ELEMENTS AND SYMBOLS

- ***Elements*** are primary substances from which all other substances are built. Elements ***cannot*** be ***broken down*** into simpler substances.

- Over time some elements have been named for planets, mythological figures, minerals, colors, geographic locations and famous people. Some examples are shown below:

Element	Source of Name
Uranium	The planet Uranus
Titanium	Titans (mythology)
Chlorine	*Chloros,* "greenish-yellow" (Greek)
Iodine	*Ioeides,* "violet" (Greek)
Magnesium	Magnesia, a mineral
Californium	California
Curium	Marie and Pierre Curie

- The ***symbol*** for most elements is the one- or two-letter abbreviation of the name of the element. Only the first letter of an elements symbol is capitalized. If the symbol has a second letter, it is written as lowercase.

  Co (cobalt)
  CO (carbon and oxygen)

- Although most of the symbols use letters from current names, some of the symbols of the elements are based on their Greek or Latin names.

  Na sodium (natrium)
  Fe iron (ferrum)

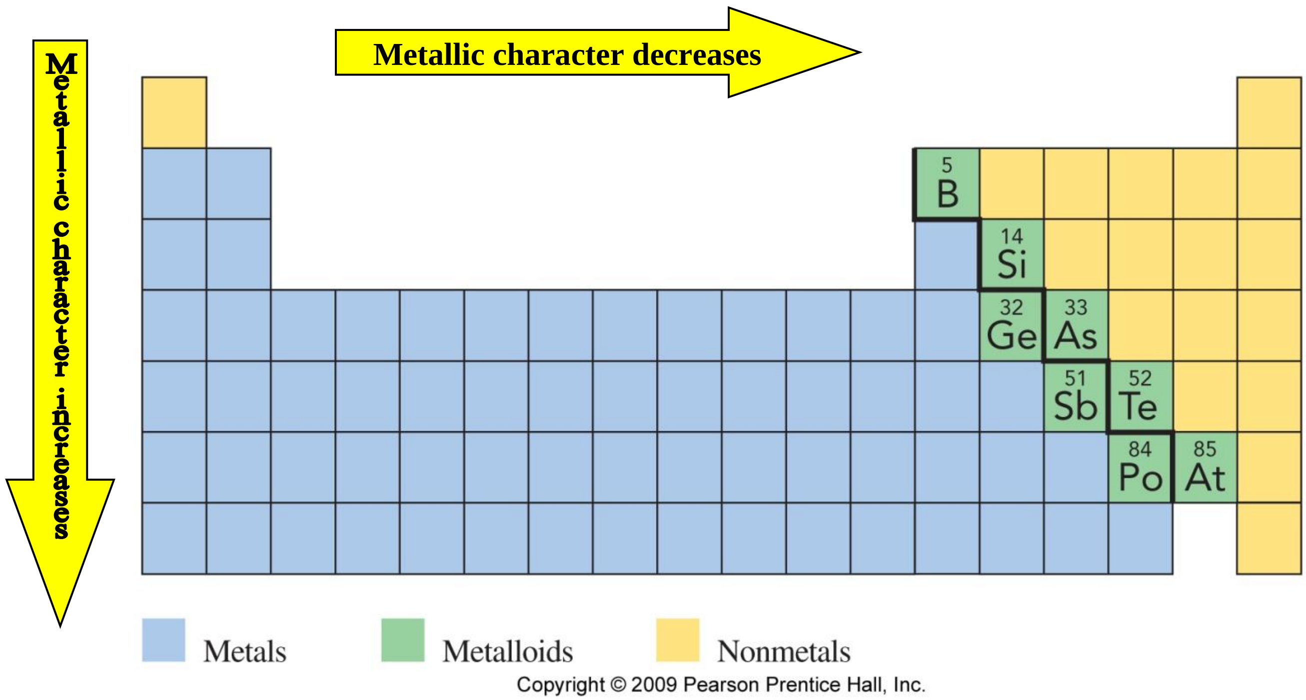
- Some elements have formulas that are ***not single*** atoms. ***Seven*** of these elements have ***diatomic*** (2-atoms) molecules.

Hydrogen	H_2	Chlorine	Cl_2
Oxygen	O_2	Fluorine	F_2
Nitrogen	N_2	Bromine	Br_2
		Iodine	I_2

PERIODIC TABLE OF THE ELEMENTS

- Arrangement of elements based on their atomic masses was first proposed by the Russian chemist, ***Dmitri Mendeleev*** in 1869.

- In the modern ***periodic table*** the elements are arranged according to their ***atomic numbers***. The elements are generally classified as ***metals***, ***nonmetals*** and ***metalloids***.



Metals	Nonmetals
1. Mostly solid	1. Can be solid, liquid or gas
2. Have shiny appearance	2. Have dull appearance
3. Good conductors of heat and electricity	3. Poor conductors of heat and electricity
4. Are malleable and ductile	4. Are brittle (if solid)
5. Lose electrons in a chemical reaction	5. Gain or share electrons in a chemical reaction

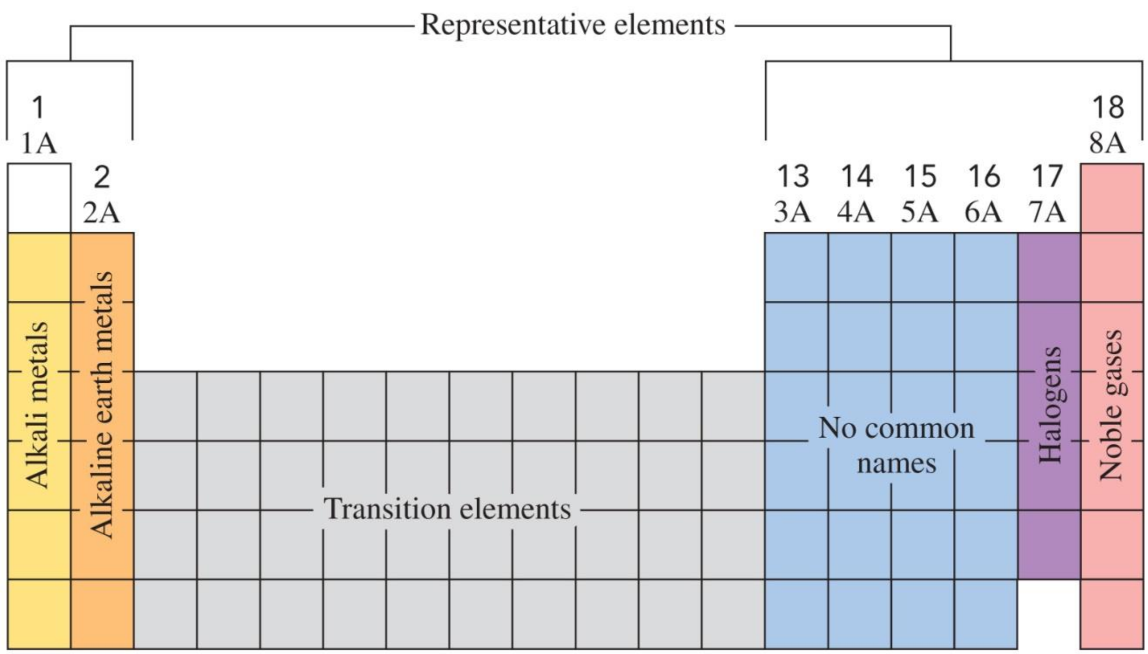
- **Metalloids** are elements that possess some properties of metals and some of non-metals. The most important metalloids are silicon (Si) and germanium (Ge) which are used extensively in computer chips.

- Seven elements (H_2, N_2, O_2, F_2, Cl_2, Br_2 and I_2) exist as **diatomic molecules**. All others exist as monatomic (single atom).

PERIODIC TABLE OF THE ELEMENTS

- The periodic table is composed of **periods** (rows) and **groups or families** (columns).

- Elements in the same family have similar properties, and are commonly referred to by their traditional names.



- Elements in groups ***1-2*** and ***13-18*** are referred to as ***main-group*** or ***representative elements***.

- ***Alkali metals*** are ***soft*** metals that are ***very reactive***. They often react explosively with other elements.

- ***Noble gases*** are ***un-reactive*** gases that are commonly used in light bulbs.

- ***Halogens*** are the most ***reactive nonmetals***, and occur in nature only as compounds.

- Group 2 elements are called ***alkaline-earth metals***. These metals are less reactive than alkali metals.

- The group of metals in between the main group elements are called the ***transition metals***.

THE ATOMIC THEORY

- The smallest particle of matter that still retains its properties is called an atom.

In the fifth century B.C., the Greek philosopher ***Democritus*** proposed that matter is composed of a ***finite number of discrete particles***, named ***atomos*** (meaning un-cuttable or indivisible)

In 1808, ***John Dalton***, built on ideas of Democritus, and formulated a ***precise definition of the building blocks of matter***.

John Dalton

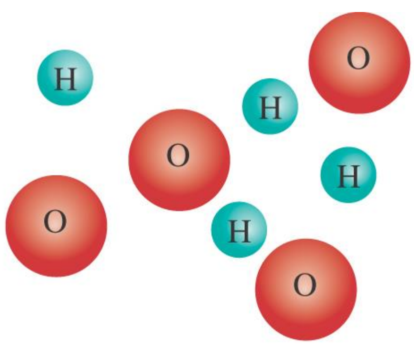
- Dalton's model represented the atom as a featureless ball of uniform density. This model is referred to as the "***soccer ball***" model.

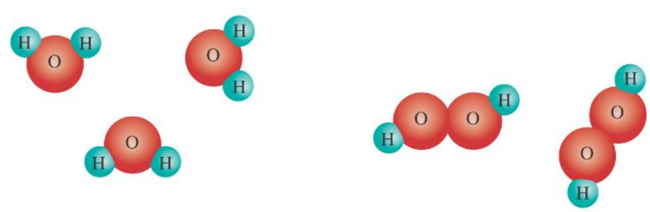
- Dalton's atomic theory consisted of 3 parts:

  1. Each ***element*** is composed of tiny indestructible particles called ***atoms***.

  2. All the ***atoms of a given element*** are ***similar*** to one another, but different from atoms of other elements.

  3. Atoms ***combine*** in simple, ***whole-number ratios*** to form ***compounds***.

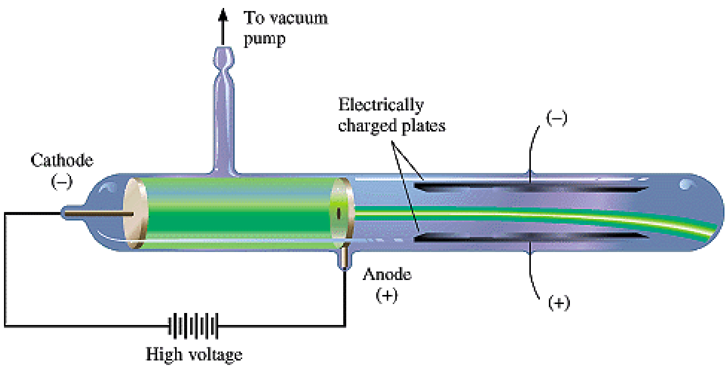


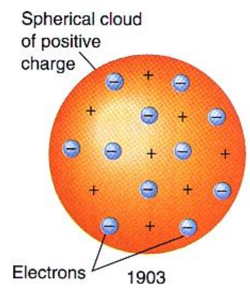


DISCOVERY OF THE ELECTRON & NUCLEUS

- Smaller particles than the atom also exist and are called subatomic particles.

- In 1897, ***J.J. Thomson*** performed experiments with a ***cathode ray*** tube. Through these experiments he discovered a negatively charged subatomic particle, that he named ***electron***.

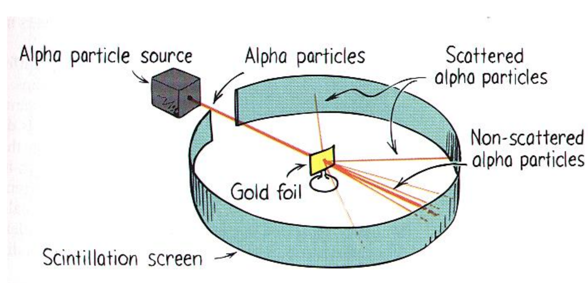




- Based on these findings, ***Thomson*** proposed an atomic model composed of ***negatively charged electrons*** embedded in a uniform ***positively charged sphere***.

- This model is called the "***plum pudding***" model.

- In 1910, ***Ernest Rutherford*** carried out a number of experiments, called gold foil experiments, to further ***probe the nature of the atom***.



- Through these experiments, he discovered that the atom was made up of a small, dense, ***positively charged nucleus*** surrounded by negatively charged electrons.

- Based on these observations, ***Rutherford*** proposed a ***model of the atom*** consisting of a ***small, massive positive center*** (***nucleus***), ***surrounded*** by ***electrons*** in ***mostly empty space***.

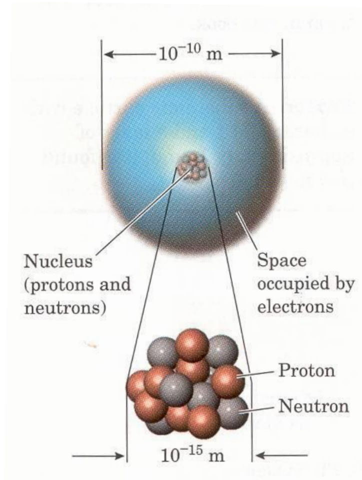
- This model is called the "***nuclear model***".

THE MODERN ATOM

In 1932, ***James Chadwick*** discovered the existence of a second nuclear particle. This ***neutral particle*** was named ***neutron***.

Current Model of the Atom:

- The atom is an electrically ***neutral*** spherical entity.
- It is composed of a ***positively charged nucleus*** surrounded by ***negatively charged electrons***.
- The ***electrons (e^-) move rapidly*** through the atomic volume, ***held*** by the ***attractive forces to the nucleus***.
- The ***atomic nucleus*** consists of positively charged ***protons*** (p^+) and neutrally charged ***neutrons*** (n^0).

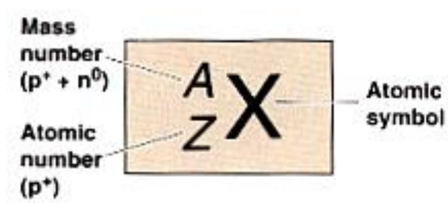


- The modern atom consists of 3 subatomic particles:

Particle	Charge	Relative Mass
PROTON	+1	~1800
NEUTRON	0	~1800
ELECTRON	–1	1

Mass Relationships in the Atom:

- The number of protons in an atom determines its identity, and is called ***atomic number (Z).***
- In a neutral atom, the number of protons (+) are equal to the number of electrons (–).



- Almost all the mass of the atom rests in the nucleus. Therefore the number of protons and neutrons in an atom is called the ***mass number (A)***.

ISOTOPES & ATOMIC MASS

- Atoms of the ***same element*** that possess a ***different*** number of ***neutrons*** are called ***isotopes***.

$^{1}_{1}H$ hydrogen $^{2}_{1}H$ deuterium $^{3}_{1}H$ tritium

- Isotopes of an element have the same atomic number (Z), but a different mass number (A).

- The ***mass*** of an atom is measured ***relative*** to the mass of a chosen ***standard*** (***carbon-12 atom***), and is expressed in ***atomic mass units*** (***amu***).

- The ***average atomic mass*** of an element is the ***mass*** of that element's natural occurring ***isotopes weighted*** according to their ***abundance***.

- Therefore the ***atomic mass*** of an element is ***closest*** to the mass of its ***most abundant isotope***.

Examples:

1. Determine the number of protons, neutrons and electrons in $^{35}_{17}Cl$.

number of p^+= number of e^- = number of n^0 =

2. Which two of the following are isotopes of each other?

$^{410}_{186}X$ $^{410}_{185}Y$ $^{412}_{183}Z$ $^{412}_{185}R$

3. Based on the information below, which is the most abundant isotope of boron (atomic mass = 10.8 u)?

Isotope	^{10}B	^{11}B
Mass (amu)	10.0	11.0

BOHR MODEL OF THE ATOM

- Protons and neutrons are contained in the small, dense nucleus of the atom, while the electrons occupy the large volume of space surrounding the nucleus.

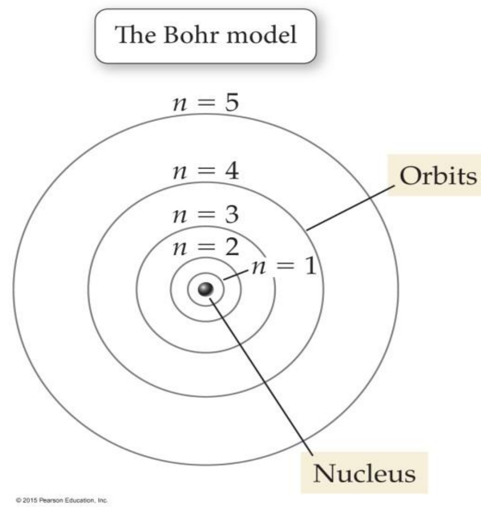
- The arrangement of the electrons within this volume is what determines the physical and chemical properties of the element.

- Scientists have now determined that electrons surrounding the nucleus occupy energy levels, with specific value of energy for each.

- A new model of the atom that embodies this concept was developed by ***Neils Bohr,*** a Danish physicist.

- ***Bohr's model*** of the atom consisted of electrons orbiting the nucleus at different distances from the nucleus, called ***energy levels***.

- In this model, the electrons could only occupy particular energy levels, and could ***"jump" to higher levels*** by ***absorbing energy***.



- The lowest energy level is called ***ground state***, and the higher energy levels are called ***excited states***. When electrons absorb energy through heating or electricity, they ***move to higher energy*** levels and become ***excited.***

- When excited electrons return to ground state, energy is emitted and is released in the form of light.

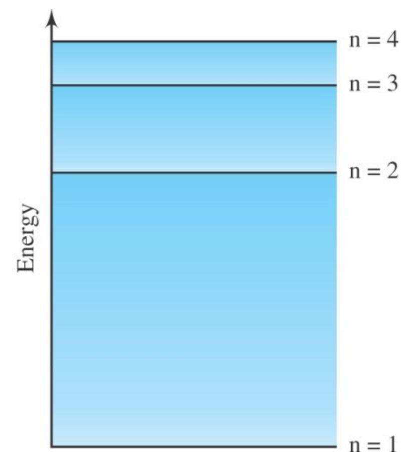
- Bohr's model of atom was called the planetary model. This model was still incomplete as it could not explain the behavior of electrons in large atoms.

QUANTUM MECHANICAL MODEL OF THE ATOM

- In 1926 ***Erwin Shrodinger*** created a mathematical model that showed ***electrons as both particles and waves***. This model was called the ***quantum mechanical*** model.

- This model predicted electrons to be located in a probability region called ***orbitals***.

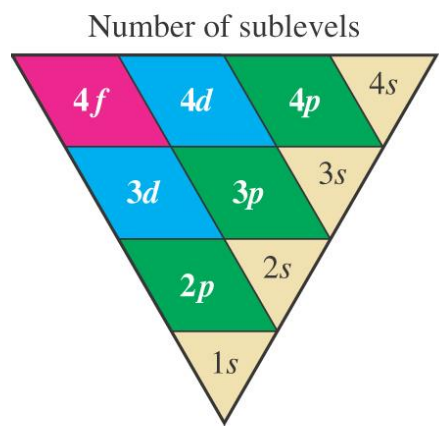
- An orbital is defined as a region around the nucleus where there is a high probability of finding an electron.



- Based on this model, there are ***discrete principal energy*** levels within the atom. Principal energy levels are designated by ***n***.

- The electrons in an atom can exist in any principal energy level. As ***n increases***, the ***energy of the electrons increases***.

- Each principal energy level is subdivided into sublevels.



- The ***sublevels*** are designated by the letters ***s, p, d and f***. As n increases, the number of sublevels increases.

- Within the sublevels, the electrons are located in ***orbitals***. The orbitals are also designated by the letters ***s, p, d and f***.

- The number of orbitals within the sublevels vary with their type.

  s sublevel = 1 orbital
  p sublevel = 3 orbitals
  d sublevel = 5 orbitals
  f sublevel = 7 orbitals

- An orbital can hold a maximum of 2 electrons.

ELECTRON CONFIGURATION

- ***Similarities*** of behavior in the periodic table are due to the similarities in the ***electron arrangement*** of the atoms. This arrangement is called ***electron configuration***.

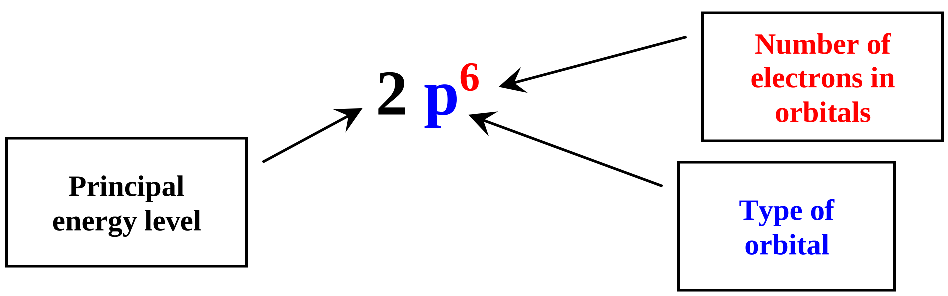
- The ***modern*** model of the atom describes the ***electron cloud*** consisting of ***separate energy levels***, each containing a ***fixed number of electrons***.

- Each orbital can be occupied by no more than 2 electrons, each with opposite spins.

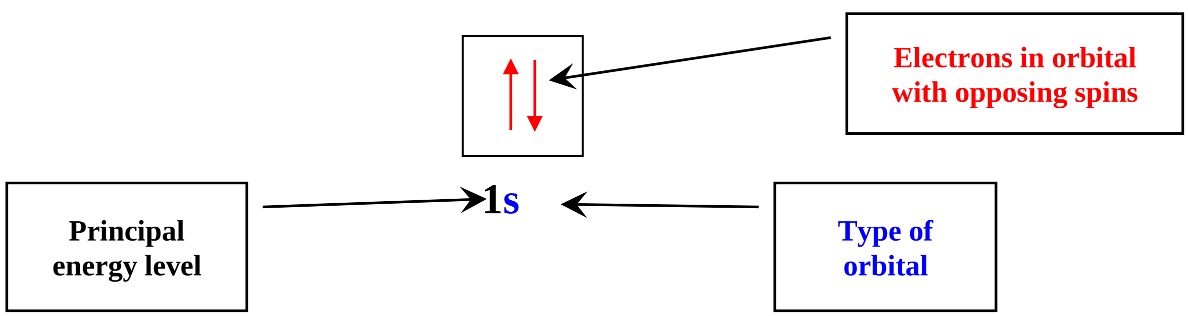
- The electrons occupy the orbitals form the lowest energy level to the highest level. The energy of the orbitals on any level are in the following order: s < p < d < f.

- Each orbital on a sublevel is occupied by a single electron before a second electron enters. For example, all three p orbitals must contain one electron before a second electron enters a p orbital.

- ***Electron configurations*** are written as shown below:



- Another notation, called the ***orbital notation***, is shown below:

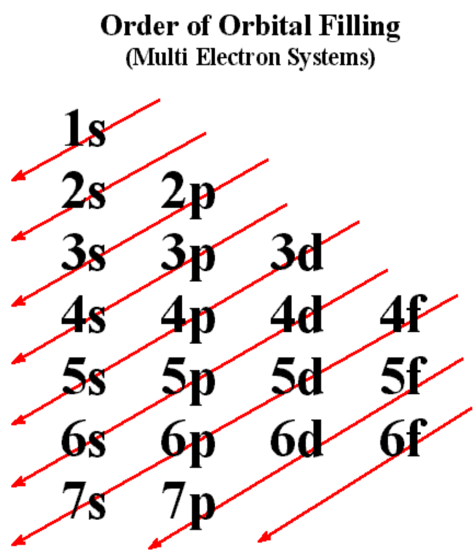


Element	***Orbital Notation***	***Configuration***
Li	□ □ 1s 2s	
Be	□ □ 1s 2s	
B	□ □ □□□ 1s 2s 2p	
C	□ □ □□□ 1s 2s 2p	
N	□ □ □□□ 1s 2s 2p	
O	□ □ □□□ 1s 2s 2p	
F	□ □ □□□ 1s 2s 2p	
Ne	□ □ □□□ 1s 2s 2p	
Na	□ □ □□□ □ 1s 2s 2p 3s	
Mg	□ □ □□□ □ 1s 2s 2p 3s	
Al	□ □ □□□ □ □□□ 1s 2s 2p 3s 3p	
Si	□ □ □□□ □ □□□ 1s 2s 2p 3s 3p	
P	□ □ □□□ □ □□□ 1s 2s 2p 3s 3p	
S	□ □ □□□ □ □□□ 1s 2s 2p 3s 3p	

ELECTRON CONFIGURATION OF LARGER ATOMS

- As electrons occupy the 3rd energy level and higher, some anomalies occur in the order of the energy of the orbitals.
- Knowledge of these anomalies is important in order to determine the correct electron configuration for the atoms.
- The following study aid is used by beginning students to remember these exceptions to the order of orbital energies.
- The order of the energy of the orbitals is determined by following the tail of each arrow to the head and continuing to the next arrow in the same manner. Listed below is the order of energy of the orbitals found in this manner:

1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s



ELECTRON CONFIGURATION AND PERIODIC TABLE

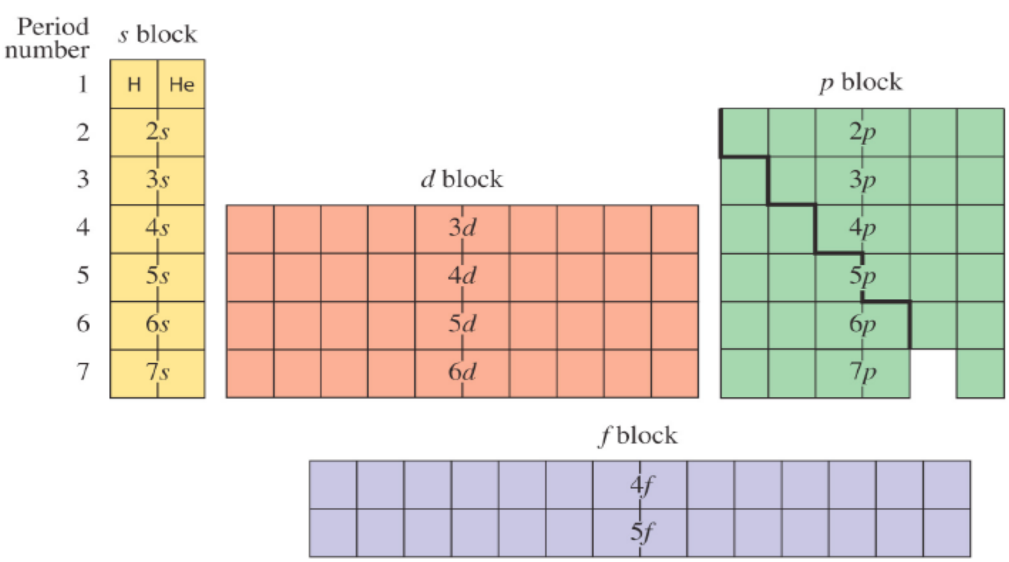
- The horizontal rows in the periodic table are called ***periods***. The period number corresponds to the number of energy levels that are occupied in that atom.

- The vertical columns in the periodic table are called ***groups*** or ***families***. For the main-group elements, the group number corresponds to the number of electrons in the outermost filled energy level (***valence electrons***).

- The ***valence electrons configuration*** for the elements in periods 1-3 are shown below. Note that elements in the same group have similar electron configurations.

IA	IIA	IIIA	IVA	VA	VIA	VIIA	Noble gases
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** $3s^23p^6$

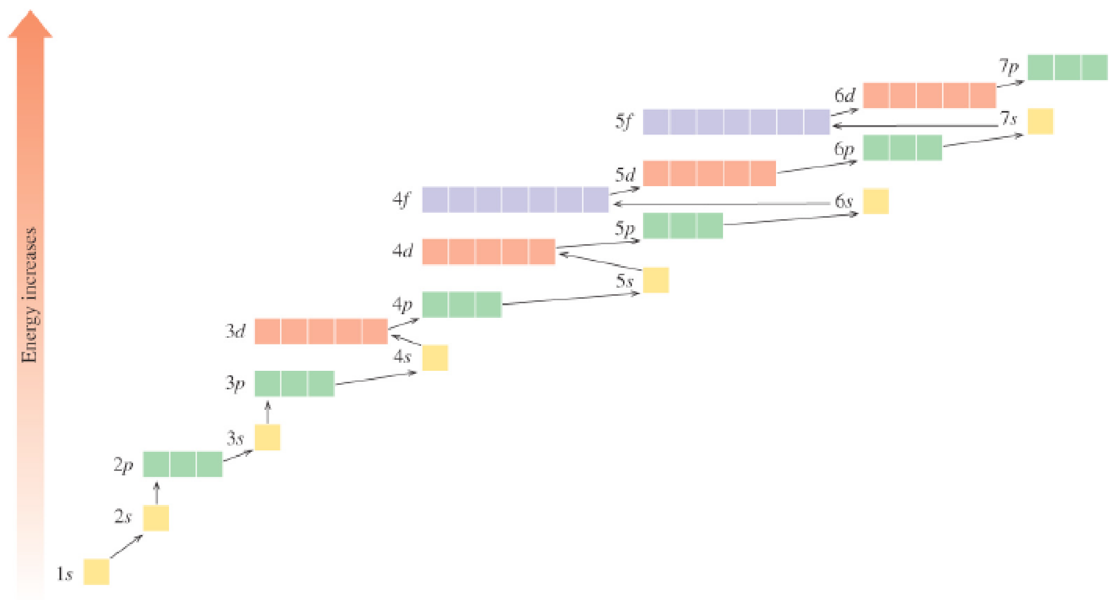
- The location of the different orbital types in the periodic table is shown below:



ELECTRON CONFIGURATION AND PERIODIC TABLE

- The electrons in an atom fill from the lowest to the highest orbitals. The knowledge of the location of the orbitals on the periodic table can greatly help the writing of electron configurations for large atoms.

- The energy order of the sublevels are shown below. Note that some anomalies occur in the energy order of "d" and "f" sublevels.



Examples:

1. Use the periodic table to write complete electron configuration for phosphorus.

   phosphorous, Z =

   electron configuration =

2. Draw an orbital notation diagram for the last incomplete level of chlorine and determine the number of unpaired electrons. Be sure to label each orbital clearly.

   chlorine, Z =

   orbital notation = _____ _____ ______ _____

ABBREVIATED ELECTRON CONFIGURATION

- When writing electron configurations for larger atoms, an ***abbreviated*** configuration is used.

- In writing this configuration, the non-valence (***core***) electrons are summarized by writing the symbol of the noble gas prior to the element in brackets followed by configuration of the valence electrons. For example:

K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$	or	$[Ar]\ 4s^1$
	complete configuration		abbreviated configuration
Br	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$	or	$[Ar]\ 4s^2 3d^{10} 4p^5$
	complete configuration		abbreviated configuration

Examples:

1. Write abbreviated electron configurations for each element listed below:

   a) Fe (Z=26):

   b) Sb (Z=51):

2. Give the symbol of the element with each of the following electron configurations:

   a) $[Ne]\ 3s^2\ 3p^1$

   b) $[Ar]\ 4s^2\ 3d^8$

TRENDS IN PERIODIC PROPERTIES

- The ***electron configuration*** of atoms are an important factor in the physical and chemical properties of the elements.

- Some of these properties include: ***atomic size, ionization energy and metallic character***. These properties are commonly known as ***periodic properties*** and increase or decrease across a period or group, and are repeated in each successive period or group.

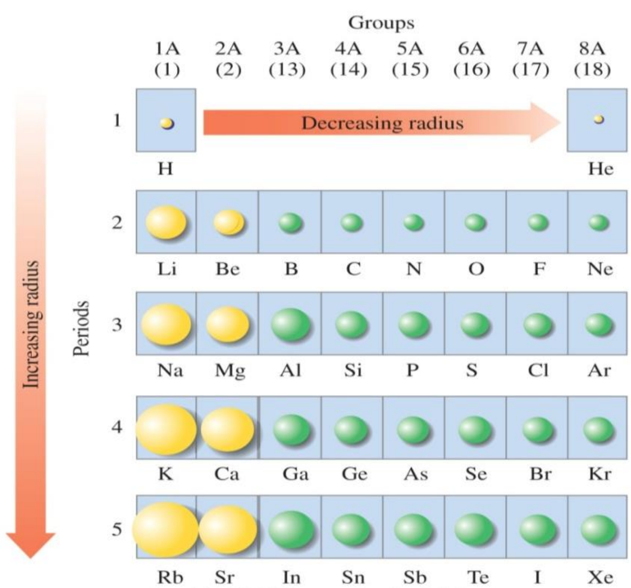
Atomic Size:

- The size of the atom is determined by its ***atomic radius***, which is the distance of the valence electron from the nucleus.

- For each ***group*** of the representative elements, the atomic size ***increases going down the group***, because the valence electrons from each energy level are further from the nucleus.

- The atomic radius of the representative elements are affected by the number of protons in the nucleus (***nuclear charge***).

- For elements going ***across a period***, the ***atomic size decreases*** because the increased nuclear charge of each atom pulls the electrons closer to the nucleus, making it smaller.



TRENDS IN PERIODIC PROPERTIES

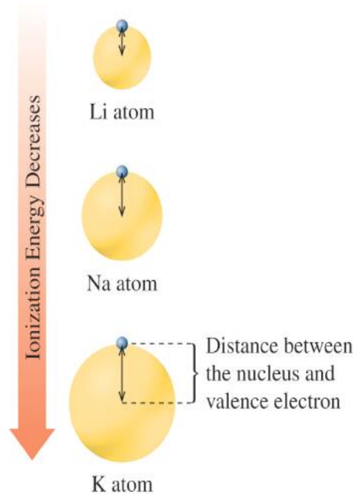
Ionization Energy:

- ***The ionization energy*** is the energy required to remove a valence electron from the atom in a gaseous state. When an electron is removed from an atom, a ***cation*** (+ ion) with a 1+ charge is formed.

$$Na\ (g) + \text{energy (ionization)} \rightarrow Na^{+}\ (g) + e^{-}$$

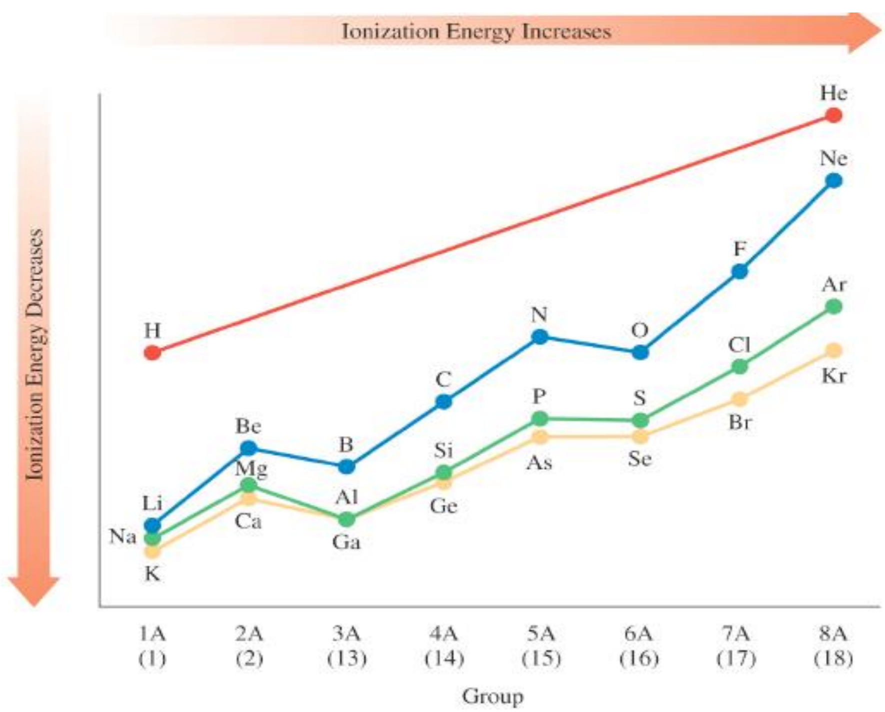
- ***The ionization energy decreases going down a group***, because less energy is required to remove an electron from the outer shell since it is further from the nucleus.

- Going across a period, the ionization energy increases because the increased nuclear charge of the atom holds the valence electrons more tightly and therefore it is more difficult to remove.



- In general, the ionization energy is low for metals and high for non-metals.

- Review of ionization energies of elements in periods 2-4 indicate some anomalies to the general increasing trend. These anomalies are caused by more stable electron configurations of the atoms in groups 2 (complete "s" sublevel) and group 5 (half-filled "p" sublevels) that cause an increase in their ionization energy compared to the next element.

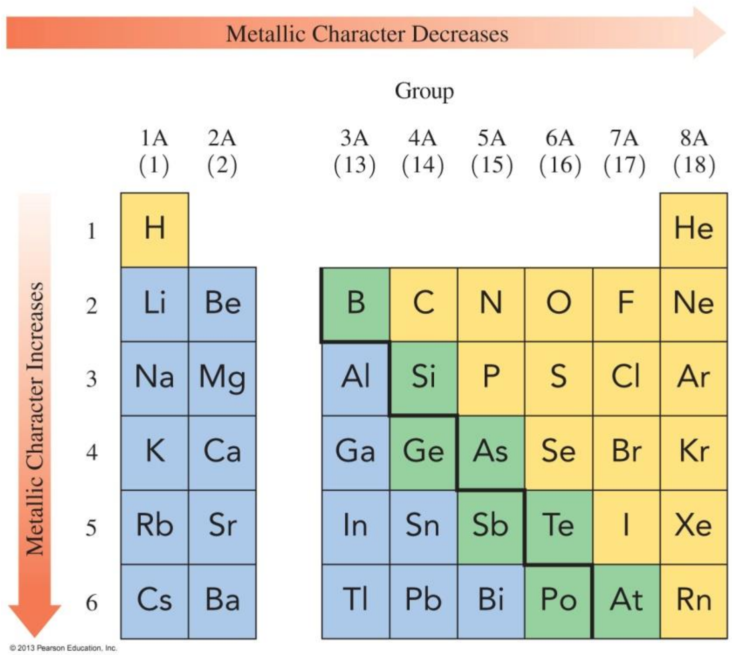


TRENDS IN PERIODIC PROPERTIES

Metallic Character:

- ***Metallic character*** (discussed earlier in this chapter) is the ability of an atom to lose electrons easily.

- This character is more prevalent in the elements on the left side of the periodic table (metals), and decreases going across a period and increases for elements going down a group.



Examples:

1. Select the element in each pair with the larger atomic radius:

   a) Li or K b) K or Br c) P or Cl

2. Indicate the element in each set that has the higher ionization energy and explain your choice:

   a) K or Na b) Mg or Cl c) F, N, or C