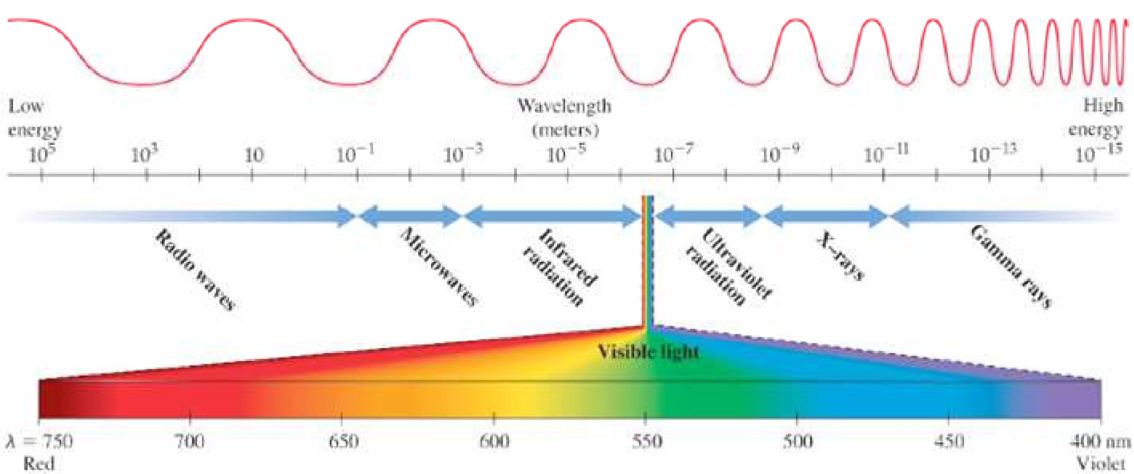
- Visible light λ usually measured in nanometers (1 nm = 1 x 10^{-9} m)

Frequency (ν): the number of waves that pass a point in a second

- The number of __waves per second__ = number of __cycles per second__
- Units for frequency = hertz (Hz) or cycles/sec (1 Hz = 1 cycle/sec)

All electromagnetic radiation obeys the equation:

$$\boxed{\lambda \times \nu = 3.00 \times 10^{8} \text{ m/sec}}$$

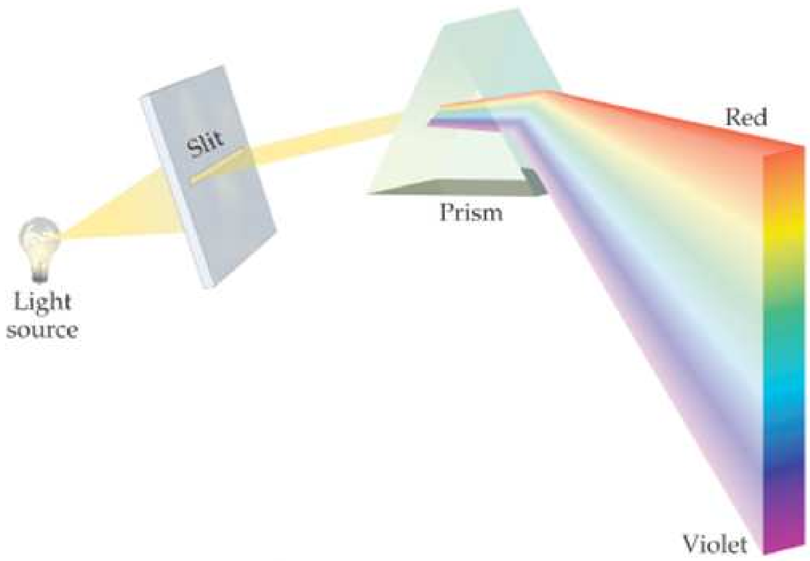


Wavelength and frequency are __**inversely related**__

- when λ increases, ν decreases; when λ decreases, ν increases

Visible Light and Color

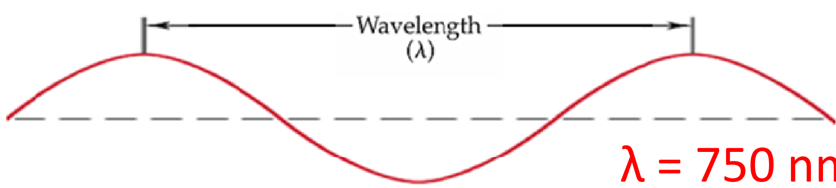
The Visible Spectrum



R ≅ 750 nm
O
Y
G
B
I
V ≅ 380 nm

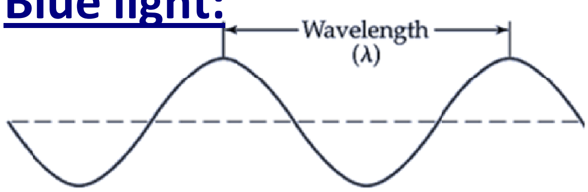
Visible light is separated into its constituent colors—red, orange, yellow, green, blue, indigo, and violet—when it is passed through a prism. The color of the light is determined by it's wavelength.

Red light:



λ = 750 nm

Blue light:



λ = 450 nm

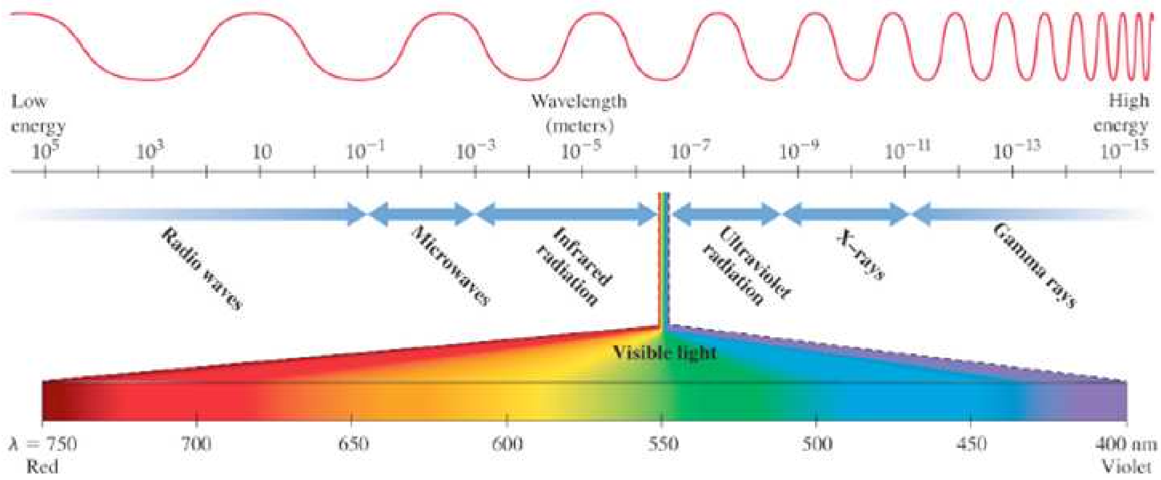
When an object absorbs some of the wavelengths of white light while reflecting others, it appears colored.

- The observed color is predominantly the colors reflected

The Nature of Light - It's a Particle!

Scientists in the early 20th century showed that electromagnetic radiation was composed of particles we call **photons,** giving rise to the **wave-particle duality** of electromagnetic radiation.

- *Each wavelength* has photons that have a *different amount of energy*



What is the relationship between energy and wavelength?

Write the order of the following types of electromagnetic radiation: microwaves, gamma rays, green light, red light, ultraviolet light

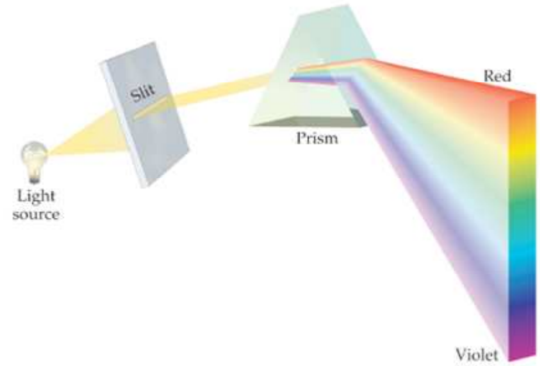
- By wavelength (short to long)

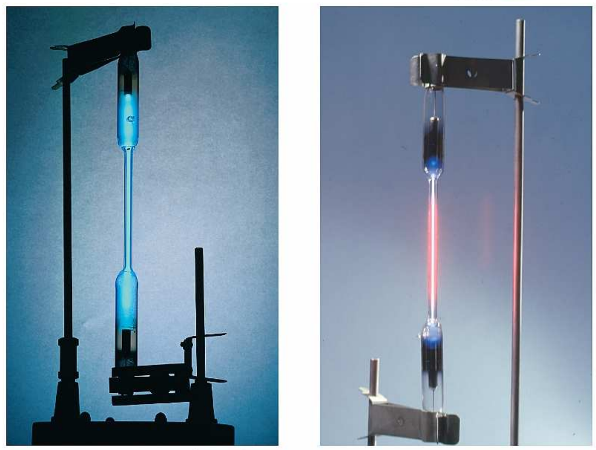
- By frequency (low to high)

- By energy (least to most)

Atomic Line Spectra

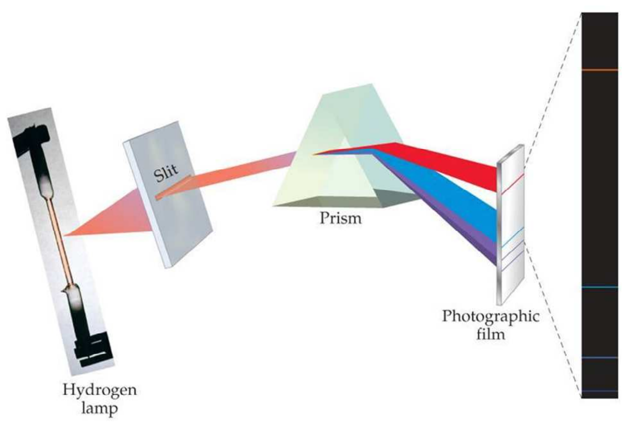
If white light is passed through a prism (or diffraction grating) it separates into a *continuous spectrum* - a rainbow of colors.



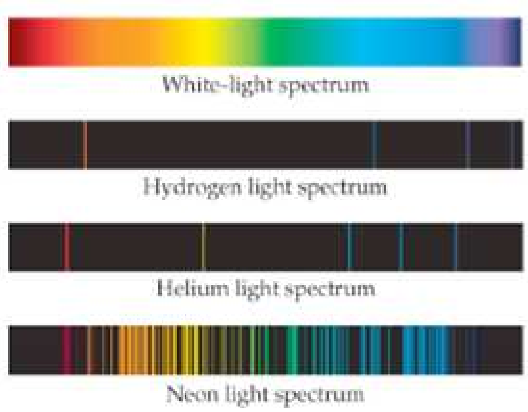


$Hg(g)$ $H_2(g)$

If an element is heated enough so that it gives off light and that light is passed through a diffraction grating, only narrow, colored lines are observed - called a **LINE SPECTRUM**.



The wavelengths of light emitted depends on the element.



Bohr Model of the Atom

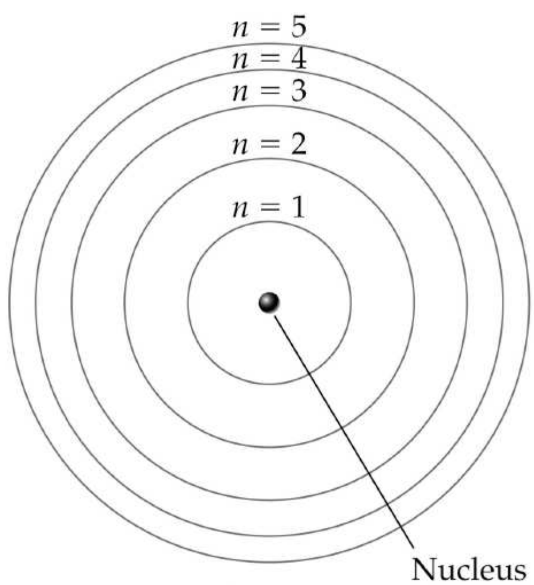
Why are only certain colors of light given off by glowing elements?

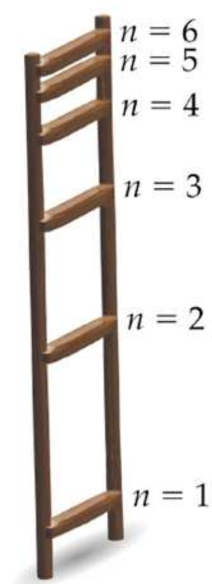
The **<u>Bohr model</u>** can help explain why:

- Electrons orbit at specific fixed distances from the nucleus, called **<u>shells</u>**

- The energy of each shell is **<u>quantized</u>**, which means that an electron in a particular shell can only have a **<u>very specific energy</u>**.
  (The energy of a shell is specified by the quantum number n, where n = 1, 2, 3, etc).

- The farther the electron is from the nucleus the more energy it has.



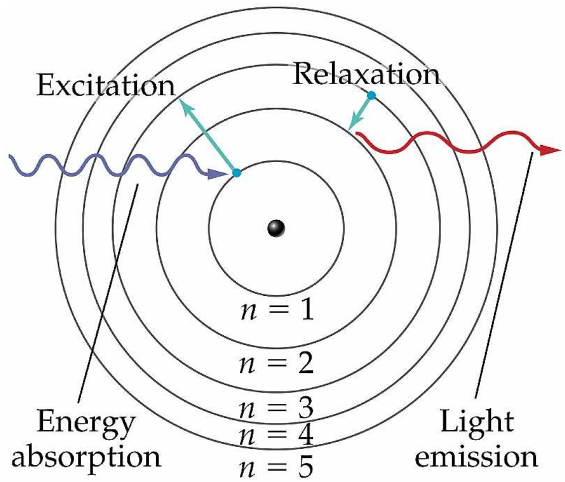


Quantized energy levels in the Bohr model of the atom are like steps on a ladder: you can stand on one step or another, but <u>you can't stand in between</u>.

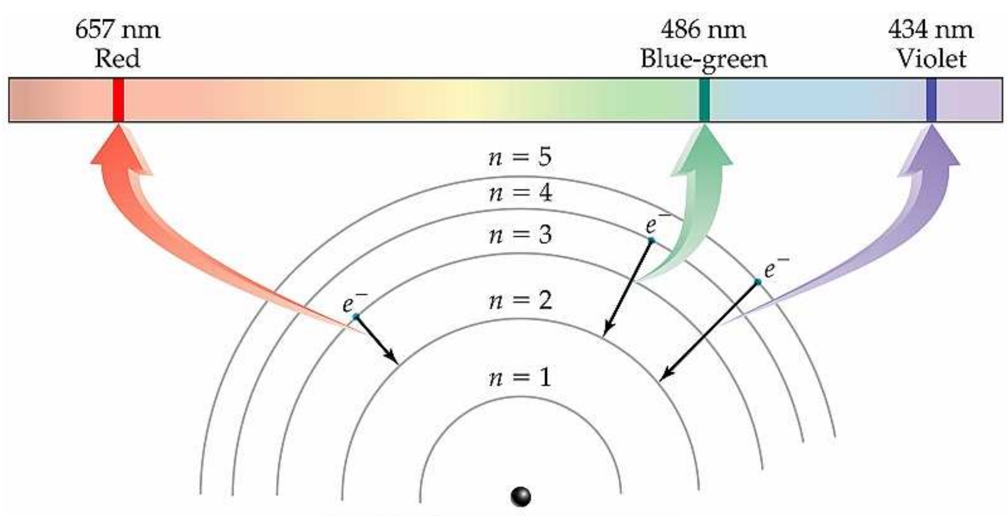
Bohr Model of the Atom

- Electrons **jump** to higher energy shells when they **absorb** energy (called "excitation")

- Electrons **drop** down to lower energy shells and they **give off** a <u>**defined amount of energy**</u> (called "relaxation")



- The defined amount of energy given off during an electron's transition <u>to a lower energy shell</u> corresponds to the emission of a <u>photon</u> of a particular wavelength of light (i.e. n= 2 to n=1, n=4 to n=2, etc).
- The closer the shells are in energy, the lower the energy of the photon emitted (Lower energy = longer wavelength)



The Bohr model helped explain the SHARP LINE SPECTRA of the elements and introduced the idea of quantization of energy levels.

The Quantum-Mechanical Model of the Atom

Shortcomings of the Bohr model:

- Can predict the energies of the shells in a hydrogen atom, but <u>not for any other element</u>
- Assumes electrons travel in single circular paths, which doesn't fit with the known laws of physics
- Does <u>**not**</u> explain the fact that certain sets of element properties reoccur periodically

However, the Bohr model of the atom did introduce <u>quantization of energy levels</u>, which was correct.

<u>Quantum mechanical model:</u>

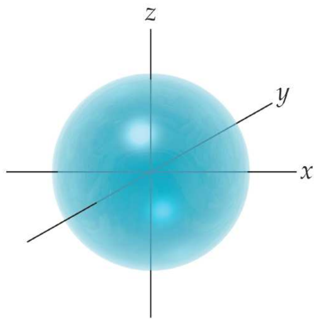
- Assumes electrons can act like particles <u>**or**</u> waves
- Erwin Schrödinger treated electrons as matter waves and applied the mathematics of probability to the physics equations that describe waves.
- Solutions to the 'Wave Equation' describe the <u>probability of finding an electron within a certain region of space</u> called an **ORBITAL**
- Quantization of energy levels came naturally from solutions to the 'Wave Equation'

Erwin Schrödinger

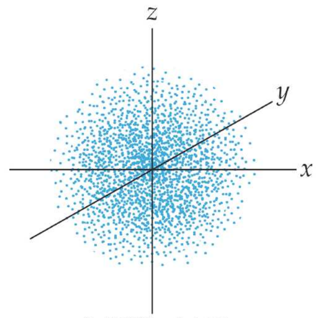
Exact path of a single electron is impossible to predict with the quantum mechanical model

Shells, Subshells, and Orbitals

Orbital: 3-dimensional region around the nucleus where an electron is most likely to be found 90% of the time



1s orbital



1s orbital electron probability map

The quantum mechanical model has **shells** like the Bohr model

- *n* = 1, 2, 3, 4, … (principal quantum number)
- As *n* increases, energy of the electron increases, and **average** distance from the nucleus increases.

But shells alone cannot explain electron behavior. Each new row on the periodic table is a new shell, and the major sections (main group, transition, inner transition) each have their own **subshell**.

Each shell has *n* subshells, with each subshell within a shell having a slightly different energy ($s < p < d < f$)

Shell	# subshells	Subshell letters
n = 1	1	1s
n = 2	2	2s, 2p
n = 3	3	3s, 3p, 3d
n = 4	4	4s, 4p, 4d, 4f

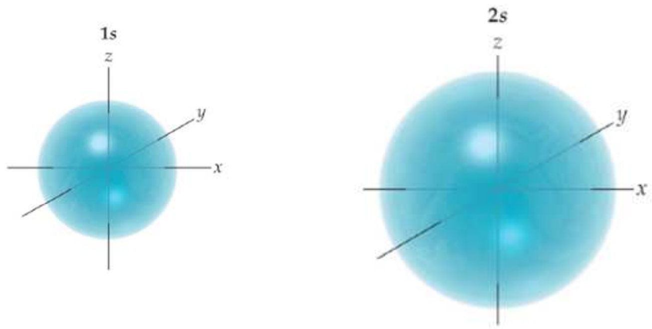
NOTE: The subshells in a shell of H all have the same energy, but for multielectron atoms the subshells have different energies.

s and p Orbitals

Each kind of subshell (*s, p, d or f*) has orbitals with a particular shape.

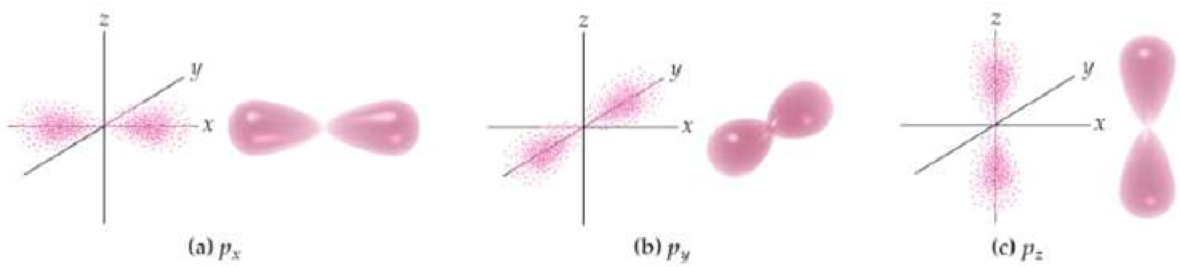
- ➢ The shape represents the probability map (90% probability of finding electron in that region)

s-orbitals



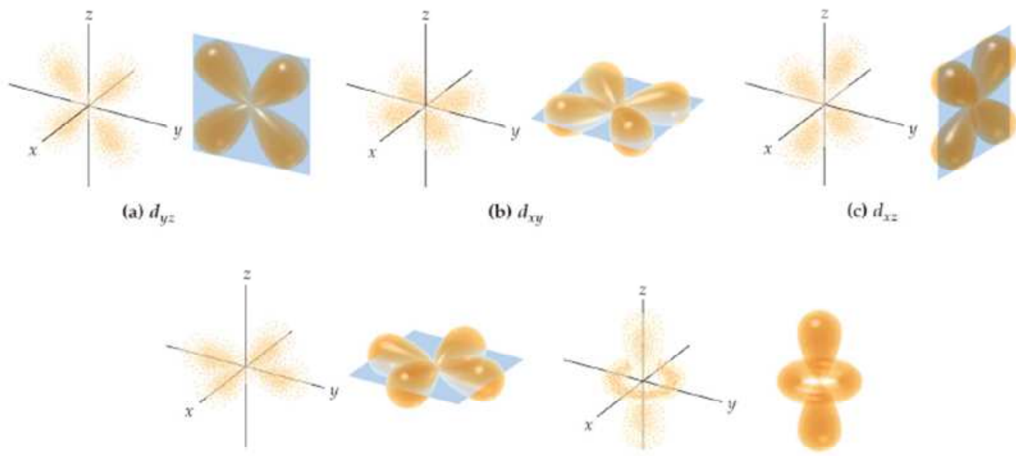
- S-orbitals are ALL spherical in shape, but the size of the s-orbital increases as n increases.
- Each energy level (n= 1, 2, 3, etc.) has ONE s-subshell that contains ONE s-orbital

p-orbitals



- There are 3 p-orbitals in each p subshell, with all 3 p orbitals having the same energy
- Each p-orbital has the same figure eight shape but has a different orientation in space
- The size of the p-orbitals increase as n increases

d-Orbitals



(a) d_{yz} (b) d_{xy} (c) d_{xz} (d) $d_{x^2-y^2}$ (e) d_{z^2}

- Each d-subshell contains <u>5-d orbitals</u>, with 4 out of the 5 orbitals having a four leaf clover shape (the other one looks like a p-orbital with a donut around it's middle).
- Each d-orbital has a different orientation in space, but all 5 d-orbitals within a d-subshell <u>have the same energy</u>
- The size of the d-orbitals increase as n increases

<u>f-orbitals</u>

- Each f-subshell has <u>7-f orbitals</u>, with each orbital having a similar shape but a different orientation in space.
- All 7 f-orbitals in a f-subshell <u>have the same energy</u>
- The size of the f-orbitals increase as n increases

Shell	# of subshells	Letters specifying subshells			
$n = 4$	4	*s*	*p*	*d*	*f*
$n = 3$	3	*s*	*p*	*d*	
$n = 2$	2	*s*	*p*		
$n = 1$	1	*s*			

Distribution of Electrons in Atoms

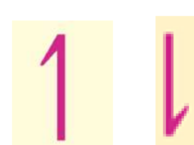
- ➢ Each orbital can hold **NO MORE THAN TWO ELECTRONS**

- ➢ Consequently, each shell and subshell has a maximum number of electrons it can hold
    - s subshell (with 1 s-orbital) = 2 e^-s max
    - *p* subshell (with 3 p-orbitals) = __ e^-s max
    - d subshell (with 5 d-orbitals) = __ e^-s max
    - f subshell (with 7 f-orbitals) = __ e^-s max

We place electrons in the shells and subshells in order of energy, from lower energy to higher, which is called the **Aufbau principle.**

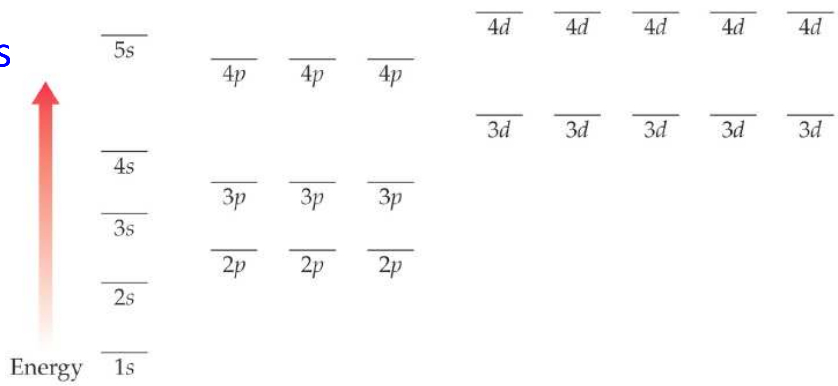
A couple more important points:

- Electrons spin on an axis generating a very small magnetic field
- When two electrons are in the same orbital, they must have opposite spins so their magnetic fields will cancel
- "Spin state" of an electron is indicated by an arrow



Electron Configurations

Electron Configuration: a notation that shows the distribution of electrons **following the Aufbau principle** into the various shells and subshells of an atom.

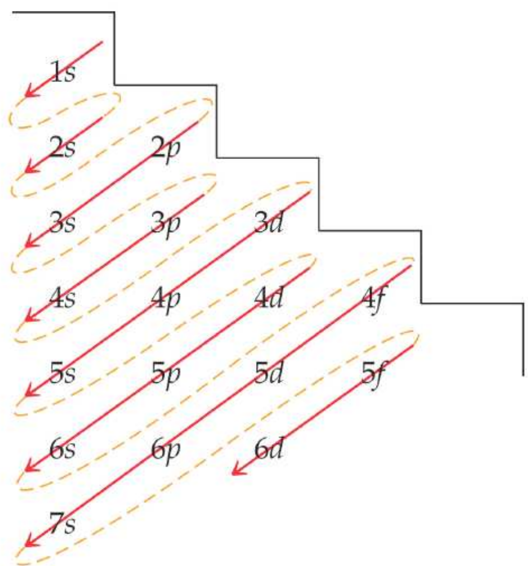


Energy Ordering of Orbitals for Multi-Electron Atoms

Order of Subshell Filling in Ground State Electron Configurations

Draw a diagram putting each energy shell in a row and listing the subshells (*s*, *p*, *d*, *f*) for that shell in order of energy (left to right).

Draw arrows through the diagonals, looping back to the next diagonal each time.



Let's use the Aufbau principle to write ground state electron configurations of a couple of elements:

Electron configuration: subshell$^{\# \, e^-}$

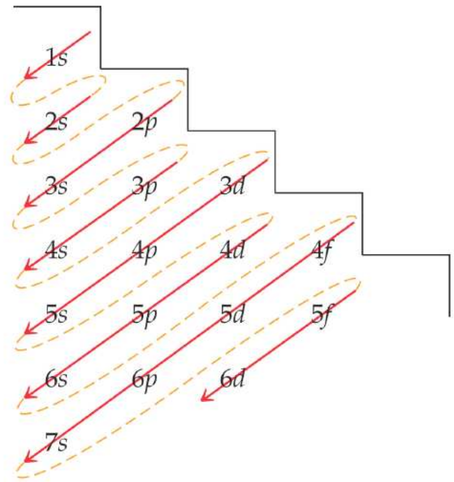
Examples:

H: $1s^1$ **He**: $1s^2$ Li: Be:

subshell # of e⁻s in subshell (3 e⁻s) (__e⁻s)

Electron Configurations

The electron configuration is a listing of the subshells in order of filling with the <u># of electrons</u> in that subshell <u>written as a **superscript**</u>



O: $1s^22s^22p^4$

($8e^-s$)

Cl : $1s^22s^22p^63s^23p^5$

($17\ e^-s$)

Fe : $1s^22s^22p^63s^23p^64s^23d^6$

($26\ e^-s$)

Kr : $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$

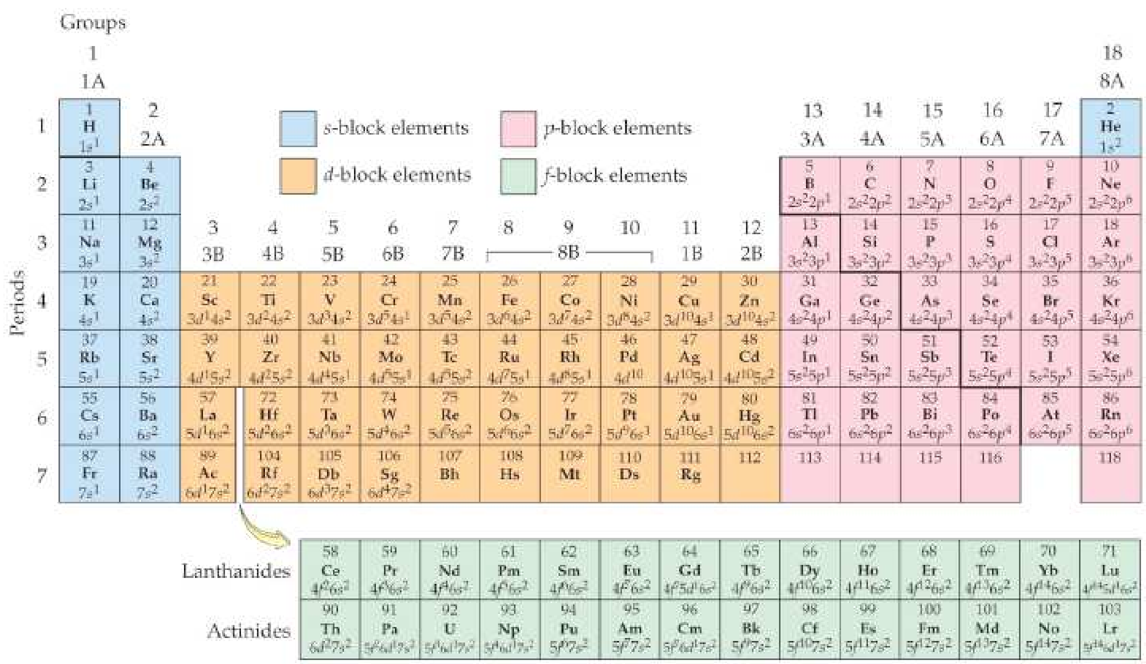
(__ e^-s)

Si :

(__ e^-s)

Ni:

Br:



You can also use the periodic table as a guide in writing the ground state electron configurations of elements and monoatomic ions

Noble Gas Shorthand Notation

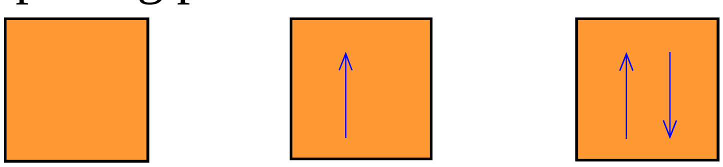
A shorthand way of writing an electron configuration is to use the symbol of the previous noble gas in brackets[] to represent all the inner electrons, then just write the electron configuration of the last set.

$$\text{Rb:}1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1 \quad = \quad \text{Rb: [Kr]}5s^1$$

What is the electron configuration using the NOBLE GAS SHORTHAND NOTATION of phosphorus, P? of Se?

Orbital diagrams show the how many electrons occupy **each orbital in a subshell**

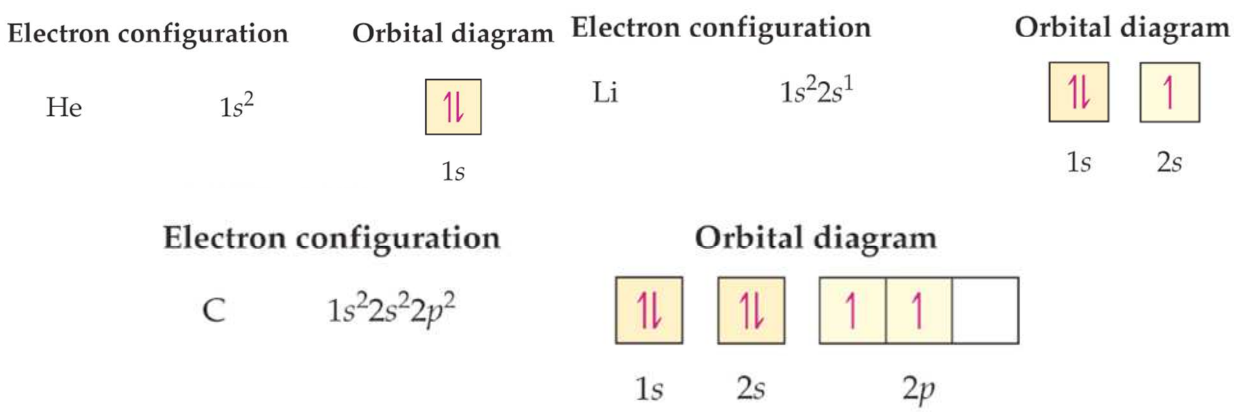
- We often represent an orbital as a square and the electrons in that orbital as arrows
- The direction of the arrow represents the spin state of the electron.
- When **filling orbitals that have the same energy**, place one electron (with the same arrow direction) in each orbital before completing pairs - called Hund's rule



Unoccupied orbital | Orbital with 1 electron | Orbital with 2 electrons

Orbital Diagrams

Comparison of Electron Configurations & Orbital Diagrams



Write the orbital diagrams of:

N: Cl:

O: As:

Write the electron configuration AND orbital diagram of:

A) the fluoride ion, F^-

B) the magnesium cation, Mg^{2+}

C) the scandium (III) cation, Sc^{3+}

Valence Electrons

<u>Valence electrons</u>: electrons in the outermost **<u>shell</u>**
(outmost shell has the highest principle quantum #, n)

- Electrons in lower energy shells are called **core electrons**.
- Chemists have observed that one of the most important factors in the way an atom behaves, both chemically and physically, is based upon the <u>number of valence electrons the atom contains.</u>

Examples:

Rb: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$

- The outermost shell of Rb that contains electrons is the 5th shell, therefore, Rb has 1 valence electron and <u>36 core electrons</u>

Kr: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

- The outermost shell of Kr that contains electrons is the 4th shell, therefore, Kr has 8 valence electrons and <u>28 core electrons</u>.

Determine the number and types of valence electrons in:
a) helium, He

b) iron, Fe

c) arsenic, As

Relationship Between Valence Electrons, the Periodic Table and Element Properties

1A	2A	3A	4A	5A	6A	7A	8A
1 **H** $1s^1$							2 **He** $1s^2$
3 **Li** $2s^1$	4 **Be** $2s^2$	5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$
11 **Na** $3s^1$	12 **Mg** $3s^2$	13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$

- Elements in the same period have valence electrons in the same principal energy shell.
  (The number of valence electrons increases by one as you progress across the period in the **main group elements**).

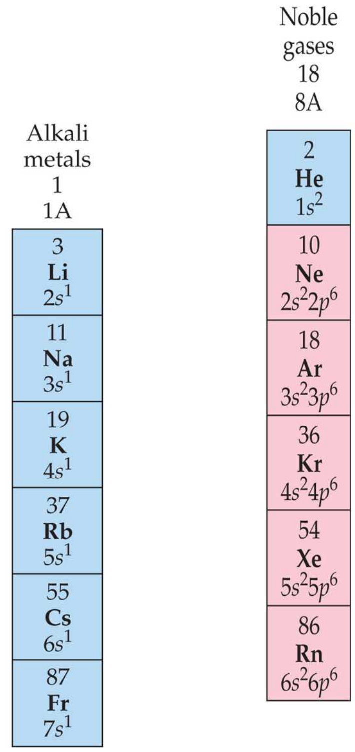
- Elements in the **same group** have the **same number of valence electrons** and the valence electrons are in the **same type of subshells** (exception: He)

Elements in the same column have similar chemical and physical properties because their valence shell electron configuration is the same.

The number of valence electrons for the main group elements is the same as the group number.

Explanatory Power of the Quantum Mechanical Model

- Noble gases are especially unreactive because they have a FULL VALENCE SHELL - 8 electrons
- The alkali metals have one more electron than the previous noble gas.
- In their reactions, the alkali metals tend to <u>lose their extra electron</u>, resulting in the same electron configuration as a noble gas, forming a cation with a 1+ charge.



Stable Electron Configuration and Ion Charge

- Metals form cations by losing valence electrons to get the **<u>same electron configuration</u>** as the <u>previous</u> noble gas.

Atom	Atom's electron config	Ion	Ion's electron config
Na	$[Ne]3s^1$	Na^+	[Ne]
Mg	$[Ne]3s^2$	Mg^{2+}	[Ne]
Al	$[Ne]3s^23p^1$	Al^{3+}	[Ne]
O	$[He]2s2p^4$	O^{2-}	[Ne]
F	$[He]2s^22p^5$	F^-	[Ne]

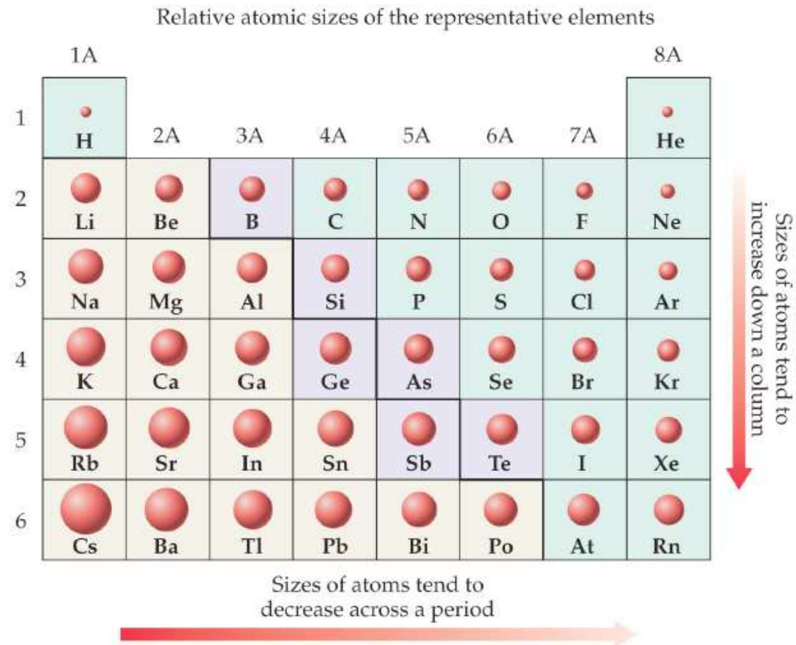
- Nonmetals form anions by gaining valence electrons to get **<u>the same electron configuration</u>** as the <u>next</u> noble gas.

Periodic Trends: Atomic Size & Ionization Energy

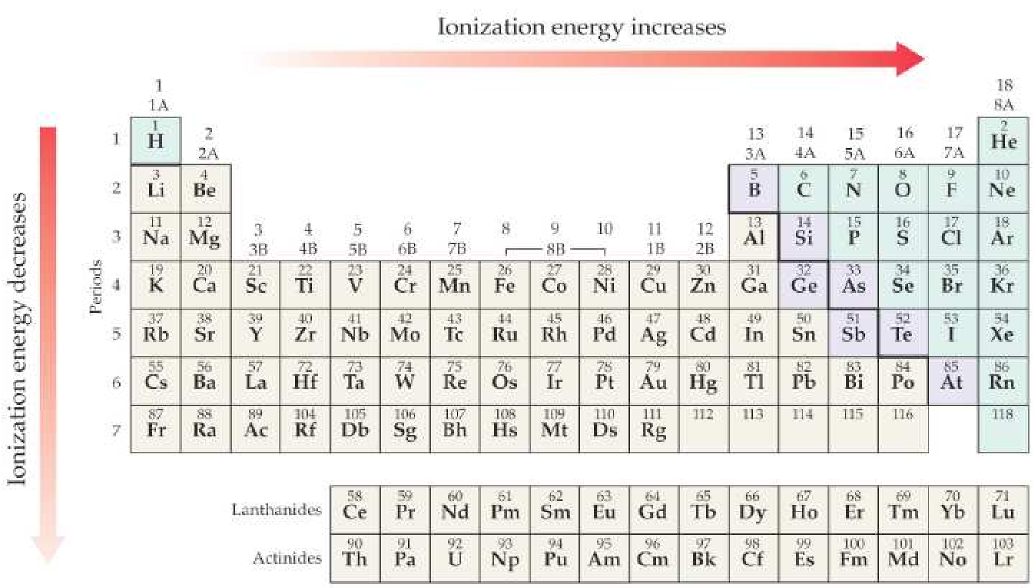
Atomic Size

As you go down a column on the periodic table, the size of the atom **increases**.
WHY?

As you go from left to right across a period (in the main group elements), the size of the atom **decreases**.



Ionization Energy (IE): the energy required to remove an electron from an atom in the gas state



As you go down a column, the first IE decreases
(Valence electron farther from nucleus and easier to remove)
Moving from left to right across a period, the IE increases

Periodic Trends: Metallic Character

Metallic Character - how well an element's properties match the general properties of a metal.

General Properties of Metals:

- Malleable and ductile
- Shiny, lustrous, and reflect light
- Conduct heat and electricity
- Form cations

General Properties of Nonmetals:

- Brittle in solid state.
- Solid surface is dull, nonreflective.
- Solids are electrical and thermal insulators.
- Form anions and polyatomic anions.

- As you go left to right across a period, the elements become less metallic.
- As you go down a group, the elements become more metallic.