

# Modern Atomic Mechanical Theory

## Objectives for Modern Atomic Mechanical Theory

Properly prepared students, in random order, **will**

1. Be able to describe and illustrate the elementary structure of atoms and atomic theory;
2. Know the names and chemical symbols of the first 106 elements and their locations on the periodic table;
3. Be able to determine the names of compounds discussed in the course;
4. Be able to describe and demonstrate the Bohr model of atoms;
5. Be able to find, read and interpret the information on a simple periodic table;
6. Be able to describe, define and apply the properties of metals, non-metals and metalloids;
7. Be able to determine the average atomic mass of elements using the naturally occurring abundances of isotopes in nature;
8. Be able to describe, illustrate, and apply the concepts of ionization and electron affinity (to include first ionization energy) of elements and electrolytes;
9. Be able to explain what cations and anions are and how they are formed at the elementary level;
10. Be able to explain and illustrate fundamental concepts behind the mass spectrometer;
11. Be able to develop/illustrate a simple/crude mass spectrum of a simple compound;
12. Be able to define freezing, melting, boiling and flash points; critical temperature and pressure;
13. Be able to list the names of the groups and periods on the periodic table;
14. Be able to explain and diagram how electrons fill the orbitals of elements/atoms;
15. Be able to explain and apply Hund's Rule to electronic configuration through the first 20 elements on the periodic table;
16. Be able to explain the subdivision of orbitals into subshells;
17. Be able to explain and illustrate applications of Lewis Structures to chemical structures;
18. Be able to describe the characteristics of elements in the representative groups;
19. Be able to explain and describe the properties of the noble gases and atmospheric composition;

20. Be able to explain the importance of using of noble gases to diving;
21. Be able to explain and illustrate light refraction through substances of differing densities;
22. Be able to explain and illustrate wavelengths and frequencies;
23. Be able to explain, describe and diagram discrete, line, spectra;
24. Be able to apply the concept of line spectra to flame tests for ion identification;
25. Be able to explain black body radiation and the photoelectric effect;
26. Be able to explain the significance of the Bohr equation and the keys to its application;
27. Be able to explain, describe and illustrate the Lyman, Balmer and Paschen series of light emission from atoms;
28. Be able to explain the right hand rule and how it applies to generating a magnetic field in elementary terms;
29. Be able to describe, recognize and diagram s, p, d, and f orbitals at the elementary level;
30. Be able to use the four quantum numbers to identify electrons by orbital, shell, shape and spin;
31. Be able to explain and apply the Pauli Exclusion Principle to orbital organization;
32. Be able to recognize, explain and diagram  $sp$ ,  $sp^2$ ,  $sp^3$ ,  $dsp^2$ ,  $dsp^3$ ,  $d^2sp^3$  orbitals;
33. Be able to recognize the difference between  $\sigma$  and  $\pi$  bonds in geometrically correct multi-covalently bonded substances;
34. Be able to explain why oxygen undergoes  $sp^3$  hybridization in water; and why nitrogen undergoes  $sp^3$  hybridization in ammonia;
35. Be able to illustrate and recognize the bond angles between atoms in  $sp$ ,  $sp^2$ ,  $sp^3$  hybridizations;
36. Be able to explain what an ionic compound is and how an ionic compound is formed;
37. Be able to balance simple chemical reactions;
38. Be able to define what a "redox" reaction is;
39. Be able to explain and identify reducing reagents and oxidizing reagents;
40. Be able to identify substances that are reduced or oxidized;
41. Be able to explain what a covalent compound is and how covalent substances are formed;
42. Be able to explain what a coordinate covalent substance/compound is and how it is formed;

43. Be able to define terms related to covalent molecules;
44. Be able to use the  $\Delta EN$  method and eyeball method from memory for determining if a compound is non-polar covalent, polar covalent or ionic;
45. Be able to explain the difference between binary and polyatomic compounds or ions;
46. Be able to determine and write the name, formula and charge on ions;
47. Be able to determine and write perceived charge on atoms in compounds;
48. Be able to explain the difference between electron-rich and electron-deficient;
49. Be able to explain what a reactant and a product are;
50. Be able to explain what a subscript and coefficient are;
51. Be able to list each metal and its ionic form in each analytical group of metals;
52. Be able to explain the difference between paramagnetic, diamagnetic and ferromagnetic;
53. Be able to discuss each metal in each analytical group in detail, e.g., uses and applications;
54. Be able to illustrate a crude Guoy balance that demonstrates paramagnetism and diamagnetism;
55. Be able to construct a flow chart to separate the analytical groups from each other; and to be able to construct a simple flow chart from data obtained in lab or in a problem to separate each ion one at a time from the others until all are separated;
56. Be able to discuss and define the different tooth filling materials;
57. Be able to explain the differences between low, medium and high carbon steels;
58. Be able to discuss the differences between grey, malleable and brittle tin;
59. Be able to explain and diagram the thermite reaction;
60. Be able to explain the difference between malleable and ductile.

and demonstrate that comprehension and applications thereof per assessment tool at no less than a score of 75%.

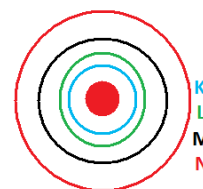
## Atomic Theory

Atoms of the same element have the same average mass. Those of different elements have different average masses. Atom comes from "atomos" which means "not cuttable". Atoms consist of protons, neutrons and electrons:

Atoms		
Protons	Neutrons	Electrons
Positively charged; located in the nucleus of an atom; we write it as either $p$ or $p^+$	Have no charge; found in the nucleus; written as $n$	Negatively charged; in the energy shells/orbitals around the nucleus; written as $e^-$
In an element, the number of protons = the number of electrons; the number of protons is also equal to the atomic number of that element.	Effects atomic mass.	Has no effect on atomic mass; in elements, the number of electrons equals the number of protons.
The number of protons + the number of neutrons = the atomic mass		

When we talk about atoms, the Bohr model, aka the solar system model (right), is the easiest with which to work. It is set up like our solar system with the "sun" being the nucleus and the "planets' orbits" the orbitals for the electrons.

Figure, below, illustrates how to read a simple element from the periodic table. In short, the atomic number is always a non-decimal, whole, number and the atomic mass represents an average mass of each atom. In this example, we are looking at hydrogen.



Nucleus	K	L	M	N
H $1p^+$	$1e^-$			
He $2p^+$ $2n$	$2e^-$			
Li $3p^+$ $4n$	$2e^-$	$1e^-$		
Na $11p^+$ $12n$	$2e^-$	$8e^-$	$1e^-$	
Sc $21p^+$ $24n$	$2e^-$	$8e^-$	$9e^-$	$2e^-$

1 <b>H</b> 1.00797
1.00797 <b>H</b> 1

1 = Atomic Number; 1.00797 = Atomic mass; H = hydrogen

The periodic table is organized into blocks based on the subshell being filled:

- s block:** Groups 1 and 2 (IA and IIA).
- d block:** Groups 3 through 10 (IIB through VIII).
- p block:** Groups 13 through 18 (IIIA through VIIIA).
- f block:** Groups 3 and 4 (Lanthanide and Actinide series).

The table also shows the following labels:

- GROUPS:** 1 through 18.
- PERIODS:** 1 through 7.
- TRANSITION ELEMENTS:** Groups 3 through 10.
- Lanthanide series:** Elements Ce through Lu.
- Actinide series:** Elements Th through Lr.

### In General

Elements on the periodic table below and to the left of the metalloids are metals. Elements on the periodic table above and to the right of the metalloids are non-metals. The dark (heavy) black “stair-step” in the periodic table, above separates the metals from the non-metals; the metalloids make up the “stairs”. The periodic table is sub-divided into groups that have similar properties. To help with this explanation, see also the following: <http://education.jlab.org/itselemental/ele001.html>

### Periodic Table – Metals

Most of the elements are metals. They have high thermal conductivity (they conduct heat readily). Metals have high electrical conductivity (they transmit electricity readily). The metals are ductile (they can be drawn into wires) and malleable (they can be pounded/hammered into thin sheets). They also have a metallic luster (a characteristic "metallic" appearance).

### Periodic Table – Non-Metals

About 20 elements are non-metals. Their properties are typically the opposite that of the metals. They occur as brittle, powdery solids or as gases.

### Periodic Table – Metalloids

Seven elements (depending on the author) are metalloids: B, Si, Ge, As, Sb, Te, At. The properties of metalloids range between those of the metals and the non-metals.

## Elements and Isotopes

Every element consists of a mixture of its isotopes. Isotopes are atoms which are identical in atomic number and electronic configuration, but different in atomic mass. The difference between the two is in the number of neutrons. One example of an isotope is Cl. It has an atomic number of 17. There are two isotopes of chlorine:

Isotope	Number of Neutrons
$^{35}\text{Cl}$	18
$^{37}\text{Cl}$	20

Again, note that the difference is in neutron number. The atomic mass of an element reflects the abundance of all isotopes of that element. There are three times as much  $^{35}\text{Cl}$  as  $^{37}\text{Cl}$ . We can determine the average atomic mass of chlorine in the following manner:

$$3 \times 35 = 105 \text{ for } ^{35}\text{Cl}$$

$$1 \times 37 = 37 \text{ for } ^{37}\text{Cl}$$

$$4 = 142$$

Since there are 3 parts  $^{35}\text{Cl}$  and 1 part  $^{37}\text{Cl}$ , simply multiply each isotope's mass by the ratio observed in nature. Add up both the parts (4) and the total masses due to those parts (142). Divide the total mass (142) by the total parts (4) and the average atomic mass is obtained (35.5).

$$\text{Avg atomic mass} = \frac{142}{4} = 35.5$$

We can do the same thing with  $^{79}\text{Br}$  and  $^{81}\text{Br}$ . Both are present in nature in equal parts. The average atomic mass is determined as follows. The operation is carried out just as before with the chlorine.

$$1 \times 79 = 79$$

$$1 \times 81 = 81$$

$$2 = 160$$

### **Ionization**

In order for most metals and non-metals to be reactive, they must find a way to be charged, i.e., they must either lose electrons (and be positively charged) or gain electrons (to be negatively charged). When this occurs, they are said to ionize. Ions are charged particles. **Cations are positively charged ions.**

$$\text{Avg mass} = \frac{160}{2} = 80$$

**Anions are negatively charged ions.**

When looking at the periodic table and remembering the electron shells from discussions, the periodic table tells you very valuable information about how those elements in the representative group will ionize.

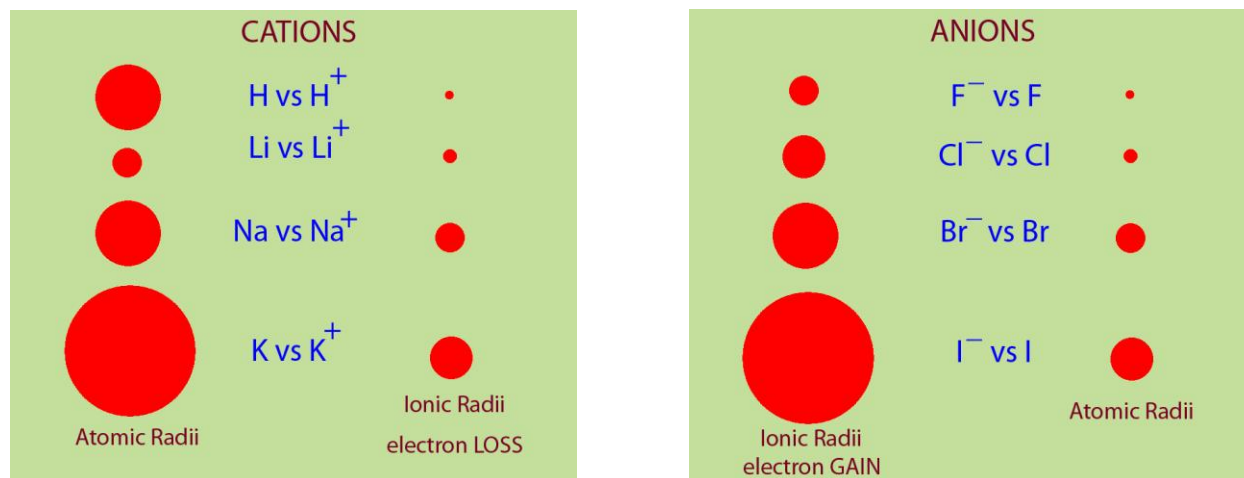
We first have to accept that when an element's **outer most shells** are **filled or are empty** is when that element is **most stable**.

## CATIONS

Elements in Groups I, II and III have less than 4 electrons in their outer shells (specifically, they have 1, 2 and 3 electrons, respectively). In order for these elements to be most stable when they ionize, i.e., to attain the electronic configuration of the previous noble gas, they must lose 1, 2 and 3 electrons, respectively. This will yield charges of +1, +2 and +3, respectively. The positive charges are due to the excess of protons left in the nucleus after the electrons are removed. Cations have smaller ionic radii than their original atomic radii (below).

## ANIONS

Elements in Groups V, VI and VII have more than 4 electrons in their outer shells (specifically, they have 5, 6 and 7 electrons, respectively). In order for these elements to be most stable when they ionize, i.e., to attain the electronic configuration of the following noble gas, they must gain 3, 2 and 1 electrons, respectively. In other words, the outer shell, unlike that of the cations, must fill up with electrons to make a total of 8 electrons in the outer shell. When this happens, the charges obtained are -3, -2 and -1, respectively. The negative charges are due to the accumulation of electrons which out-number the protons in the nucleus. Anions have a larger ionic radius than their original atomic radius (below).



## Electrolytes & Biological Application

Cations and anions are collectively known as electrolytes: charged particles that conduct an electrical current in aqueous solution. Electrolytes are necessary for muscles, brain, heart and nerve tissues to function.

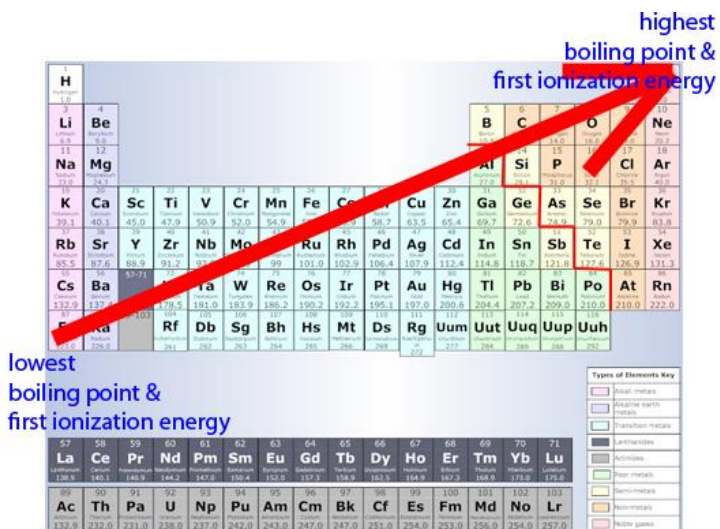
## Group IV Elements

One issue that was not discussed in earlier discussion was that of what happens when the atom is in Group IV? Elements in this group will gain or lose electrons depending on the reaction conditions.

Chemical bonds are strong forces of attraction between reorganized atoms. Oppositely charged particles (ions) attract each other on ionic compounds. Remember that ions form from metal and non-metal atoms.

### Electron Affinity: Ionization, Revisited

Continuing with the theme of ionization, figure, right, illustrates the electron affinity of elements across the periodic table. Electron affinity is the measure of an atom's tendency to gain (or lose) an electron. In general, across periods, the first ionization energy increases from left to right with increasing atomic number.



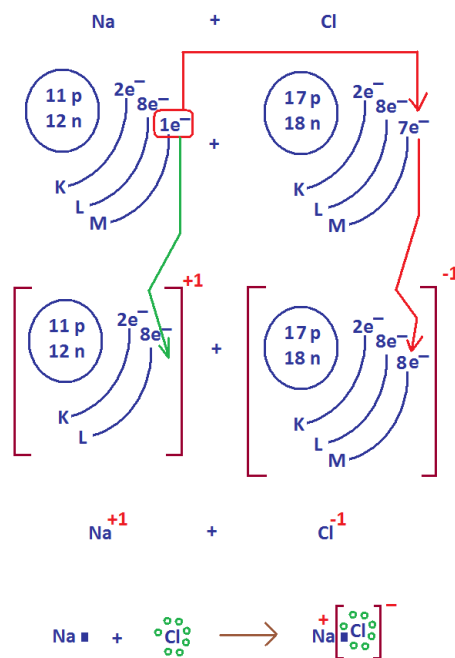
The first ionization energy decreases from top to bottom with increasing atomic number. The first ionization energy is the amount of energy needed to remove the most loosely bound electron from the valence shell of the neutral atom. While there is a small increase in electron affinity from top to bottom of the periodic table, the biggest increase is from left to right.

Electron affinity follows first ionization energies: lower left has the lesser electron affinity while upper right has the higher electron affinity. Fr is the most electro-positive element on the periodic table and has the lowest electron affinity. F is the most electro-negative element on the periodic table and has the highest electron affinity.

What this means is that Fr gives up its outer electron the easiest and F gains its last electron the easiest to attain stability, i.e., the electronic configuration of the previous and the following noble gases, respectively. For practical purposes, this means that elements in low left easily give up their electrons and elements in upper right hang on to their electrons.

### Ionic Compounds

Ionic compounds form when electrons shift between atoms, Figure, right. In this example, elemental sodium reacts with elemental chlorine to form sodium chloride. The movement of electrons is that the sodium loses its single outer electron and chlorine gains it. Note the final Lewis Structure of sodium chloride. This occurs only when the sodium chloride is solid. Otherwise, when this is in water, table salt (sodium chloride)





exists as two distinctly separate ions. Indeed, one does not actually taste NaCl, rather, when NaCl hits the mouth and mixes with the saliva, it dissolves and separates (dissociates) into its two ionic components. Each ion is detected by different portions of taste buds. The detected presence is converted into an electrical signal that is sent to our brains for interpretation. The message we receive following interpretation is that we have "tasted" salt.

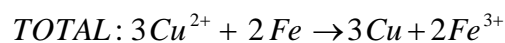
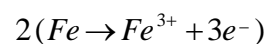
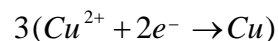
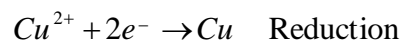
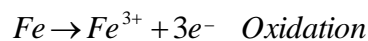
The reactions between magnesium and fluorine and potassium and sulfur, right, respectively, form magnesium fluoride and potassium sulfide. To name these binary compounds (compounds which contain only two kinds of atoms), put

the metal's name first, add the non-metal stem and add "ide". The non-metal stem for fluorine and sulfur are "fluor" and "sulf". These two examples also bring up something else we need to discuss: balancing chemical formulas.

Balancing chemical formulas is not difficult. All of us can balance checkbooks, hence, we can all balance formulas. There are a number of methods by which to do this. Probably the simplest is to remember that for neutral compounds the sums of the charges for ALL of the ions involved must equal zero.

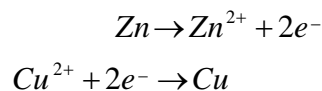


Looking at magnesium fluoride, for example, below, note that it took 2 fluorides to balance the single magnesium. This is because the magnesium ionized to a +2 charge while the fluorine ionized to the -1 charged fluoride ion. Since +2 -1 is +1, it follows that there has to be 2 fluorides to give a total of -2 for the negative charges, i.e. +2-2 = 0 for a neutral compound. The same thing applies to the potassium in potassium sulfide. In this case, two potassium atoms are required to balance out the -2 charge on the sulfide ion.



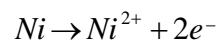
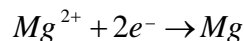
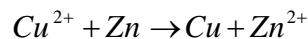
Thus far, these two specific reactions are examples of what are called "redox" reactions. Redox reactions are oxidation (a reactant loses electrons) and reduction (a reactant gains electrons) reactions. If there is reduction, there MUST be oxidation -- i.e., ya can't have one without the other. Each part of a redox reaction may be written separately as a half reaction. When the oxidation half reaction and the reduction half reaction are added up and adjusted

(balanced) for electron exchange, the complete redox reaction is obtained. Note that when the electrons are not equal between each half reaction that they must be made equal -- this includes multiplying the whole half reaction by the exact same number you multiply the electrons by.



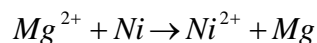
Whenever you work with redox reactions, there are 4 things to remember:

**1) the oxidizing agent is always reduced; 2) the reducing agent is always oxidized; 3) the oxidizing agent gains electrons; 4) the reducing agent loses electrons.** Additional redox examples are illustrated at right.



### How Do We Know About (Or Measure) the Presence of Isotopes?

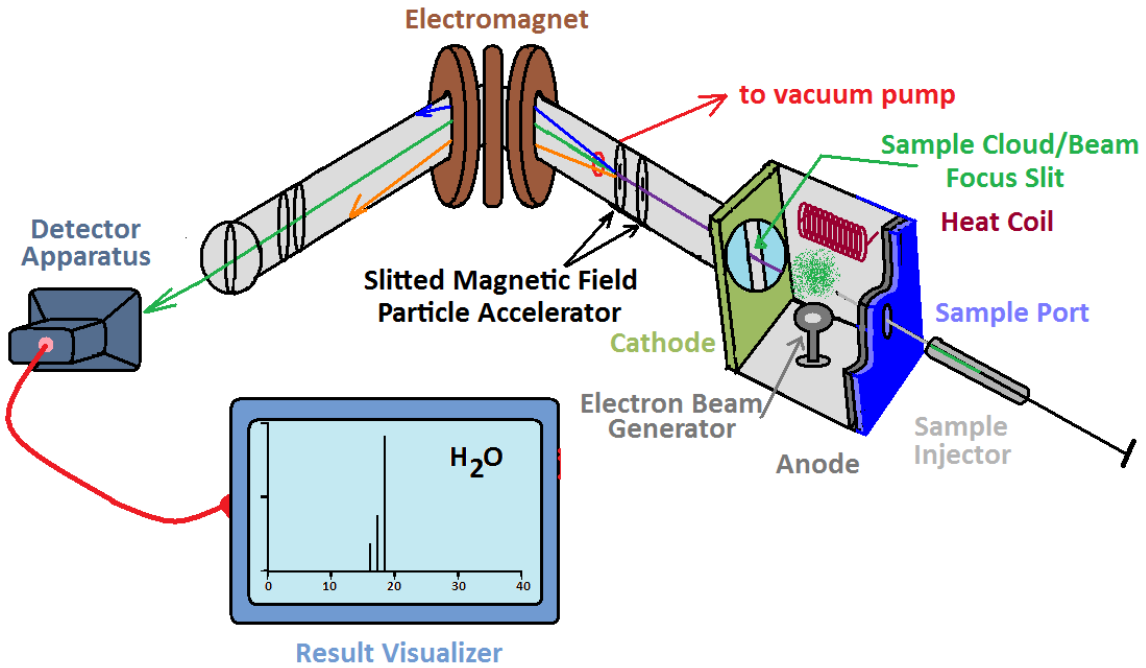
Isotopes (be they radioactive or not) are relative new-comers to Chemistry and Physics. Isotopes were first studied following the initial development of the mass spectrometer by Wilhelm Wien in 1899, improved upon by J.J. Thomson ca 1912. Ca 1912, Thomson proved the presence of the first isotopes known to man:  $^{22}\text{Ne}$  and  $^{20}\text{Ne}$ . The concept/existence of "isotopes" was first proposed by Frederick Soddy, another eventual Nobel Prize winner.



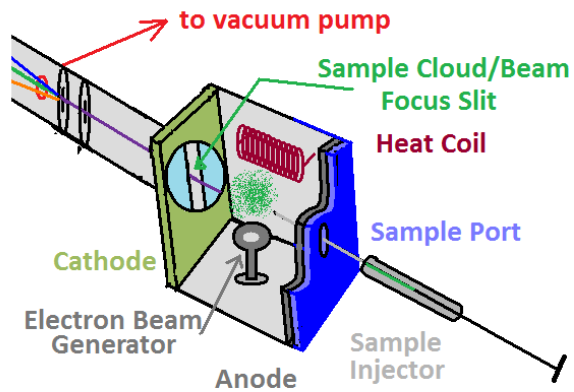
### Why Discuss Mass Spectrometry?

Mass Spectrometry makes determining chemical formulas and structures fairly simple. Mass Spectrometry makes determining isotope presence and abundance fairly simple.

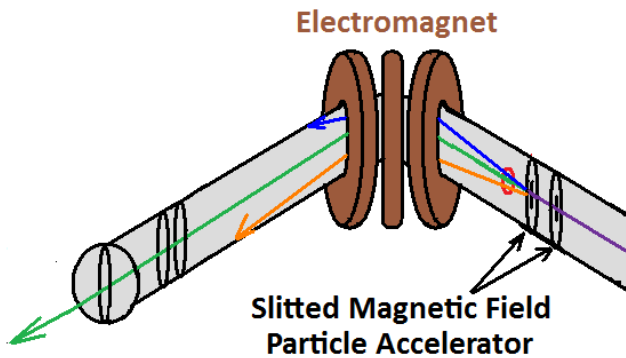
**Mass Spectrometer Layout:**



Sample is injected into the sample holder/chamber. Sample is vaporized by heat coil. Sample is bombarded by electrons from electron beam generator (some folks call this an electron “gun”). Excess electrons boil off generator and are attracted to the Anode (positively charged). Remainder of electrons interact with (bombard) the sample cloud to generate sample cations. Cations are attracted



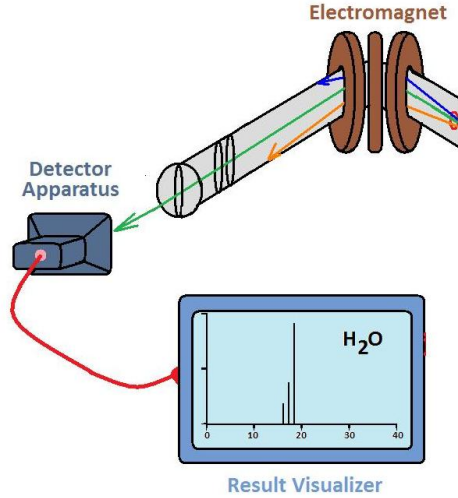
to the cathode (negatively charged) and are partially focused into a beam to enter the MS Tube.



The sample beam travels under vacuum to the slitted magnetic field particle accelerators – further focuses the beam of the sample cations and prepares the beam to “slingshot” around the tube bend. As the beam

approaches the electromagnet, the magnetic field causes the beam to split into “three” beams: Heavy beam (blue beam); “Correct beam” (green beam); Light beam (orange beam)

“Correct beam” of fragments and/or molecules and/or ions hits the detector apparatus where it’s digitized for read-out on the result visualizer. The mass spectrum on the visualizer is that of water: tallest peak is 18/1 (m/z; for water, itself); second highest is 17/1 (for OH fragment); third highest is 16/1 (for atomic oxygen).

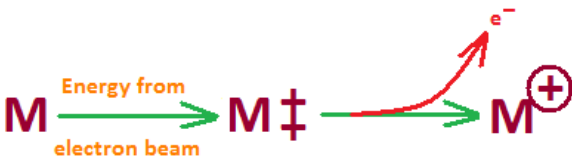


**Mass Spectrometry: Some Examples**

The mass spectrum identifies molecular mass and molecular particle/fragment masses of a parent compound. When a mass spectrum is visualized, it is a “bar/line” graph set on a set of axes as at upper right. The mass spectrum for water, BTW, is lower right of upper right image for an illustration of the “bar/line” graph. Newer MS’ seem, in some instances, to be using “peaks” in some instances, rather than “lines”. The vertical axis is Relative Abundance – usually the highest peak is given the value of “100”. The horizontal axis is the mass/charge ratio, often expressed as m/z or m/e (immediate right image).



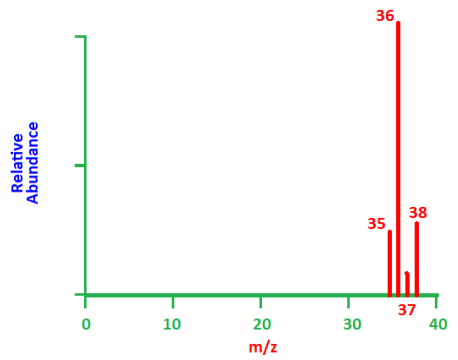
How does one obtain the mass spectrum of a compound or a mixture of isotopes, barring the techno-mechanical portion of this activity?



“Circle +” = radical ion

**Mass Spectrometry Used to Determine Abundance of Isotopes in Nature**

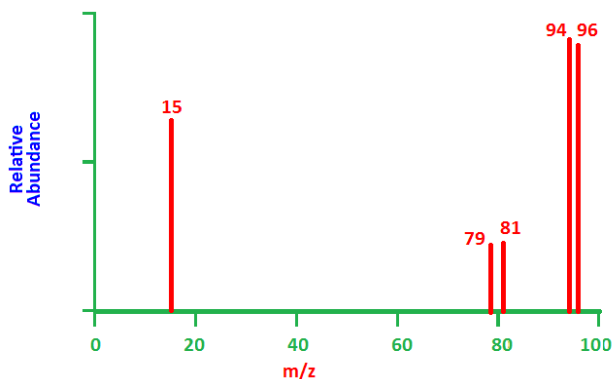
Sample at bottom right contains HCl – what’s the proportion of <sup>35</sup>Cl to <sup>37</sup>Cl in nature?



35 peak is  $^{35}\text{Cl}$ ; 37 peak is  $^{37}\text{Cl}$ ; Ratio of peak heights = 3:1, hence 3X as much  $^{35}\text{Cl}$  as there is  $^{37}\text{Cl}$ ; 36 peak is from  $\text{H}^{35}\text{Cl}$ ; 38 peak is from  $\text{H}^{37}\text{Cl}$ ; 35 peak called "A" peak; 37 peak called "A+2" peak since is 2 mass units heavier/beyond the 35 peak.

### Mass Spectrometry Used to Determine Abundance of Isotopes in Nature

Sample contains methyl bromide (bromomethane) – what's the proportion of  $^{79}\text{Br}$  and  $^{81}\text{Br}$  in nature? 15 peak is  $\text{CH}_3$  peak; 79 and 81 peaks in almost equal amounts – big clue; Hence almost equal – one can actually calculate the numbers based on the actual peak heights;  $94 - 79 = 15$  for the methyl;  $96 - 81 = 15$  for the methyl. Since the 94 and 96 peaks differ only by methyl, the peak heights are also another clue as to the abundance of the two isotopes.



### Review – Update – Terms to Know

**Freezing Point:** is the temperature at which a liquid becomes a solid at normal atmospheric pressure.

**Melting Point:** is the temperature at which a solid becomes a liquid at normal atmospheric pressure.

**Boiling Point:** the temperature at which a pure solvent's or solution's vapor pressure = atmospheric pressure. The properties of the elements are a periodic function of their atomic numbers. In general, across periods, boiling point increases from left to right with increasing atomic number. Boiling point decreases from top to bottom with increasing atomic number.

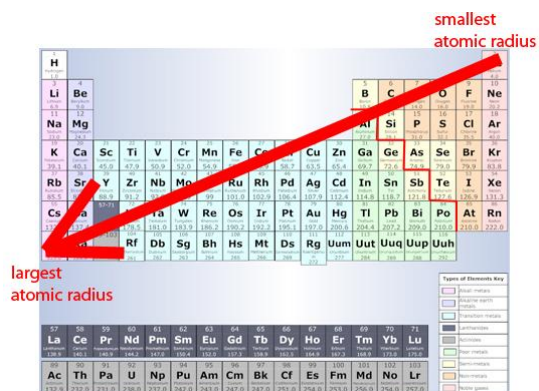
**Flash Point:** is the lowest temperature at which a liquid can form an ignitable mixture in air near the surface of the liquid. The lower the flash point, the easier it is to ignite the material.

**Critical Temperature:** is the temperature above which a substance can not be liquified REGARDLESS how much pressure is applied. **Critical Pressure:** is the pressure required to liquify a gas at its critical temperature.

Common Substance	Critical Temperature (K)	Critical Pressure (atm)	Common Substance	Critical Temperature (K)	Critical Pressure (atm)
$\text{H}_2 - ^1$	33.24	12.8	$\text{NH}_3$	405.5	111.5
$\text{N}_2$	126	33.5	$\text{H}_2\text{O} - ^2$	647.1	217.7
$\text{O}_2$	154.3	49.7	$\text{SO}_2$	430.3	77.7
$\text{CO}_2$	304.2	73.0			

<sup>1</sup> -- has weak intermolecular forces; <sup>2</sup> - has high intermolecular forces. Note the differences these forces make between critical temperatures and pressures between the two related substances.

The atomic radius increases from right to left and top to bottom of the periodic table with those elements in the lower left having the greatest atomic radius. Those elements in the upper right of the periodic table have the smallest atomic radii.



Group I	Group II	Group VII	Group VIII
Alkali metals	Alkaline earth metals	Halogens	Noble (Inert) gases
These react vigorously with water to form an alkaline or caustic solution.	These are found in "earthy" substances like limestone.	React to make salts.	These are primarily unreactive. Some have been shown to be reactive -- that's another course.

Groups I through VIII are the representative elements. The three-row cluster interrupting the representative elements are the transition elements. Groups run vertically on the periodic table and are marked with Roman Numerals. Periods run horizontally across the periodic table and are marked with Arabic numerals.

Substances may be sub-divided into two classes: Molecular compounds and Ionic compounds. These are summarized in the table, below:

Molecules	Ions
Electrically neutral particles made up of 2 or more atomic nuclei surrounded by enough electrons to make the tiny package neutral.	Oppositely charged particles which are attracted to each other.

## Part II

### Electronic Configuration of Atoms

We start with the nucleus and work outward from there. Each element has a unique electron configuration. It is this uniqueness that determines the chemical properties of each element. Every element has a different electron arrangement that requires placement of electrons from the inner shell thence to the next shell, thence to the next shell.

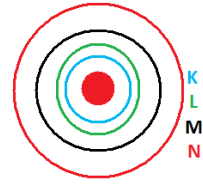
The innermost shell (K shell) must fill first. It holds a maximum of 2 electrons. Once the K shell is filled, the L shell begins to fill.

The L shell holds a maximum of 8 electrons. When the L shell is filled, the M shell begins to fill.

The M shell holds a maximum of 8 or 18 electrons. The reason for this will become apparent in a short while. Once the M shell is filled, the N shell begins to fill.

The N shell holds a maximum of 8 or 32 electrons. As with the M shell, this will become clearer in a bit.

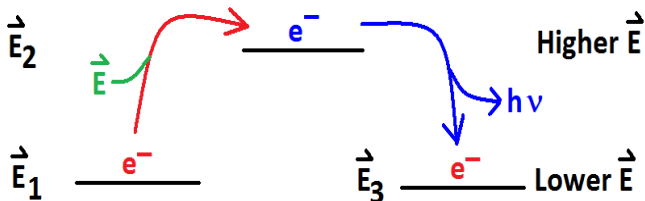
These energy shells (I use orbitals interchangeably) do not explain, though, how electrons move between energy levels. When electrons become excited, they "jump" to a higher level. As they relax, they give off energy in the form of heat or light energy ( $h\nu$  in the graphic).



Nucleus	K	L	M	N
H $1p^+$	$1e^-$			
He $2p^+$ $2n$	$2e^-$			
Li $3p^+$ $4n$	$2e^-$	$1e^-$		
Na $11p^+$ $12n$	$2e^-$	$8e^-$	$1e^-$	
Sc $21p^+$ $24n$	$2e^-$	$8e^-$	$9e^-$	$2e^-$

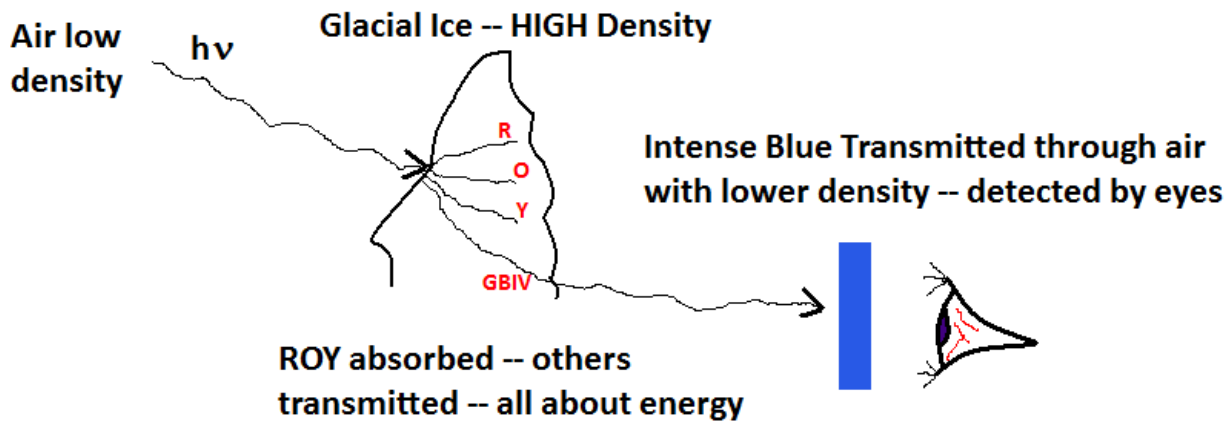
	maximum electrons orbital can hold
N _____	8 or 32
M _____	8 or 18
L _____	8
K _____	2

Shells or Orbitals



How do we know that the electrons give off this light energy?

## Light, Atoms and Energy

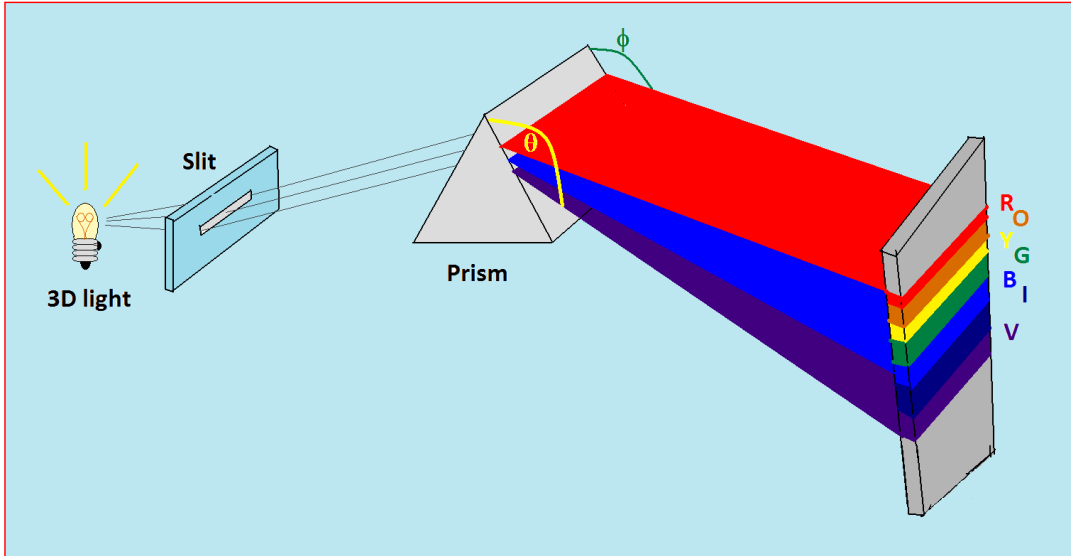


How is it that light does this, i.e., how is it possible that we may separate out individual colors from "white" light? Prisms or tiny slits or gratings "split", bend (refract) polychromatic light into mono-chromatic light. Another example of how this occurs is in the graphics, above (glacier ice refracted to permit only blue light to be transmitted) and below:



The Light from the sun we see as "white" light. This light may be "split", though, into many colors, of which 7 are primary: Red, Orange, Yellow, Green, Blue, Indigo and Violet, or the colors of the rainbow. An easy way to split light color components from white light -- visible light -- is to pass it through a prism, top next page:

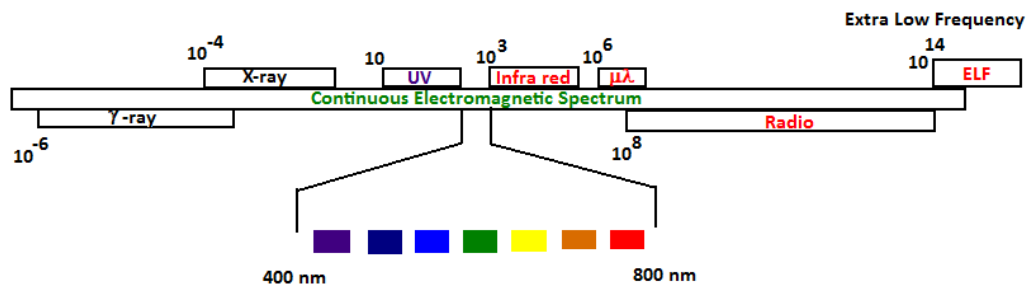




This is from the visible portion of the electromagnetic spectrum. Note that the angle,  $\phi$ , is less than the angle,  $\theta$ . Red light is bent the least and blue light is bent the most. The blue light has the most energy. Examples of "light" from the continuous electromagnetic spectrum are tabulated below:

Characteristic	"Color"
Warmth; night vision goggles	IR
Eyes, film	Visible light
Sun burned skin	UV
Radios	Radio waves

The next graphic lays out an overview of the electromagnetic spectrum:



The center region of the graphic has been exploded so you may see how the visible spectrum fits into the scheme of things. Remember, too, that there are shades of varying colors between each specific primary color (from red-red-red to red-red-orange to red-orange to red-orange-orange to orange-orange-orange): 700 nm: RRR → RRO → RO → ROO → OOO: 600 nm

Note that different densities also effect the "splitting of light". (The visible spectrum, BTW, spans 400-800 nm, more or less.) These tools "split" the light because light has dualistic properties. Planck and Einstein showed that in an elegant manner when they each developed their own equations to explain electromagnetic energy:

$$E = h \nu \text{ and } e = m c^2$$

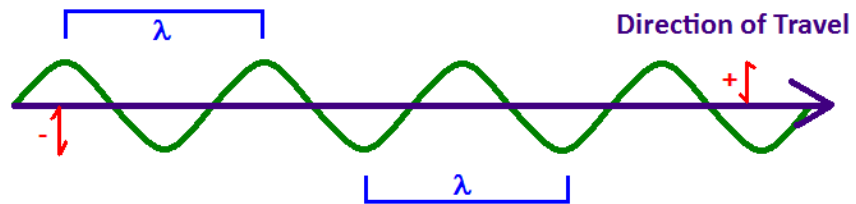
The former equation is Planck's equation that describes the wavelike properties of electromagnetic energy/radiation and the latter is Einstein's equation that describes the particle like properties of electromagnetic energy/radiation. When the two equations are combined:

$$h \nu = m c^2$$

which says that electromagnetic energy/radiation has mass and acts like a wave.

A wave is described as a progressive, repetitive motion that moves from a point of origin to farther points. We've all seen what happens when a pebble is dropped in a pond: ripples form and spread out from the point of origin (where the pebble hit the water). The ripples are waves -- light -- Electromagnetic energy -- moves in the same way. If we look at electromagnetic energy (EME), the wave form looks like this:

Where the distance between 2 peaks OR two valleys is the wavelength:  $\lambda$ .



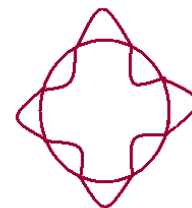
The EME wave forms originate from the movement of electrons with respect to the atomic nucleus. The electron-ic movement produces oscillations in the electric/magnetic fields that are propagated over the electron's orbit:



The circular orbit, overlapped by the wave-like oscillations is based on the hydrogen atom by Bohr in 1913. Keep in mind that as the electron is moving it is inducing a magnetic field perpendicular to the direction of its flow. (Remember the electromagnets you made as a kid with a battery, some wire and a nail? and the right hand rule?)

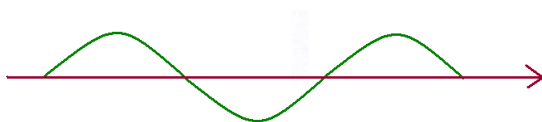


Induced Electromagnetic Field; Perpendicular to electron flow

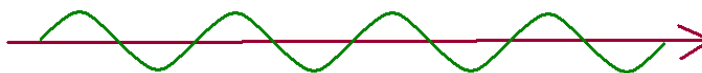


Electron-ic oscillations around orbit

Wavelengths may be far apart:



Wavelengths may be close together:



The closer together, the greater the frequency ( $f$  or  $\nu$ ).

The frequency is defined as the number of wavelengths that pass a point in time, e.g., if  $4\lambda$  pass the same point in 0.25 seconds, then the frequency is equal to:

$$\nu = \frac{4\lambda}{0.25 \text{ sec}} = 16 \text{ Hz}$$

Hz = cps or is denoted as  $s^{-1}$ . The farther apart the less the frequency, e.g., if  $4\lambda$  pass a point in 2 seconds, then the frequency is equal to:

$$\nu = \frac{4\lambda}{2 \text{ sec}} = 2 \text{ Hz}$$

Since frequency is in units of reciprocal seconds ( $s^{-1}$ ) and wavelength ( $\lambda$ ) is in distance units (m in SI -- nm, for practical purposes), the product of the two is in units of m/s or velocity:

$$\lambda * f = c \text{ OR } \lambda * \nu = c$$

"c" is a special number: the speed of light in a vacuum, which is  $3.0 \times 10^8$  m/s. Notice that both  $f$  and  $\nu$  are equal to the quotient of  $c/\lambda$ , where, again,  $f$  and  $\nu$  are frequency in Hz, cps or  $s^{-1}$ .

Concepts to remember:

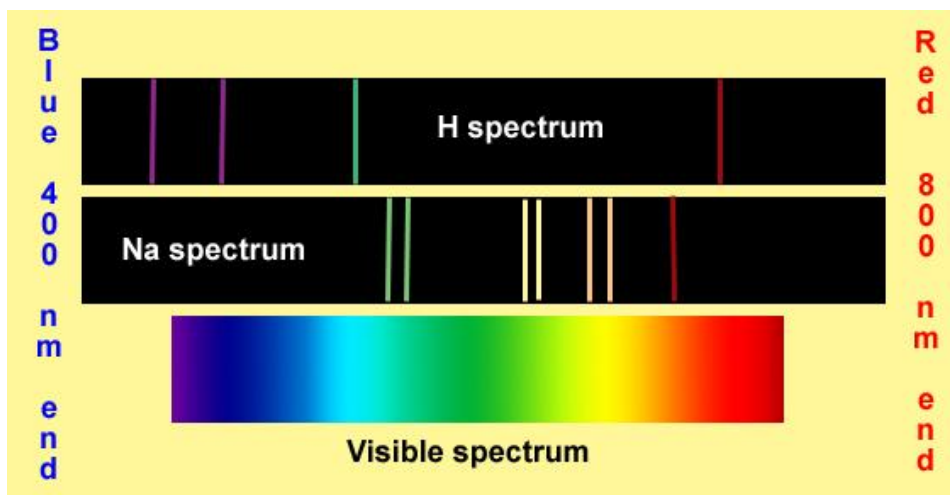
$\lambda$	$f$	color	energy
Long	Small	Red end	Low
Short	Large	Blue end	High

The red end may be likened to the bass end of the musical scale; the blue end may be likened to the treble end of the musical scale.

Why are atoms important in understanding colored light? Among other reasons, raindrops split visible sun light into a familiar continuous spectrum we know as a rainbow (angle of refraction is between 40 and 42 degrees (blue vs red, respectively)).

Specific elements, however, do not give continuous (rainbow) spectra when light released from a gas discharge tube (a light "bulb" with a specific element in it) is sent through a slit and viewed. The light emitted by the gas discharge tube consists of discrete wavelengths (colors) of light, e.g., light from a hydrogen discharge tube appears fuschia and consists of light of 4 specific wavelengths: 410.1 nm, 434.0 nm, 486.1 nm and 656.3 nm or violet, violet, indigo and carmine (next slide). When viewed, this spectrum has ONLY 4 LINES of light (at the above wavelengths) -- this set of spectral data is called a LINE spectrum or an emission spectrum. The line spectrum of an element is its fingerprint and is unique to itself.

## Continuous and Discrete (Line) Spectra



Some elements can be forced to emit light quite simply by placing them in a fire (bunsen burner) and allowing them to burn, e.g.,

Element	Flame	Element	Flame
Na	Bright yellow	Rb	Red
Sr	Carmine	Ba	Green
K	Violet	Cu	Greenish blue
Cs	Blue		

These are called flame tests and are used to qualitatively identify compounds containing these elements. We shall see later why the line spectra of the element and its ion are identical.

### Black Body Radiation

In the 19th century, classical physics (includes mechanics, thermodynamics, electricity, magnetism and light) failed to explain line spectra.

People watched "red hot" poker emit red light; when the temperature was raised to 1200° C, the poker emitted white light as more yellow and blue light were emitted and mixed with the red light. This sort of radiation that depends on the solid's temperature and NOT on what elements make up the solid is called black body radiation. To explain black body radiation, Max Planck in 1900 developed a revolutionary theory to cover all aspects of black body radiation: the energy of vibrating atoms in a system of unique atoms is fixed. EME emitted by these vibrating atoms corresponds ONLY to the difference between 2 permitted energy levels.

The quantum, or smallest packet, of energy that can be emitted is expressed by Planck's equation,  $E = h \nu$ , where

$E$  = energy emitted

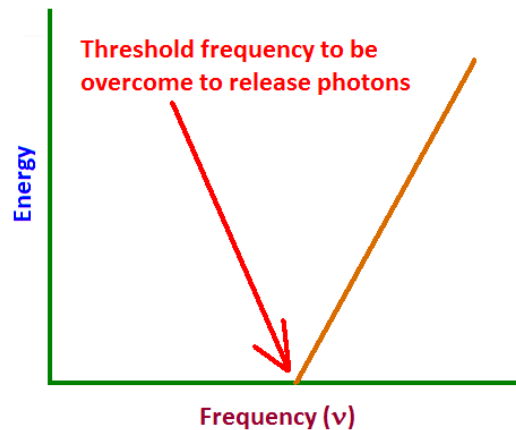
$h$  = Planck's constant

$\nu$  = frequency of radiated light

emitted light is emitted as a single quantum that has energy that is EXACTLY an integral multiple of the simplest quantum (packet):  $h \nu$ , i.e.,  $h \nu$ ,  $2h \nu$ ,  $3h \nu$ ,  $4h \nu$ ,  $5h \nu$ ,  $6h \nu$ ,  $7h \nu$  and NOT  $0.5 h \nu$ ,  $2.4h \nu$ ,  $3.7h \nu$ ,  $7.65h \nu$ .

A useful analogy to understanding this is to think of vending machines, stamp machines that increase in exact nickel or quarter increments and NOT in penny or 3 cents increments.

Einstein expanded on Planck's quantum theory to explain the photoelectric effect. The photoelectric effect is defined as occurring when a beam of light that is shining on certain metals causes the release of a beam of electrons. The photoelectric effect is dependent on the frequency (color) of the light: weak blue light releases photoelectrons with higher energy than does bright red light. In addition, if the frequency is less than a minimal value (defined as the threshold frequency) no photoelectric effect is observed:



When the photoelectric effect occurs, the release of light is quanta called photons.

Analogy: VW stuck in mud: lot's of people can push one at a time to no avail -- a tractor can pull it out and give extra energy (kinetic energy) to it. Planck's equation allows us to also calculate the energy of these photons.

The importance of the structure of atoms as applied to line spectra became apparent in 1913 when Bohr discovered how to explain the line spectrum of hydrogen. He first derived an equation to explain the energy of electrons [ $E_n$ ]. Each energy value ( $E_1$ ,  $E_2$ ,  $E_3$ ,  $E_4$ ,  $E_5$ ,  $E_6$ , ...) is called an energy level.

$$E_n = \frac{-B}{n^2}$$

The ONLY permitted values for  $E_n$  are, below:

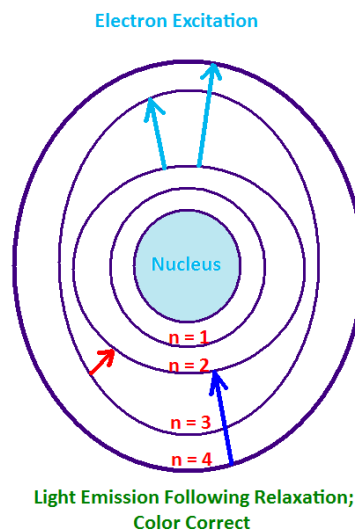
Where  $E_n$  = electron energy at level "n"

$B$  = a constant based on Planck's constant and the mass and charge of an electron; =  $2.2 \times 10^{-18} \text{J}$

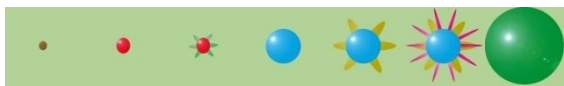
$N$  = an integer and corresponds to the primary quantum number; may also = "n". More on this later.

The negative sign on B (-B) is there to remind us that energies of attraction are negative (this is due to nuclear and electron co-attraction).

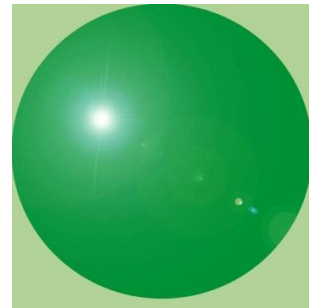
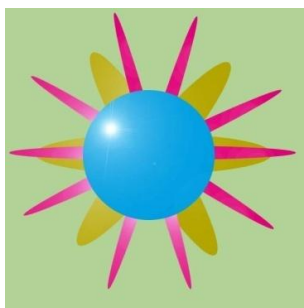
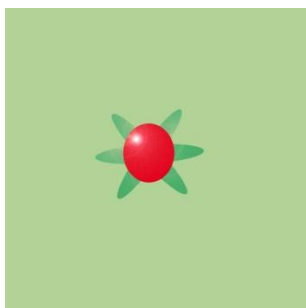
Bohr felt that the hydrogen atom was similar to the solar system meaning that all energy levels were present with electrons in a specific level on numerous quantum mechanical explanations, image right:



“n” in the graphic stands for energy level or orbital. And the 2-D illustrations represent the layout of overlapping subshells from nucleus outwards to outer shells:



Starting above left: 1s, 2s, 2s2p, 3s, 3s3p, 3s3p3d, 4s;



Larger images: above, left to right, 2s2p, 3s3d3p, 4s

In order to understand line spectra, we must look at energy changes or differences ( $\Delta E$ ):

$$\Delta E = E_f - E_i$$

$\Delta E$  = difference in energy between 2 states

$E_f$  = final energy state

$E_i$  = initial energy state

We can apply this concept to the Bohr equation, as well, right.

$$\Delta E = \frac{-B}{n_f^2} - \frac{-B}{n_i^2} = B \left( \frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$$

KEY:

If  $n_f > n_i$ : the electron ABSORBS energy and goes from a lower level to a higher energy level (also called an excited state)

If  $n_i > n_f$ : the electron releases energy and goes from a higher level to a lower level (gives off light at specific frequencies and, hence, colors)

In ALL movements (transitions) electrons move all at once, i.e., they move from  $n_i$  to  $n_f$  ALL AT ONCE without stopping at  $n_x$  in between the two ( $n_i$  and  $n_f$ ). Every movement to a lower level causes a spectral line. All the spectral lines give the emission spectrum for each element.

E.g., Calculate the difference in energy in Joules, that occurs when an electron falls from  $n=6$  to  $n=2$  level in a hydrogen atom.

$$\Delta E = B \left( \frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$$

$$\Delta E = 2.2 * 10^{-18} \left( \frac{1}{6^2} - \frac{1}{2^2} \right)$$

$$\Delta E = 2.2 * 10^{-18} (-0.2222) = -4.86 * 10^{-19} J$$

Calculate the frequency in Hz of the EME released by the above electron energy-level change. Calculate the wavelength of this light, as well.

First, lose the negative sign on the energy -- it's not useful, here

$$E = h \nu$$

Now, solve for  $\nu$  :

$$\nu = \frac{E}{h} = \frac{4.89 * 10^{-19} J}{6.63 * 10^{-34} Js} = 7.38 * 10^{14} Hz$$

$$\lambda = \frac{c}{\nu} = \frac{3.0 * 10^8 m/s}{7.38 * 10^{14} Hz} * \frac{10^9 nm}{1m} = 406.5 nm$$

This is at the "edge" of the violet portion of the spectrum. This is one line of a complete line spectrum for hydrogen.

When discussing line spectra -- particularly of that Nobel-prize-winning element, hydrogen -- it is necessary to understand that emission spectra consist of several SERIES of lines. The usual series observed are IR, visible and UV, image, top next page.

Bohr's work only explains the single electron atom, H, and does not work with atoms that are larger than H.

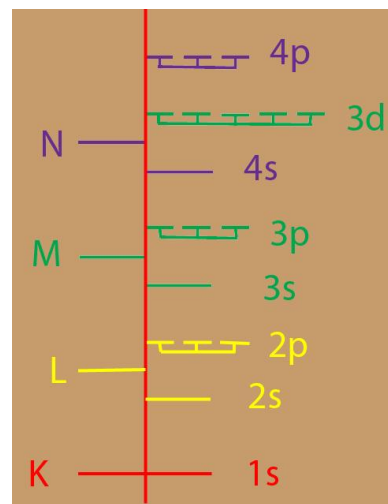
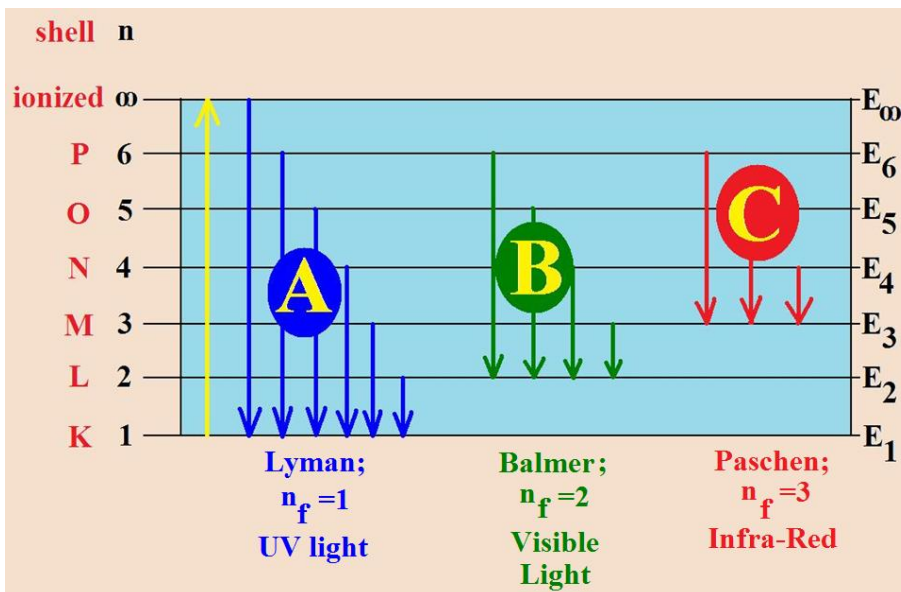
It does, however, tell us that, although not all are occupied by electrons, all energy levels are present in all atoms.

Knowing that electron movement gives us "color", there must, therefore, be another way to examine this atomic phenomenon.

It turns out that each energy shell (orbital) has sub-shells within it, much like how trees grow in layers or laminates, with one exception: the K shell. The K shell has one and only one subshell in it. It is called the 1s subshell. **All s subshells hold a maximum of only 2 electrons.** Once the 1s subshell is filled (at element #2, He), the next electron has to go into the L shell.

The L shell consists of 2 subshells: the 2s and the 2p subshells. The third electron goes into the 2s subshell. Once the 4th electron has filled the 2s subshell, the 5th electron has to go into the 2p subshell. Here's where it gets a bit messy. The 2p subshell consists of 3 sub-subshells. All subshells and sub-subshells hold a maximum of 2 electrons. Since the 2p subshell consists of 3 sub-subshells, it actually holds a maximum of 6 electrons. Remember that **the L shell holds a maximum of 8 electrons.**  $2 + 6 = 8$ . The 8 electrons come from the two 2s and six 2p electrons when that shell is filled. When the last 2p electron is in place (element 10, Ne), the 11th electron must begin to fill the M shell.

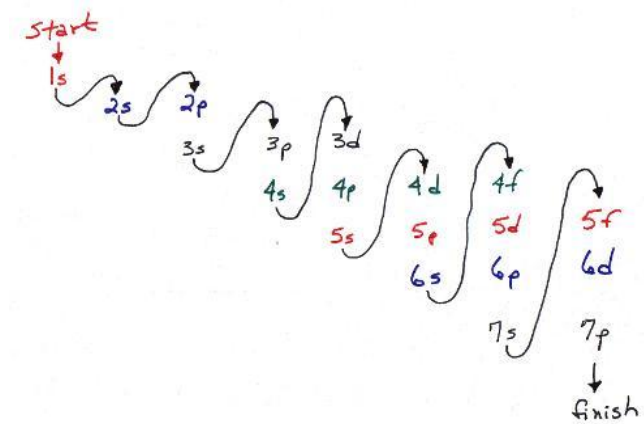
The M shell, remember, is sort of strange: **it holds a maximum of 8 or 18 electrons.** How does it do that? The M shell consists of the 3s and 3p subshells. A third subshell is also available as necessary: the 3d subshell. We already know that s subshells hold a maximum of 2 electrons and that p subshells hold as maximum of 6 electrons by virtue of having 3 sub-subshells, apiece. What about the 3d subshell? All d subshells hold a maximum of 10 electrons. Since electrons prefer to pair up, that tells us that there are 5 sub-subshells in any d subshell. Since we already know that 8 electrons come from the 3s and 3p subshells, and we now know that there may be 10 more from the 3d subshell, it follows that  $8 + 10 = 18$ .





Even though the 3d shell is a part of the M shell, energy-wise, it begins to fill AFTER the N shell (4s subshell) begins to acquire electrons. You'll see why in a bit.

The N shell, remember, **may hold 8 or 32 electrons**, maximally. It may hold this many electrons because



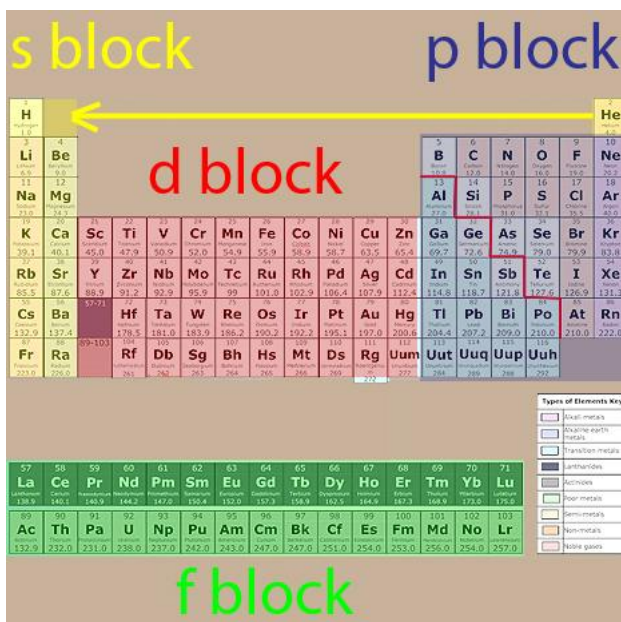
not only does it contain 4s, 4p and 4d subshells, it also contains a 4f subshell. The total number of electrons that an f subshell may hold is 14. Again, since electrons like to pair up, this says that there are 7 sub-subshells in the 4f subshell. Hence,  $2 + 6 + 10 + 14 = 32$  total electrons.

Groups I and II	Transition elements	Groups III through VIII	Lanthanides and Actinides
These elements make up the "s" block elements.	These elements make up the "d" block elements.	These elements make up the "p" block elements.	These elements make up the "f" block elements.

Once we look at the periodic table, sub-shell filling becomes elementary. The subshells have a specific order in which they fill. The periodic table is, bar none, the best "cheat sheet" any chemist has available to him or her.

This means that every element in a specific Group has electrons, successively, in a specific subshell that is easily identifiable. In terms of periods, this is even easier.

Off the left of the periodic table are Arabic numbers that run from top to bottom, starting at 1. Since periods go horizontally across the periodic table, every element in a specific period starts with that number. Let's look at some examples.



Hydrogen is the first element on the periodic table. It is in Group I and period 1. Put the period first and follow it with the block letter: 1s. That means that this element has an electron in the 1s subshell. We also have to put in how many electrons there are. This is done by inserting the number in

the subshell as a superscript. Since there is only one electron, the electronic configuration of H is  $1s^1$ . Again, this says that there is one electron in the 1s subshell.

Let's look at helium, now. He is in period 1. It is element number 2. It's electronic configuration is  $1s^2$  -- period 1, s block and has 2 electrons. The 1s subshell is now full. Note: He is the exception to the p block elements -- it has no p electrons.

The next element on the periodic table is Li. Note that it is in group I and period 2. In order to write the electronic configuration for Li, you MUST start at H and go all the way through to Li:  $1s^22s^1$ . This takes into account all of the electrons necessary to make up Li.

The 4th element is Be. It is in group II and period 2. Its electronic configuration is  $1s^22s^2$ . For this period, this closes out the s block elements.

Boron (B) is the 5th element on the periodic table. It is in group III and period 2. Its electronic configuration is  $1s^22s^22p^1$ .

The 6th element is carbon (C). It has one more electron than B:  $1s^22s^22p^2$ .

Nitrogen's (N) electronic configuration is  $1s^22s^22p^3$ . At this point, the 2p subshell is half full.

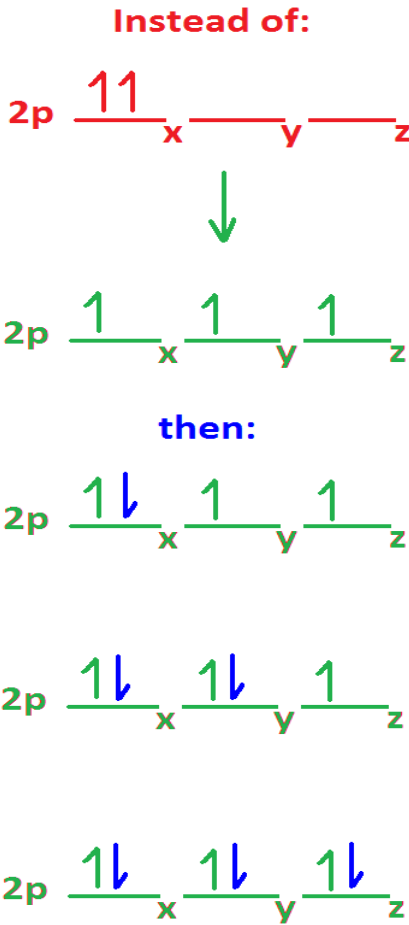
**Hund's Rule**

There is a specific rule dealing with the order in which sub-subshells fill. This rule is Hund's Rule. Hund's Rule says that the electrons enter each sub-orbital of a given type singly and with identical spins before any pairing of electrons of opposite spin occurs within those orbitals:

This graphic shows that each sub-subshell (p, d or f -- p in the graphic) must half fill, first, with electrons of one spin before the sub-subshells can completely fill with electrons having opposite spin.

In other words, for a p subshell, all three sub-subshells must half fill with one electron each before the 4th electron can fill the first sub-subshell, the 5th electron can fill the second sub-subshell and the 6th electron can then fill the last sub-subshell.

Oxygen (O) has one more electron:  $1s^22s^22p^4$ ; while fluorine (F) has one more than O:  $1s^22s^22p^5$ . Neon (Ne) fills up the p subshell and closes out the second period:  $1s^22s^22p^6$ .



The 11th element is sodium (Na). It starts out the third period and has one more electron than Ne:  $1s^2 2s^2 2p^6 3s^1$ . There is another way in which to write this electronic configuration that is a sort of shorthand: rather than going through all of the electronic configuration, use the previous noble gas as a core and add to it, e.g., for Na, then, the electronic configuration can be written as  $[\text{Ne}]3s^1$ .

For Mg, the electronic configuration is  $[\text{Ne}]3s^2$ . The remainder of the third period elements' electronic configuration are tabulated below in both formats:

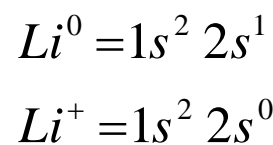
Element	Traditional notation	Shorthand notation	Element	Traditional notation	Shorthand notation
Al	$1s^2 2s^2 2p^6 3s^2 3p^1$	$[\text{Ne}] 3s^2 3p^1$	S	$1s^2 2s^2 2p^6 3s^2 3p^4$	$[\text{Ne}] 3s^2 3p^4$
Si	$1s^2 2s^2 2p^6 3s^2 3p^2$	$[\text{Ne}] 3s^2 3p^2$	Cl	$1s^2 2s^2 2p^6 3s^2 3p^5$	$[\text{Ne}] 3s^2 3p^5$
P	$1s^2 2s^2 2p^6 3s^2 3p^3$	$[\text{Ne}] 3s^2 3p^3$	Ar	$1s^2 2s^2 2p^6 3s^2 3p^6$	$[\text{Ne}] 3s^2 3p^6$

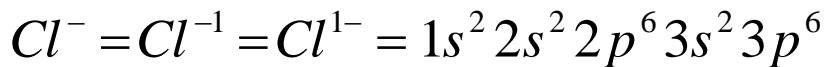
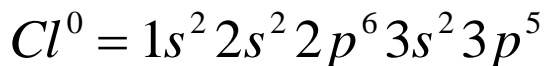
Potassium (K) starts out period 4 in Group I. In this case, the shorthand method for its electronic configuration is  $[\text{Ar}]4s^1$ . The last element for which we will consider electronic configuration is Ca. Its configuration is  $[\text{Ar}]4s^2$ .

### Periodic Law

Short periods		
Period	Contents	Orbitals
First period	H ® He (2 elements)	1s
Second period	Li ® Ne (8 elements)	2s2p
Third period	Na ® Ar (8 elements)	3s3p
Long periods		
Fourth period	K ® Kr (18 elements)	4s3d4p
Fifth period	Rb ® Xe (18 elements)	5s4d5p
Sixth period	Cs ® Rn and Lanthanides (32 elements)	6s5d4f6p
Seventh period	Fr ® Ac and Actinides (Incomplete; holds 16 elements)	7s6d5f7p

Let's first start examining atoms with some of their ions. Ions are electrically charged particles that result from the gain or loss of electrons. Figure, right, illustrates the difference between elemental lithium and the monovalent





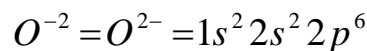
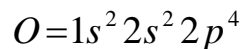
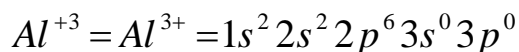
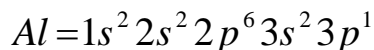
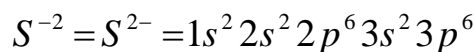
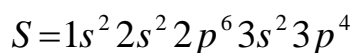
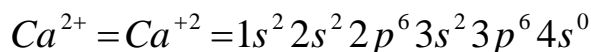
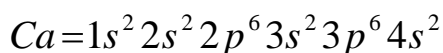
lithium cation utilizing electron configuration. Note that in order for Li to ionize that it loses the single 2s electron. I left the empty 2s shell in that figure to

illustrate where the electron used to be. It is not necessary to write out the empty orbitals.

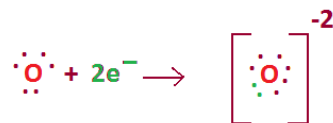
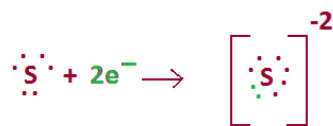
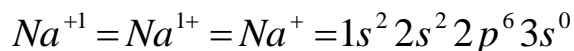
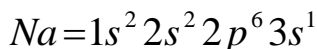
Likewise, the figure, left, illustrates what happens when elemental chlorine is ionized to the monovalent chloride anion. In this case, an extra electron is gained in the 3p shell, filling it. What's important, here, is to remember which electron shells contain which orbitals/sub-shells.

The reason this is important goes right back to our earlier discussion that explained that atoms will gain or lose sufficient electrons to achieve an outer electron arrangement identical to the preceding or following noble gas. **This means that the K shell wants to be full at 2 electrons; the L shell at 8 electrons; the M shell at 8 electrons -- these correspond to He, Ne and Ar, respectively.**

Examples of elemental states going to their respective ionic states using electronic configuration:

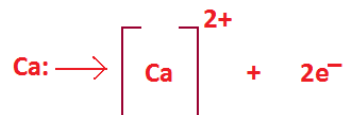


There is another method in which to approach this concept, as well. It is called Lewis Structures.



### Lewis Structures

Lewis Structures, for all intents and purposes, utilize different symbols, e.g., dots, squares, "x's", triangles, and nauseum, to represent the electrons in the outer valence shell of the atoms. Remember that it takes 2 electrons to make a single bond. Whenever possible, you need to place 8 electrons around your atoms (= octet rule). At times you



will only get 2 electrons around some atoms, e.g. H, Li, Na, K and that fulfills the octet rule, too. Fortunately for us, the first 20 of the representative elements are the easiest and the ones we'll focus on. Remember that these groups are in numerical order from I to VIII. The number of the group tells you exactly how many electrons are in the outer valence shell of the atom with which you are working. Thus, we may draw the ionization of each of the elements described above in the figure at right. Note that cations lose their outer shell electrons and anions gain electrons in their outer shells.

All of these are examples of simple ions: atoms that have acquired a net positive or negative charge by losing or gaining electrons, respectively. Ions and atoms with identical electronic configuration are called isoelectronic:

The last Group, Group VIII, has 8 electrons in its outer shell -- with the exception of He which only has 2 (in the 1s subshell).

### Characteristics of Elements in The Groups of The Periodic Table

#### Group I

Element	Configuration	Atomic radius	Ionic radius	Density	Melting point	Boiling point
Li	[He]2s <sup>1</sup>	Smallest	Smallest	Least	Highest	Highest
Na	[Ne]3s <sup>1</sup>		Ionic radii are smaller than their atomic radii.			
K	[Ar]4s <sup>1</sup>					
Rb	[Kr]5s <sup>1</sup>					
Cs	[Xe]6s <sup>1</sup>					
Fr	[Rn]7s <sup>1</sup>	Largest	Largest	Most	Lowest	Lowest

#### Hydrogen

Note that I did not include hydrogen in the table above. It's a funny element that has numerous contradictory characteristics that make it in a class by itself. In fact, some of the older periodic tables put H in the middle above the periodic table because of that.

Group I elements are also known as the alkali metals. They are known as that as they produce alkali when reacted with water. The table below summarizes the types of reactions that alkali metals undergo with examples of each reaction. Note: "M" represents any alkali metal. "X" represents any halogen, e.g., F, Cl, Br, I.

Reaction type	Example
With water (rapid! hot!) -- Forms hydroxides.	$2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2$
With halogens (rapid! hot!) -- Forms halides.	$2M(s) + X_2 \rightarrow 2MX(s)$
With hydrogen -- Forms hydrides.	$2M(s) + H_2(g) \rightarrow 2MH(s)$
With oxygen -- Forms oxides.	$4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$ $2Na(s) + O_2(g) \rightarrow Na_2O(s)$ $4K(s) + O_2(g) \rightarrow 2K_2O(s)$
With nitrogen -- Forms nitrides.	$6Li(s) + N_2(g) + \text{heat} \rightarrow 2Li_3N(s)$

## Group II

Elements in Group II are called the alkaline earth metals as they are found in earthy substances. They also react with water to form alkalis. The table, below, summarizes the properties of these elements:

Element	Configuration	Atomic radius	Ionic radius	Density	Melting point	Boiling point
Be	[He]2s <sup>2</sup>	Smallest	Smallest	Higher	Highest	Highest
Mg	[Ne]3s <sup>2</sup>					
Ca	[Ar]4s <sup>2</sup>			Smallest	Way lowest	Way lowest
Sr	[Kr]5s <sup>2</sup>			-	-	-
Ba	[Xe]6s <sup>2</sup>			-	-	-
Ra	[Rn]7s <sup>2</sup>	Largest	Largest	Highest	Lowest	Lowest

For some strange reason, Ca has really weird properties. Otherwise it fits into Group II. The table, below, summarizes the reaction types and example reactions for this group:

Reaction type	Example
With water (vigorous) Forms hydroxides.	$M(s) + 2H_2O(l) \rightarrow 2M(OH)_2(aq) + H_2$ When "M" = Mg, add heat for the reaction to go
With halogens (rapid! hot!) Forms halides.	$M(s) + X_2 \rightarrow MX_2(s)$
With hydrogen. Forms hydrides.	$M(s) + H_2(g) + \text{heat} \rightarrow MH_2(s)$

With oxygen. Forms oxides.	$2M(s) + O_2(g) \rightarrow 2MO(s)$ When "M" = Be or Mg, add heat for the reaction to go $Ba(s) + O_2(g) + \text{heat} \rightarrow BaO_2(s)$
With nitrogen. Forms nitrides.	$3M(s) + N_2(g) + \text{heat} \rightarrow M_3N_2(s)$

Alkaline earth compounds have numerous uses:

Compound	Formula	Common name and use
Magnesium hydroxide	Mg(OH) <sub>2</sub>	Milk of magnesia; laxative
Magnesium oxide	MgO	Magnesia -- fire bricks
Magnesium sulfate	MgSO <sub>4</sub>	Magnesium sulfate heptahydrate is Epsoms salts and used for laxative; magnesium sulfate is used to treat pre-term labor
Calcium carbonate	CaCO <sub>3</sub>	Primary ingredient in limestone, marble, coral, shells, Tums
Calcium hydroxide	Ca(OH) <sub>2</sub>	Slaked lime -- aqueous is white wash
Calcium chloride	CaCl <sub>2</sub>	Desiccant (drying agent); melts ice, too
Calcium sulfate	CaSO <sub>4</sub>	The dihydrate salt is called gypsum (sheet rock); hemihydrate is plaster of Paris
Barium sulfate	BaSO <sub>4</sub>	Opaque to x-rays; x-rays of GI tract
Calcium oxide	CaO	Lime -- used in mortar, cement, plaster

## Metals of Significance

The table, below, summarizes the ions in their ANALYTICAL groups – NOT groups from the periodic table:

Group 1	Group 2	Group 3	Group 4	Group 5	"Extras"
Metals: Ag, Pb, Hg	Metals: Bi, Cu, As, Sb, Sn, Hg	Metals: Ni, Co, Mn, Fe, Al, Cr, Zn	Metals: Ca, Sr, Ba	Metals: Mg	Au, Ti, Zr
Ionic forms: Ag <sup>+</sup> , Pb <sup>2+</sup> , Hg <sub>2</sub> <sup>2+</sup>	Ionic forms: Bi <sup>3+</sup> , Cu <sup>2+</sup> , As <sup>3+</sup> , Sb <sup>3+</sup> , Sn <sup>2+</sup> , Hg <sup>2+</sup>	Ionic forms: Ni <sup>2+</sup> , Co <sup>2+</sup> , Mn <sup>2+</sup> , Fe <sup>2+</sup> , Al <sup>3+</sup> , Cr <sup>3+</sup> , Zn <sup>2+</sup>	Ionic forms: Ca <sup>2+</sup> , Sr <sup>2+</sup> , Ba <sup>2+</sup>	Ionic forms: Mg <sup>2+</sup>	

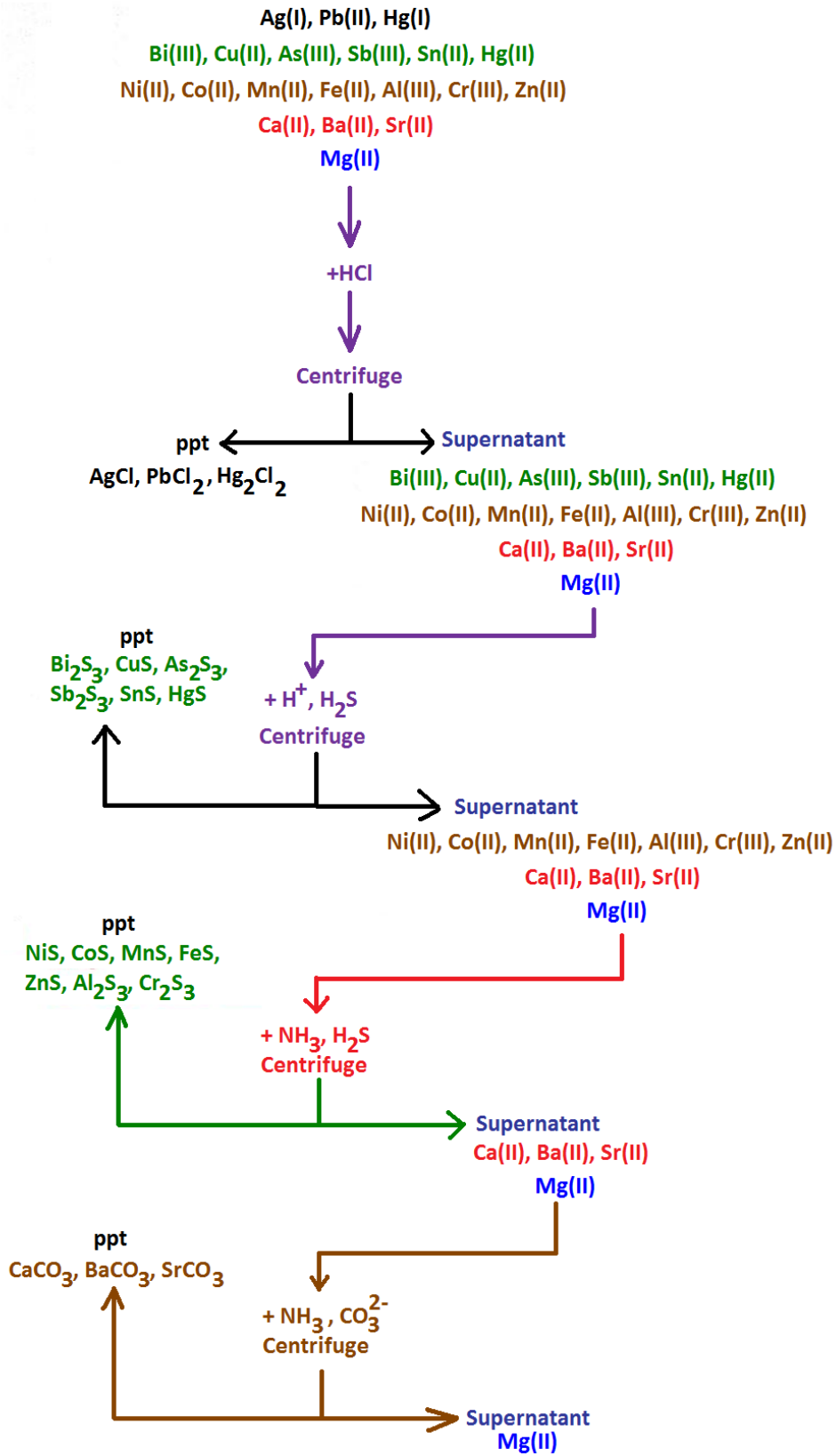
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This grouping is not inclusive and is intended to serve only as an introduction. You'll get the remainder in CHEM 122.

These groups were organized after observing that these ions were separable in specific groups by specific reagents (chemicals in solution). How best to approach this? Well, keep in mind that you can't just go out and buy a bottle of a cation or an anion. You have to buy a SALT -- remember those things? When you put a salt in water, it dissociates into a cation and an anion. In this case, we're interested in the cations. If all of these cations were in solution, could we separate them into the groups, above?

YES! Image on next page.





## Group 1

### Ag -- silver

Horn silver is silver chloride. Tarnished silver is silver sulfide ( $\text{Ag}_2\text{S}$ ). Typically, silver is obtained as a by-product of other metals. It is a white lustrous metal. Its polished surface reflects light. It is the best conductor of heat and electricity other than gold. It's not used as such as it's too expensive for that. It is malleable (can be pounded out into a flat sheet) and ductile (can be drawn into a wire). It is used for coins, silverware and ornaments. Sterling silver is about 7.5% Cu. Jewelry silver is 20% copper. It has been used in dentistry in amalgam filling; it is used in photography and mirrors. It is also used to stop bleeding after calves are de-horned and has been used in solution to scarify tissue in the nasal cavity for those susceptible to epistaxis.



### Pb -- lead

Lead pipes led to Rome's downfall -- pun intended. Galena is fool's silver and is lead sulfide. It is the heaviest of the metals except for mercury and gold. Tetraethyl lead [ $\text{Pb}(\text{C}_2\text{H}_5)_4$ ] is a catalytic converter poison and was used in leaded gasoline to prevent knocking. Pb is used in car batteries. Lead is used in stained glass windows. It is no longer used in pigments for paints, nor is it used in solders, much. Silver has replaced it in many solders.



### Hg – mercury

Both 1+ and 2+ forms are considered, here. Aristotle called this liquid silver or quicksilver. It's used to extract gold due the high solubility of gold in mercury. It's a silvery white metal and is the only metal in the liquid state at room temperature. It has enjoyed wide-spread use in thermometers. That is changing as we learn more and more how dangerous even a little mercury can be. It is not used as much in dentistry to make amalgam fillings (Note: whenever you read "amalgam", think mercury).



Hg is used in mercury vapor lamps and fluorescent lamps.  $\text{Hg}_2\text{Cl}_2$  is calomel -- used in calamine lotion.  $\text{HgCl}_2$  is corrosive sublimate and is very toxic. Mercury salts have been used as antiseptics -- not so much any more.

Mercury is a cumulative poison. A fatal dose is 0.2-0.4 g (200-400 mg). The treatment, at least initially, is to give egg whites and milk (both contain proteins with LOTS of SH groups that LOVE Hg and remove it from binding with proteins in the GI tract).

## Tooth Filling Materials:

**Amalgams:** A mixture of silver, tin, zinc, copper, and mercury. Mercury is nearly 50% of the mixture. **Lasts** at least seven years, usually longer; for back teeth fillings.

**Composite Resin:** A mixture of plastic and fine glass particles. **Lasts** at least five years; for visible portions of teeth.

**Cast Gold:** Gold alloy (gold mixed with other metals). **Lasts** at least seven years, usually longer; withstands the forces of chewing.

**Gold Foil:** Gold. **Lasts** approximately 10 to 15 years; small fillings with lesser chewing forces; crown repair.

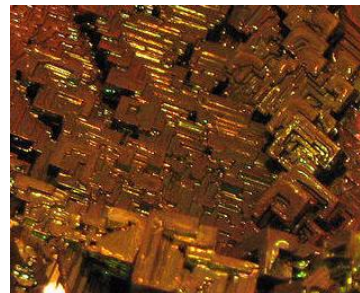
**Ceramics:** Porcelain, most commonly. **Lasts** five to seven years; matches tooth color.

**Glass Ionomer:** Acrylic and a component of glass called fluoroaluminosilicate. **Lasts** five years or more; for filling below the gum line in root caries. [http://www.colgate.com/app/CP/US/EN/OC/Information/Articles/Oral-and-Dental-Health-Basics/Checkups-and-Dental-Procedures/Fillings/article/Types-of-Fillings.cvsp]

## Group 2

### Bi -- bismuth

Alloys of Bi, Sn and Pb have low melting points. Uses of these alloys include electrical fuses; safety plugs for boilers and automatic sprinkler systems. Bismuth subcarbonate is  $(\text{BiO})_2\text{CO}_3$  and bismuth subnitrate is  $(\text{BiO})\text{NO}_3$ . Many bismuth compounds (Pepto BISMol) are used to treat gastritis and ulcers; others are used to treat skin diseases such as eczema.



### As -- arsenic

Gray As is mono-atomic and sublimates at  $615^\circ\text{C}$ , where it becomes tetra-atomic and takes on a tetrahedral conformation. When As vapor is cooled FAST, it becomes unstable yellow As. The yellow As is soluble in carbon disulfide and has an odor of garlic. Yellow As is very poisonous. Bronze and lead shot contains As as a hardening agent. As is used mostly as a weed killer, cattle/sheep dip and in insecticides.

### Sb -- antimony

The Egyptians used Sb for eye makeup as early as 3000 B.C. It's been used in medicines. It is a lustrous white metal and is brittle and readily pulverized. Lead plus 10-20% antimony makes "hardened lead" which is used in shrapnel, bullets and bearing. Babbit metal contains Sb (Sb, Sn and Cu) and is used as an anti-friction alloy in the production of ball bearings.



Cu -- copper

Copper is reddish-yellow metal. It is ductile and malleable. It is the best electrical conductor of the cheap metals. It needs, though, to be very pure: impurities readily decrease its conductivity. Biologically, Cu is used with vitamin C, lysine for connective tissue biosynthesis and in the synthesis of hemoglobin. The following alloys of copper are commercially useful:

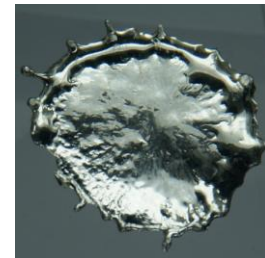


Brass	Bronze	Aluminum bronze	German silver
60-82% Cu; 18-40% Zn	70-95% Cu; 1-25% Zn; 1-18% Sn	90-98% Cu; 2-10% Al	50-60% Cu; 20% Zn; 20-25% Ni

Note that in German silver there is no silver.

Sn -- tin

Tin cans are really tin-plated steel cans. There are three forms of tin:



Gray Tin

Gray tin is formed from malleable tin after prolonged exposure to temperature extremes below 13.2° C. It becomes powdery and spread like a rash -- hence its name: tin disease or tin pest. This is a problem for organ pipes in northern Europe. The pipes are made of tin for the best sound and are regularly replaced as a result of the cold temperatures.

Malleable Tin

This tin is a silver-white with bluish tinge. It is also known as white tin. When this tin is bent it emits what is called the "tin cry". When this tin is heated it forms brittle tin. When malleable tin is exposed to temperatures below 13.2° C, it forms gray tin.

Brittle Tin

Formed from heated malleable tin.

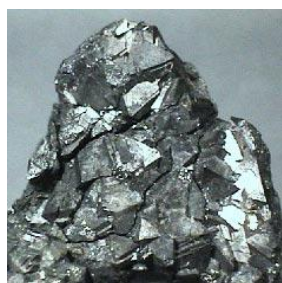
Tin is also found in different alloys:			
Bronze	Type metal	Solder	Chrome plate substitute
Cu, Zn, Sn	Used in the old presses to print newspapers; Sn, Pb, Sb, Cu	Sn and Pb (the latter is being slowly replaced with Ag)	One part nickel and 2 parts tin approximate a cheap replacement for chrome plating.

Group 3



Co -- cobalt

Similar to iron. Like Fe and Ni, Co is magnetic. Permanent magnets are made from Alnico (Al, Ni, Co and Fe), Hiperco (Co, Fe, Cr) and Vicalloy (Co, Fe, V). Cobalt is alloyed with iron and other metals for high speed cutting tools and surgical instruments. Black CoO gives glass a blue color. This glass is called cobalt glass.  $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$  in alcohol is used as "invisible ink". To make it show up, heat the paper. As the salt dehydrates, the writing becomes blue. It fades upon rehydration from the environment.



Ni -- nickel

Nickel is a silvery-white metal. It is hard, malleable and ductile. It is used in plating iron, steel and copper. Permalloy consists of Ni and Fe and is used in sound reproduction. Nichrome and Chromel consist of Ni, Cr, and Fe. Alnico consists of Al, Ni, Co and Fe.

Platinite and Invar are used to "seal in" wires in light bulbs as they expand with the heat and maintain the vacuum in the light bulb; conversely, it shrinks with cooling temperature and does the same thing (maintains vacuum) as the light bulb gets cooler. Ni powder is a hydrogenation catalyst of, among other things, unsaturated fatty acids.

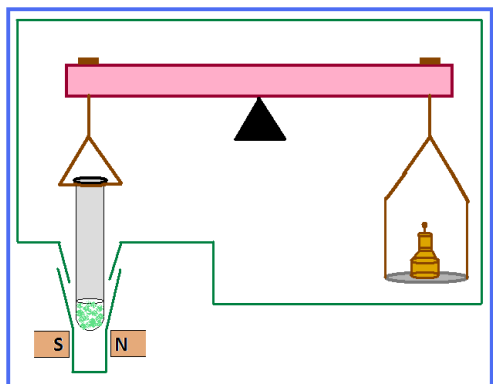


Fe

Iron (III) tannate is nearly colorless. It is made by reacting tannic acid with ferrous sulfate. It is used as ink. When the salt is oxidized to ferric tannate, it turns black. Before that happens, though a dye makes the writing visible. You may have noticed how you thought you wrote something in blue, yet when you came back to it a day later it was black. You're right: the blue dye made the writing visible and then the black came through as the real ink was oxidized by the air. The ink can be cleaned off cloth by soaking it in ammonium oxalate for several hours. This treatment reduces the water insoluble ferric tannate to water-soluble ferrous tannate.

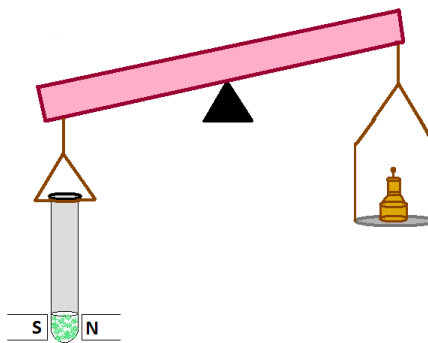
Steels will be discussed later, below.

Iron, cobalt, nickel and gadolinium are all magnetic. There are three kinds of magnetism:		
Paramagnetism	Diamagnetism	Ferromagnetism
In the case of paramagnetism, there are an odd number of unpaired electrons in the outer shell making the metal attract to a magnet.	In the case of diamagnetism, there is an even number of paired electrons in the outer shell, making the metal repel from a magnet.	Ferromagnetism is an extreme case of paramagnetism and leads to permanently magnetized magnets.



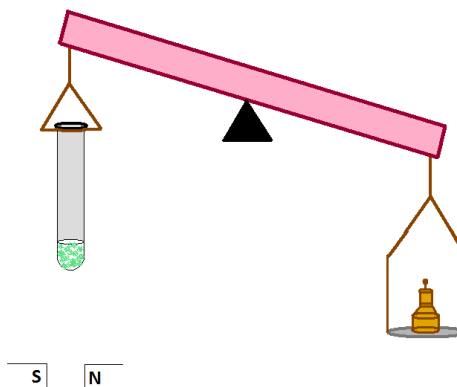
A Guoy balance (Magnetic Susceptibility Balance is what they are using, now) is a balance on a knife-edge. It has a pan on one arm and a test tube holder on the other that has an electromagnet associated with it.

One places the sample in the test tube and adds masses to the pan to balance it out. The masses are recorded and the electromagnet turned on.



If the sample is paramagnetic, the sample is attracted to the electromagnet. Masses are added to re-

balance it. The difference in the masses is used to calculate exactly how many electrons are in the outer shell and are unpaired.

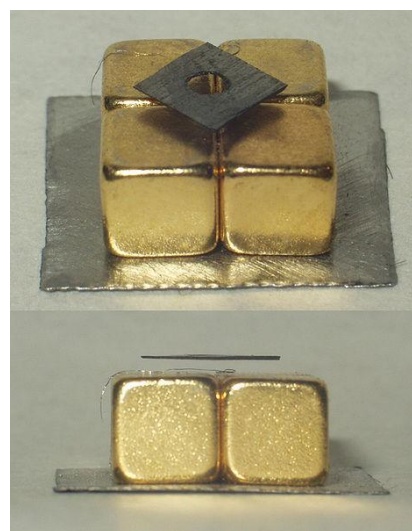


If the sample is diamagnetic, the sample is repelled from the electromagnet. Diamagnetism is demonstrated in the lower right image: levitation of pyrolytic graphite.

Guoy balance concept, top left. Magnetic susceptibility balance, bottom left. Paramagnetic, top right. Diamagnetic, bottom right.

Iron salts were also used in blueprints. When ferric ammonium citrate is reacted with potassium ferricyanide, ferriferricyanide (bronze green) is formed. Blueprints are drawn on tracing cloth in black ink. When the blue print is made, the cloth is laid over the green blueprint paper. Hit it with light and the ferric ion is reduced to the ferrous ion, which leaves a blue precipitate wherever the light hits. Since the light doesn't go through the black lines, you're left with blue paper and white lines.

Iron metals need to be protected to reduce corrosion. They may be covered with paints, grease, asphalt, Zn (galvanized), Cu, Ni, Cr, Sn, Cd, Pb, ceramic enamel like you find on tubs or with adherent oxide (where the iron is treated superheated steam to form magnetic oxide:  $\text{Fe}_3\text{O}_4$ ).



Iron pyrites are called fool's gold. It is an unsatisfactory iron source, although it is used as a sulfur source for sulfuric acid manufacture. Hematite ( $Fe_2O_3$ ) and Magnetite ( $Fe_3O_4$ ) are used as iron ores. The metal formed from a blast surface is called pig iron.

Pig iron is brittle and is generally used to make cast iron. If it is re-melted and re-cooled, it forms cast iron. If pig iron is rapidly re-cooled, it forms  $Fe_3C$ , a white cast iron that is brittle, but hard and wears well. If the pig iron is re-cooled slowly,  $Fe_3C$  (gray cast iron; soft and tough) and graphite (C) are obtained.

Is it possible to obtain both hard and wears well AND soft and tough for useful cast iron? YES! Cool the surface rapidly and the body of the casting slowly.

Steel -- coming up shortly -- consists of different alloys of iron. It's pig iron minus impurities plus some of the following: Mn, Cr, Ni, W, Mo, V and 0.04-2.5% C.

There are three primary steels: Stainless steel (low carbon content and contains 12% or more Cr), Alloy steel (gives special properties for special uses; contains 1-5% Si for magnets) and Carbon steel (primarily iron and carbon). There are three Carbon steels:		
Mild	Medium	High
This contains not more than 0.2% carbon and is ductile. It is used to make hot irons, wire and pipe. It usually contains 2-15% Si.	This steel contains 0.2-0.6% carbon. It's used to construct rails, structural steel and boilerplate. It usually contains about 10-18% Mn.	High carbon steel contains 0.6-1.5% carbon. It is hard to brittle. It is used to make surgical instruments, razor blades, springs and cutlery. It contains about 14-18% chromium and 7-9% nickel
These two steels can be forged and welded.		



Mn -- manganese

Mn is used in alloys of iron for steel. Alloys high in manganese are called ferromanganese. Alloys low in manganese are called Spiegeleisen (mirror iron). Mn is a gray-white metal with a slight reddish tinge -- it is brittle and looks like gray cast iron. It takes about 12.5 pounds of elemental manganese per ton of steel to remove oxygen, nitrogen and sulfur.

Manganin (84% Cu, 12% Mn, 4% Ni) is used for electrical measuring instruments since there is no change in the resistance following temperature changes.  $MnO_2$  colors glass black.  $KMnO_4$  is a disinfectant and has been used as a deodorant and germicide.



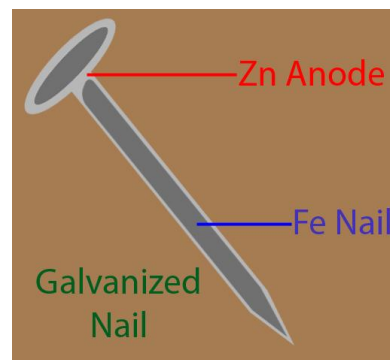
Cr -- chromium

Cr is a very hard, silvery-white, crystalline metal. Chrome alum is used to tan leather, print calico and waterproof fabrics. Cr is in stainless steel. Ni and Cr are present in Nichrome and Chromel -- 2 non-ferrous alloys. Cr is a protective and decorative coating. Cr is also used in paints as a pigment.



Zn -- zinc

Zn is a silvery metal that tarnishes to a blue-gray. It is used in dry cells, brass, bronze and used to coat iron. When it is used to coat iron or steel, it is called galvanization (making a battery). To make the nail not rust and, hence, make your fence not fall apart, the nail (cathode) is coated with zinc (anode) to protect the nail from rusting (oxidation). ZnO is used for sunblock. Zn is hard and brittle at usual temperatures, but at 100-150 C is ductile and malleable -- this is when Zn ribbon is made.



Al -- aluminum

Al has a silvery appearance, but dulls white due to oxidation. Weight for weight, it is twice as good a conductor of electricity as copper. It is resistant to corrosion. It is used in cooking utensils. Duralumin (Al, Cu, Mn, Mg) is very light and very strong and is used in aircraft.

Aluminum bronzes (Cu with Si, Mn, Fe, Ni or Zn and Al) are around, too. Al has been used for crankcases and connecting rods in automobiles. A last use of Al is in Emery cloth -- the black cloth used to smooth rough ends on copper tubing or pipe. This is a combination of  $\text{Al}_2\text{O}_3$  and  $\text{Fe}_3\text{O}_4$  and is an excellent abrasive.



Thermite is a very interesting stuff and has several uses: rail welding and bombs in WW II. The rail welding is usually called the thermite reaction. The way it works is to have some guys drive up to a remote site on the railroad where a rail has broken, they hop out, place a container over the break and add the thermite mixture. They put a magnesium fuse in some ignition powder, light it and let 'er rip.

The reaction is  $2\text{Al}(s) + \text{Fe}_2\text{O}_3(s) \rightarrow 2\text{Fe}(s) + \text{Al}_2\text{O}_3(s)$ .

In other words, the rail is welded in place at temperatures well above 3000°C. When used as bombs in WW II, they burned with high temperatures. Water doesn't quench this kind of fire, making them nasty armaments.



## Group 4

All members are silvery-white, crystalline metals that are malleable and ductile.



Ba -- barium

Was used in vacuum tubes -- if you can remember them. It is also used in spark plugs. Is softer than Pb and has a yellow/green flame.

$\text{BaSO}_4$  is used in studies of the GI tract (upper GI, lower GI [Ba enema], Ba swallow, small bowel series. Although  $\text{BaSO}_4$  is extremely poisonous it is so insoluble that it is of a non-poisonous nature for medical use.

## Sr -- strontium

Has no commercial uses.  $^{90}\text{Sr}$  was a fallout product of atomic bombs. Sr is in Group II -- same as Ca. Causes a lot of osteogenic sarcoma. Is about as hard as Pb and has a carmine flame.



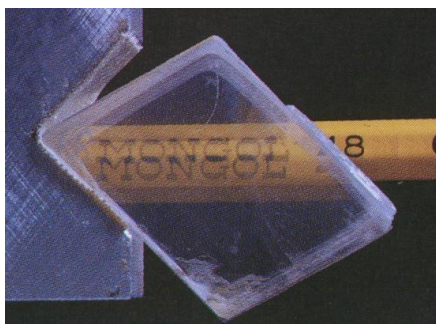
Ca – calcium

Used in dehydrating agents; alloyed with Si; used in bearings and in storage battery grids. Is harder than Pb and has a dark red flame.

Compounds of significance include  $\text{CaO}$  (unslaked lime or quicklime),  $\text{Ca(OH)}_2$  (slaked lime; hydrated quicklime), saturated  $\text{Ca(OH)}_2$  (limewater) and a suspension of  $\text{Ca(OH)}_2$  (milk of lime).

In caves, calcium reacts with bicarbonate to form calcium carbonate and water and  $\text{CO}_2$ . Stalactites hang from the ceiling and stalagmites grow up from the floors in limestone caves.

The reaction that occurs between mortar and  $\text{CO}_2$  that is in the air takes 3-4 parts sand with 1 part calcium hydroxide, 1 part water and  $\text{CO}_2$  from the air to produce  $\text{CaCO}_3$  and water. This reaction takes hundreds of years to form (think castles of Germany that are 400-600 years old and still standing), making mortar very time efficient.



Calcite

Calcite ( $\text{CaCO}_3$ ) precipitates at temperatures less than  $30^\circ\text{C}$ . It has the property of birefringence (double refraction): when a beam of light enters calcite, it is split into two beams and two images are observed.

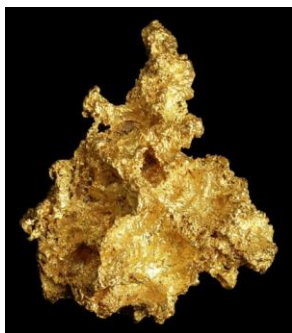


## Group 5

### Mg -- magnesium

Mg is obtained from underground brines and seawater as the chloride salt. It is a silvery-white metal. It is malleable and ductile at high temperatures. It was used in flash powders. It is still used in military flares and incendiary bombs -- it burns hot and bright. Since Al is hard at machining temperatures, Mg is alloyed with it to make Magnalium (1-15% Mg, 0-1.75% Cu and Al) or Dowmetal (8.5% Al, 0.15% Mn, 2% Cu, 1% Cd, 0.5% Zn and 87.85% Mg) which is easier to machine.

### Au -- gold



Gold is a soft, yellow metal. It is the most malleable and ductile of all metals, e.g., gold can be hammered so thin that 300,000 sheets of gold foil would make a pile one foot thick; one gram of gold wire can be pulled more than 1.5 miles long. Gold is ALWAYS alloyed because it's too soft to work with by itself.

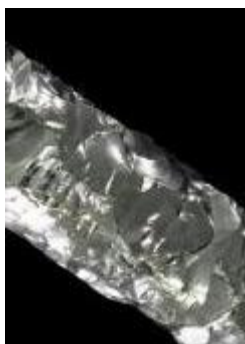
Red/yellow gold contains Cu; White gold contains Pd, Ni or Zn. Gold is very inactive, BUT it dissolves in aqua regia (solution of HCl and HNO<sub>3</sub>). Forms of Au are used to treat rheumatoid arthritis. Gold is also used in dentistry.

### Ti – titanium

TiO<sub>2</sub> is rutile. It is a white pigment for paint. It's used in the production of white rubber and white leather. Elemental Ti is very strong, light and resistant to corrosion. It's used in jet motors and high-speed aircraft. When TiCl<sub>4</sub> reacts with moist air, a dense white smoke is produced, which is used for skywriting and smoke screens.



### Zr -- zirconium



Zircon is ZrSiO<sub>4</sub> and is a semi-precious stone. ZrO<sub>2</sub> is zirconia. It has excellent refractory qualities. It is used in the production of enamels and opaque glass.

## Halogens

Jumping across to the halogens (Group VII), their characteristics are summarized, below. Halogens react to form halides -- salts.

Element	Valence shell configuration	Atomic radius	Ionic radius	Density	Melting point	Boiling point	Reactivity
F	$2s^2 2p^5$	Smallest	Smallest	Lowest	Lowest	Lowest	Highest
Cl	$3s^2 3p^5$		Ionic radii are bigger than atomic radii				
Br	$4s^2 4p^5$						



per cent composition does not vary - the PRESSURE does, e.g., at sea level, atmospheric pressure is 760 mm Hg; at 15000 feet, 400 mm Hg; at 10 miles, 40 mm Hg and at 30 miles, 0.1 mm Hg.

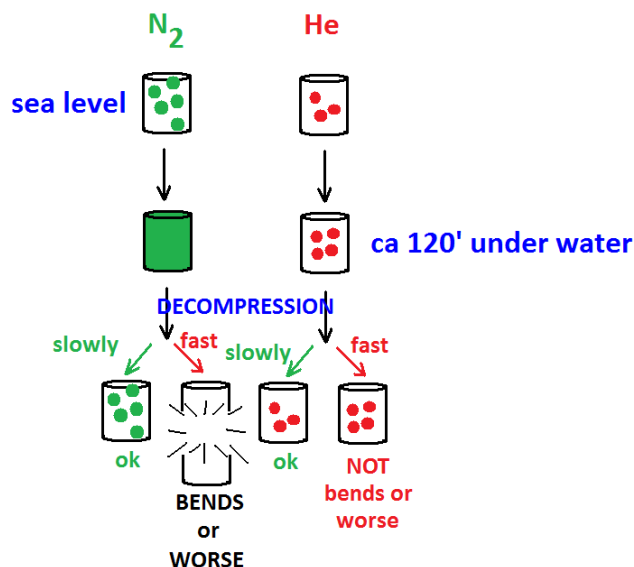
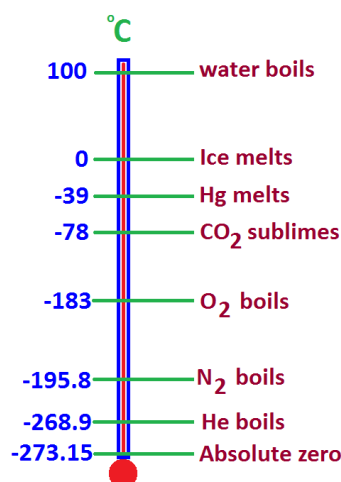
N <sub>2</sub>	78%	Ne ←	0.0012%
O <sub>2</sub>	20.99%	He ←	0.0005%
Ar ←	0.94%	Kr ←	0.0001%
CO <sub>2</sub>	0.035-0.04%	O <sub>3</sub>	0.00006%
H <sub>2</sub>	0.01%	Xe ←	0.000009%

← All noble gases

### General Comments Regarding Noble Gases

Some of the more common temperatures in chemistry are at right. The one to emphasize is that of the boiling point of He: -268.9° C.

He is an unusual noble gas. One form of <sup>4</sup>He<sub>2</sub> has no electrical resistance. It has zero viscosity, i.e., it flows up and over the edges of the container until the levels of He inside and outside the container are level. That it has no electrical resistance makes it useful as a super conductor. He also has no triple point, unlike water. He is used in arc welding, to fill weather balloons, to fill blimps/dirigibles (it's non-flammable).



He is also used in deep-sea diving. In general, He is much less soluble in water than are either oxygen or nitrogen. At the high pressures of deep sea diving, LOTS of nitrogen (N<sub>2</sub>) dissolves in water, blood and cerebrospinal fluid. If decompression is attained appropriately, there is no problem.

If, however, decompression is too rapid, the person develops the bends. In this part of the country, this is usually fatal as there is no decompression chamber at Tahoe, any more and the closest one is at Davis, I believe.

OTOH, if He is substituted for the N<sub>2</sub>, even rapid decompression is not as deadly as it is for nitrogen-containing gas mixtures. This means that for long term deep-sea diving, the likelihood of developing the bends is greatly reduced.

Ar is cheaper than He. We use about  $1.5 \times 10^9$  pounds per year. It is used in incandescent bulbs to prevent bulb "burn out". Ne, Ar and Xe are often coupled with Kr in "neon lights". The color of the light is dependent upon the composition of the gases.

Pure Ne gives an orange-red light; pure Xe gives a blue light.

Xenon is water-soluble and has found some use in compounds as an anesthetic. Xe is capable of attaining one of three hybridizations:  $sp$ ,  $dsp^2$  or  $d^2sp^3$ .

What are hybridizations, where do they come from and how do we explain them? For that we have to turn to Quantum Theory and Wave Mechanics.

### Wave Mechanics: An Introduction

Remember that light has dualistic properties. Louis DeBroglie looked at matter and wondered if matter had wave-form properties, i.e., does matter vibrate at its own unique frequency? DeBroglie's theory revolutionized the mathematical description of atoms and matter. DeBroglie's theory proposes that a particle moving has some sort of wave nature that allows for the calculation of its wavelength-equivalence:

$$\lambda = (h/mv)$$

$\lambda$  = wavelength-equivalence

$h$  = Planck's constant

$m$  = mass of the particle

$v$  = velocity (speed) of the particle

Remember that  $p = mv$

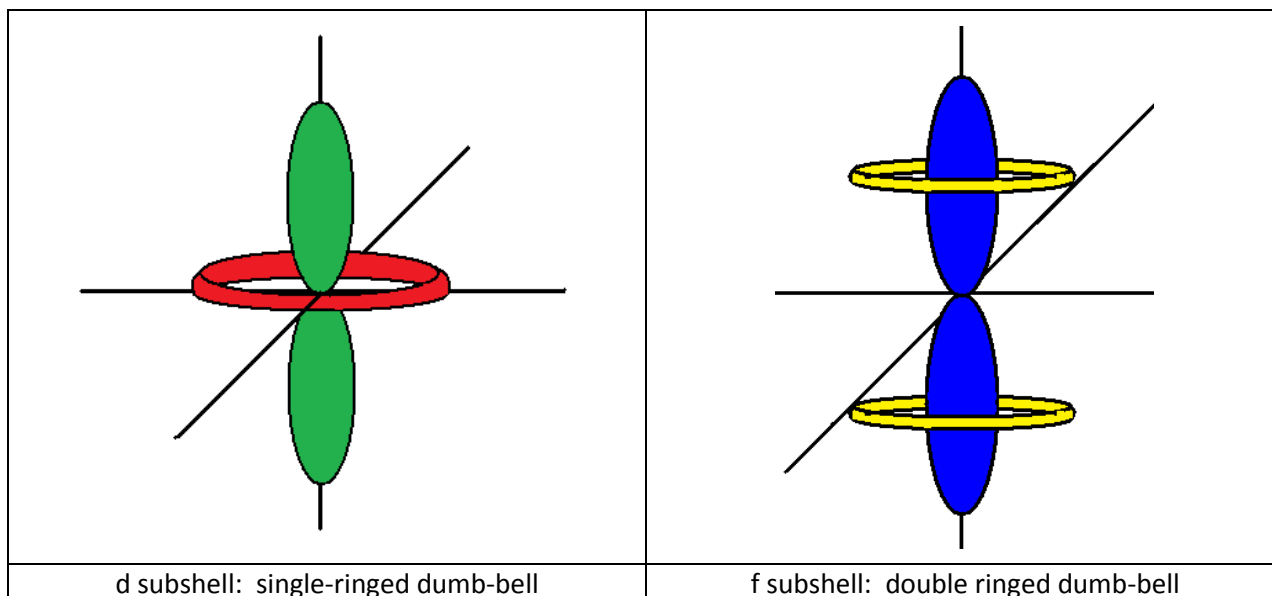
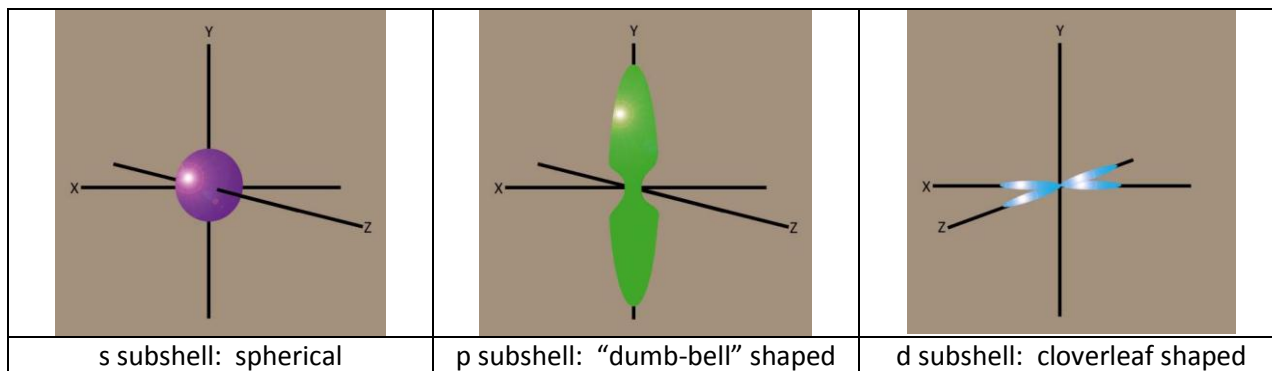
This work resulted in the manufacture and technology behind the electron microscope.

Two other scientists took DeBroglie's work and applied to the wave-like properties of sub-atomic particles. Werner Heisenberg developed his uncertainty principle at the same time. His equation allows for the determination of the degree of uncertainty when measuring a particle's location and momentum simultaneously. By virtue of his equation, if the particulate mass is large, the uncertainties in both variables are incredibly small. At best, his equation tells us where an electron might have been, but not where it will be.

Heisenberg's Uncertainty Principle predicts the probability of finding electrons within each electron's energy affinity, i.e., close to the nucleus, farther out, etc, i.e., the atomic orbital.

Erwin Schrödinger derived his wave equation that requires the utilization of the quantum numbers of each electron in an element. His equation changed the view of electron orbitals from concrete, solar system-type to "fuzzy" regions with mathematical restrictions. Solutions to Schrödinger's Wave

Equations give 3 parameters called quantum numbers. Each set of quantum numbers describes an electron cloud wave function called an orbital. Orbitals are 3-D, geometric regions. Shown below are some representative orbital shapes -- the schematics are not inclusive, merely representative. These orbitals came about as a result of Schrödinger's Wave Equations.



### Quantum Numbers

The motions of the electrons about the nucleus can be described in terms of 4 (not just three as we've been discussing) quantum numbers:

Principal Quantum Number = "n"

Subsidiary Quant. Number, Secondary Quant. Number, or Angular Momentum Quantum Number = " $l$ "

Magnetic Quantum Number = "m"

Spin Quantum Number = "s"

### Quantum Numbers in Brief

1°	2°	shape	subshell	3°
n	l			m
1	0	Sphere	s	0
2	1	Dumbbell	p	-1,0,1
3	2	Ringed dumbbell; clover leaf	d	-2,-1,0,1,2
4	3	Ringed dumbbell	f	-3,-2,-1,0,1,2,3

### Principal Quantum Number = "n"

Principal Quantum Number = "n": designates the effective volume of the space in which the electron moves; increased "n" means increased energy with the shell; values vary theoretically between 1 and infinity ( $\infty$ ).

Shell	"n"
K	1
L	2
M	3
N	4
O	5
P	6

### Secondary Quantum Number = "l"

Subsidiary Quantum Number; Secondary Quantum Number; Angular Momentum Quantum Number = "l": designates the shape of the region which the electron occupies; values vary from 0 to ("n"-1): If "n" = 1, l = 0 ONLY; If "n" = 2, l = 0,1, ETC:

"n"	l	"n"	l	"n"	l
1	0	3	0,1,2	5	0,1,2,3,4
2	0,1	4	0,1,2,3	6	0,1,2,3,4,5



### Secondary Quantum Number: Application

$\ell$	Electron designation
0	s
1	p
2	d
3	f
4	g
5	h
6	i

### Magnetic Quantum Number = "m"

Magnetic Quantum Number = "m": designates in a general way the orientation of the electron path in space; values vary from  $-\ell, 0, +\ell$ .

"n"	$\ell$	"m"	"n"	$\ell$	"m"
1	0	0	4	0	0
2	0	0		1	-1,0,+1
	1	-1,0,+1		2	-2,-1,0,+1,+2
3	0	0		3	-3,-2,-1,0,+1,+2,+3
	1	-1,0,+1			
	2	-2,-1,0,+1,+2			

### Spin Quantum Number = "s"

The fourth quantum number, "s", or Spin Quantum Number is independent of the Schrödinger wave equation. The first three quantum numbers describe the orbitals in which the electrons may be; the last quantum number describes the electrons, specifically, that are in these orbitals.

There are two values for this number:  $\pm \frac{1}{2}$ . This value specifies the direction of spin of the electron about its own axis as the electron moves around the nucleus. Spin is clockwise or counter-clockwise (+ or -, respectively) and is designated the  $\frac{1}{2}$  arbitrarily -- probably due to Hund's rule.

(Hund's Rule Reminder:  $\equiv$  every orbital in a subshell is singly occupied with one electron before any one orbital is doubly occupied, and all electrons in singly occupied orbitals have the same spin. This means that in each "sub" subshell, there are only 2 electrons with equal BUT OPPOSITE SPIN once they begin to fill: E.g.: 2p:  $\uparrow\downarrow_x \uparrow\downarrow_y \uparrow\downarrow_z$ )

This fourth quantum number came about when Stern and Gerlach (in 1921) "shot a beam" of vaporized Ag atoms through a strong electromagnet and the vapor beam split into two. This indicated that electrons behaved like tiny magnets (remember the induced magnetic field, earlier) and deflected towards the opposite, or attracting, pole. This same sort of pattern also occurs with the mono-electronic atom, H, suggesting that each atom's electron has its own spin, i.e., there is no 100% identical spin of electrons even in H.

### Pauli Exclusion Principle

The Pauli Exclusion Principle says: No two electrons in a particular atom can have the same set of 4 quantum numbers, i.e., each electron has its own quantum number "fingerprint", i.e., each electron is unique.

### Quantum Number Summary & Application

Shell	"n" or "N" Primary Quantum Number	$\ell$ (ranges from zero to n-1) Secondary Quantum Number	Subshell	"m" (ranges from zero to $\pm \ell$ ) Magnetic Quantum Number	Shape	"s" (arbitrarily assigned as $\pm\frac{1}{2}$ ; independent of Schroedinger's Wave Equation) Spin Quantum Number
K	1	0	1s	0	Spherical	$\pm\frac{1}{2}$
L	2	0	2s	0	Spherical	$\pm\frac{1}{2}$
		1	2p	-1,0,+1	Dumb-bell	$\pm\frac{1}{2}$
M	3	0	3s	0	Spherical	$\pm\frac{1}{2}$
		1	3p	-1,0,+1	Dumb-bell	$\pm\frac{1}{2}$
		2	3d	-2,-1,0,+1,+2	Clover-leaf and ringed dumbbell	$\pm\frac{1}{2}$
N	4	0	4s	0	Spherical	$\pm\frac{1}{2}$
		1	4p	-1,0,+1	Dumb-bell	$\pm\frac{1}{2}$
		2	4d	-2,-1,0,+1,+2	Clover-leaf and ringed dumb-bell	$\pm\frac{1}{2}$
		3	4f	-3,-2,-1,0,+1,+2,+3	Ringed and Double-ringed dumb-bell	$\pm\frac{1}{2}$

1)  $n$  or  $N$  does NOT mean the number of electrons. It means that Schrodinger came up with a way to give each shell in an atom a number, hence the K shell's "number" is "1" – much like we humans have a name AND a social security number.

2) " $n$ " or " $N$ " only tells us which shell a single electron has been in.

3)  $\ell$  tells us the subshell that the electron has been in, e.g., s or p or d or f and it is dependent on the value of " $n$ " or " $N$ ", e.g.,  $l$  varies from zero to  $n-1$ , e.g., if " $n$ " is 3, then  $l$  varies from 0 to 1 to 2 (which is its maximum value because  $3-1=2$ ).

4) " $m$ " depends upon the value of  $l$ , and varies from  $-l$  to 0 to  $+l$ . This means that if  $m = 2$ , then  $l$  equals -2, -1, 0, +1, +2. Among other things, the value of " $m$ " is that the number of " $m$ " values you get essentially tells you what subshell the electron was in, in addition to the shape of that subshell.

5) The spin quantum number does not depend on ANY of these 3 quantum numbers, i.e., it's independent of the Schroedinger Wave Equation. Its numerical value is  $\frac{1}{2}$ . To determine if the spin of the electron is clockwise, we assign a "+" sign; counterclockwise, we assign a "-"; this is arbitrarily assigned unless we actually ran the experiment in the lab to confirm it.

So, if we were given the following set of quantum numbers, 2,0,0,+ $\frac{1}{2}$ , we can now give an electron an "address" of sorts that would describe where it had been:

**the quantum number series is ALWAYS in the order  $n,\ell,m,s$ , so we can identify the electron from or data.**

$N = 2$ ; this means that we're looking at an electron that was in the L shell.

$\ell = 0$ ; this means that we're looking at an electron that was in the s subshell.

$m = 0$ ; this confirms that we're looking at an electron that is in a spherical subshell (orbital), which is an s subshell.

$s = +\frac{1}{2}$ ; this means that the electron we're working to identify spins in a clockwise manner.

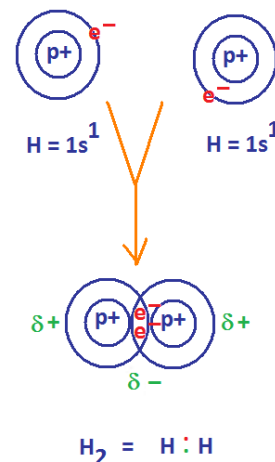
Succinctly, then, this is the quantum address for an L shell electron in a spherical s subshell that's spinning clockwise.

### Covalent Compounds

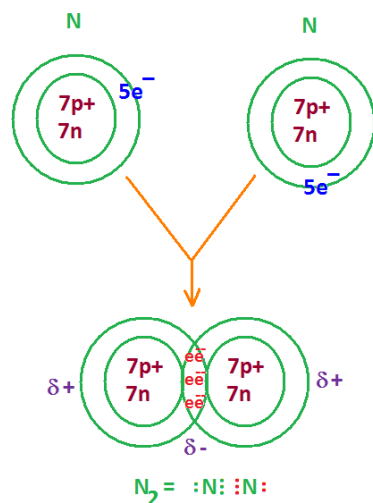
Continuing on our merry way, we've already discussed ionic bonding in a superficial sense, e.g., electrons are swapped between two entities to generate ionic bonds. We haven't really discussed a different kind of bond called a covalent bond. Covalent bonds share electrons, i.e., one electron from one element will be shared with one electron from another element to form the bond between the two.

Two electrons, then, are required to make a single bond between two atoms. That's concept 1 to walk away with prior to getting into this topic.

Covalent compounds are formed by a chemical bond that occurs between two atoms by sharing electrons -- NOT gaining or losing. The Figure at right illustrates a simple covalent bond between two atoms of hydrogen



(NOTE: hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine and iodine prefer to function as molecules in their elemental states held together by covalent bonds).



Note that at the bottom of the graphic that there are three distinct regions about the molecular hydrogen: there are 2 regions that are electron deficient and 1 region that is electron dense due to the electron overlap/sharing. The region that is deficient in electrons is said to be partially positive ( $\delta^+$ ; electropositive) and the region that is electron dense is said to be partially negative ( $\delta^-$ ; electronegative). This covalent bond still follows the octet rule.

The Figure at left shows the covalent bonds between two atoms of nitrogen to form molecular nitrogen. Note that there is a triple bond present -- note also that it takes 2 electrons to make each single bond.

When discussing covalent molecules, there are several terms one must know in order to fully appreciate the diversity of these molecules. These terms are summarized in the table, below:

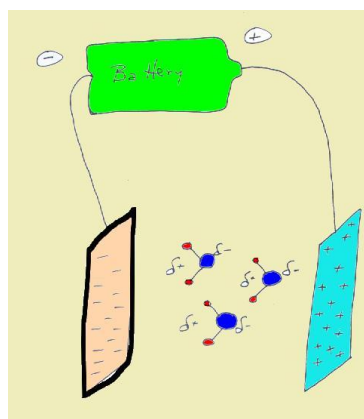
Nonpolar covalent bond	A covalent bond in which the bonding pair of electrons is shared equally by the bonded atoms.
Polar covalent bond	A covalent bond that shows bond polarization (bonding electrons are not shared equally).
Bond polarization	The attraction of the shared electrons to the more electronegative atom of a bonded pair of atoms.
Electronegativity	The tendency of an atom to attract shared electrons of a covalent bond.
Dipolar force	The attractive force that exists between the positive end of one polar molecule and the negative end of the same molecule.
Polar molecule	A molecule with polarized bonds which results in asymmetrical distribution of charge.
Nonpolar molecule	Molecules with nonpolarized bonds which results in symmetrical distribution of charge.

Remember from previous discussions that the most electronegative element on the periodic table is fluorine; the most electropositive element is Francium. Utilizing the above definitions and the previous statement, let's see if we can determine by two methods if a molecule contains a polar or non-polar covalent bond.

The first method is the "eyeball" method. Let's consider three molecules: ICl, Br<sub>2</sub> and CO. Let's set them side-by-side in the table, figure top of next page, and walk through this method a step at a time.

	ICl	Br <sub>2</sub>	CO
<b>Step 1: identify most electronegative atom</b>	Cl	Br and Br	O
<b>Step 2: draw electrical orientation</b>			
<b>Step 3: draw direction of polarization</b>			
<b>Step 4: identify type of bond (not James, either)</b>	Polar covalent bond	Non-polar covalent bond	Polar covalent bond

Another way in which to think of polar molecules is to think of water with two electrodes suspended in it attached to a battery. Polar molecules align themselves electronically, i.e., by charge, in an electrical field: the partially positive part of the molecule (red hydrogens) is aligned with the negative electrode and the partially negative part of the molecule (blue oxygens) is aligned with the positive electrode -- "in essence" closing the circuit.



The second method requires the utilization of a chart of known electronegativities:




		H 2.1					
Li 1	Be 1.5		B 2	C 2.5	N 3	O 3.5	F 4
Na 0.9	Mg 1.2		Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.9	Ca 1.0		Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0		In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9						

The chart is not complete; it will suffice for this discussion. As far as I'm concerned, the units are unimportant. It's the concept that I want you to get. The value of this method is that the extent of bond polarization is proportional to the difference in electronegativity ( $\Delta EN$ ) between the bonded atoms. When computed, a number is obtained that fits into one of the three ranges tabulated, below:

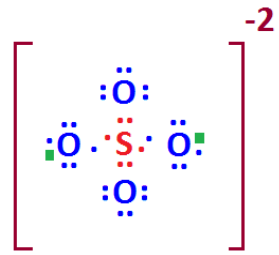
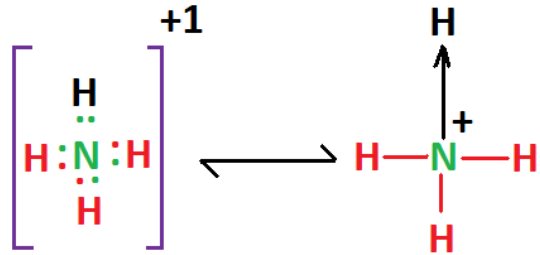
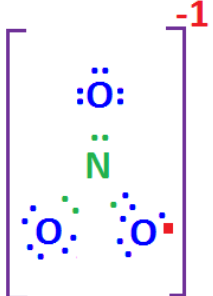
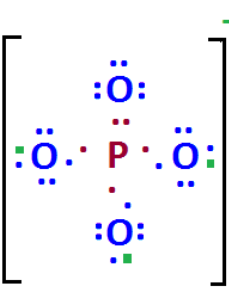
Range	Bond type
0	Non-polar covalent
0 - 2.0	Polar covalent
>2.0	Ionic

Let's run three examples as we did, before, only this time compare the "eyeball" method with this electronegativity method. We'll use ClF, MgO and  $PI_3$  as our examples, as well as to illustrate the process.

	ClF	MgO	$PI_3$
Most electronegative atom	F	O	I
Electrical orientation	$\delta^+$ $\delta^-$ Cl — F $e^- e^-$	$\delta^+$ $\delta^-$ Mg — O $e^- e^-$	

Polarization direction			
Bond type by eyeball	Polar covalent	Polar covalent	Polar covalent
$\Delta EN$	4-3=1	3.5-1.2=2.3	(2.5-2.1)*3=1.2
Correct bond-type	Polar covalent	ionic bond	Polar covalent

Binary compounds are pretty simple. Not all covalent substances are as simple. Polyatomic ions are ions that are bound covalently to each other with a net charge. Most polyatomic ions are negatively charged. Examples of polyatomic ions are tabulated, below:

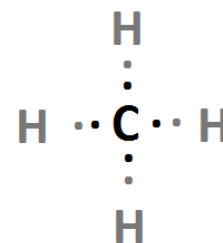
Name	Formula	Name	Formula
	SO <sub>4</sub> <sup>2-</sup>		NH <sub>4</sub> <sup>+</sup>
Sulfite	SO <sub>3</sub> <sup>2-</sup>	Bicarbonate	HCO <sub>3</sub> <sup>-</sup>
	NO <sub>3</sub> <sup>-</sup>		PO <sub>4</sub> <sup>3-</sup>
Nitrite	NO <sub>2</sub> <sup>-</sup>	Arsenate	AsO <sub>4</sub> <sup>3-</sup>
Carbonate	CO <sub>3</sub> <sup>2-</sup>		

The ammonium ion is a bit different from the other polyatomic ions -- besides being the only cation mentioned. It also contains a coordinate covalent bond. A coordinate covalent bond is a bond between 2 atoms in which one atom donates the 2 electrons (both electrons) to make the bond.

In the case of the ammonium ion, the coordinate covalent bond occurs between the unbonded pair of electrons in nitrogen and the extra hydrogen ion (proton) to make the polyatomic cation. A coordinate covalent bond is drawn as an arrow going from the two electrons' source to the electron acceptor. Remember that ammonia is neutral; the ammonium ion has the extra charge due to an extra proton and no electron to balance it out.

The second concept to walk away with prior to going through orbital hybridization is to accept as fact that all orbitals are present around each atom. Only those orbitals closest to the nucleus in sequential order fill so as to create an electron "cloud" around nucleus at optimal energies. The remainder of the orbitals is the topic for further discussion in another course.

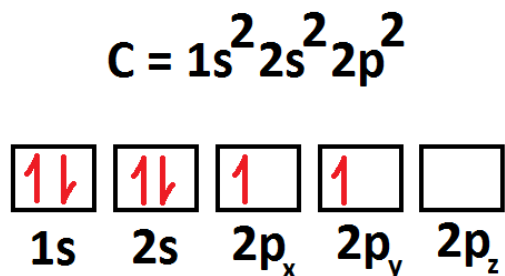
The third concept to walk away with is that when we speak of orbital hybridization, we're talking about hybridization in the same case as if we were to cross a pure red flower with a pure white flower to obtain the hybrid, the pink flower. In other words, the hybrid is somewhere in between the pure -- or elemental -- states. In the case of hybrid orbitals, the new orbitals will have energies that are lower than an energy-rich pure state and higher than an energy-poor state.



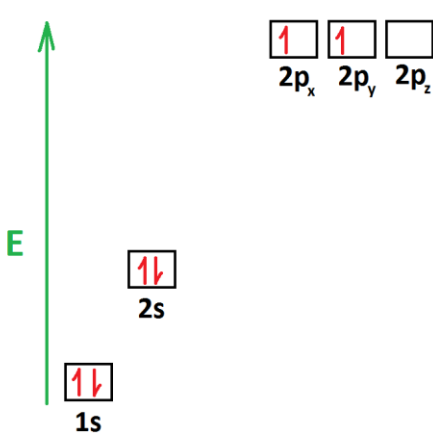
### Orbital Hybridization and Molecular Geometry

With that introduction, let's get started understanding how the sharing of electrons for the formation of a covalent bond works. As our first example, we'll examine methane, Figure, above right. Note that in this representation of methane,  $\text{CH}_4$ , that each bond consists of 2 electrons -- one electron from C (.) and one electron from H (.). These electrons in the bond are shared, remember. So, how does this happen?

The figure, right, illustrates the ground, or elemental, electronic configuration of carbon and of hydrogen. Note that, in carbon, that there are two sets of paired electrons and two sets of unpaired electrons. Remember that carbon is in Group IV on the periodic table and will make four bonds. Remember, too, that these 4 electrons have to be shared and that that will not happen until the second set of paired electrons is "split up" so that there are 4 unpaired electrons to share to make 4 single bonds with 4 hydrogen atoms.



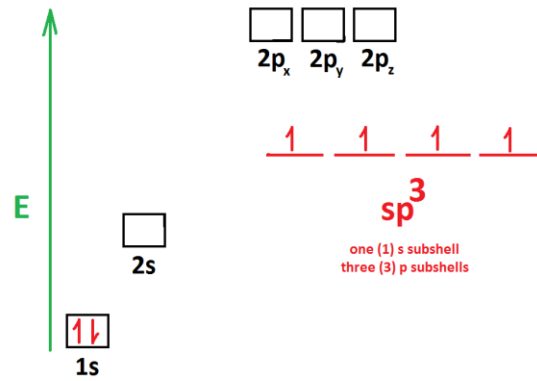
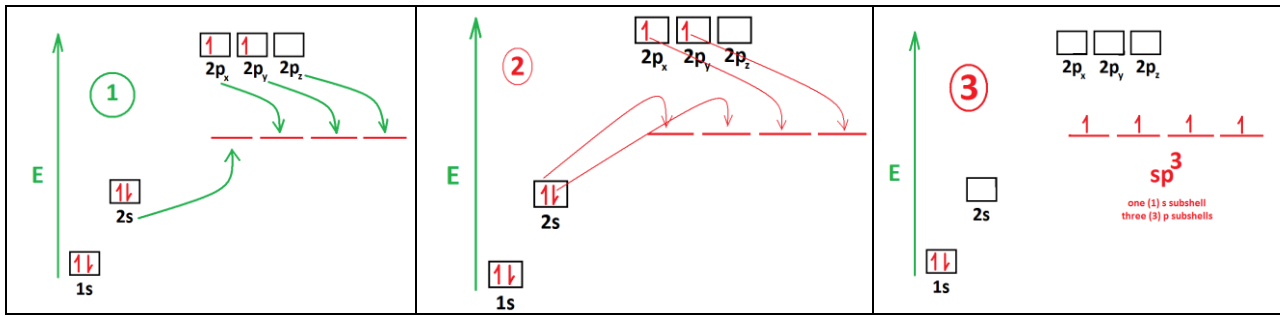




Figure, left, shows a representation of the ground state orbitals in terms of energy moving away from the nucleus (bottom of graphic) out (top of graphic).

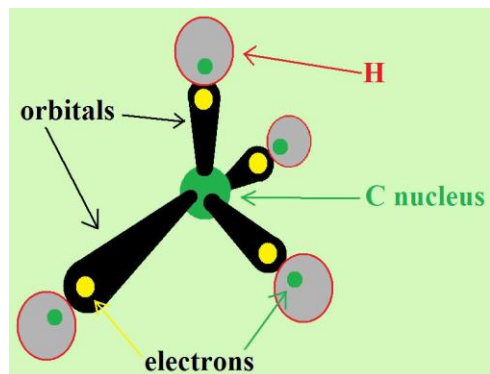
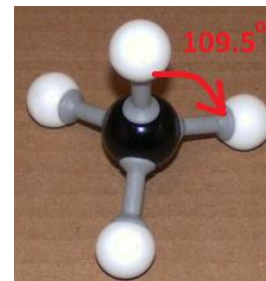
Figures, below, show what has to happen for the two 2s electrons to separate. In brief, the 2s and all three 2p sub-orbitals must gain and lose, respectively, energy **(1)** in order to force the separation of the two 2s electrons into unpaired electrons at the same energy as the two 2p electrons. In order for this to happen, the three 2p orbitals give up energy to drop down **(2)**; the 2s orbital gains the energy to come up a bit and to split the two 2s electrons up **(3)**.

The result is seen, below. While the 2s and 2p orbitals are shown as empty, the key is to remember that they are no longer "around" and that there is a new hybrid orbital. That new orbital contains all 4 electrons, unpaired, and is called an  $sp^3$  hybridization. An  $sp^3$  hybridized orbital is so called because one of the s orbitals and three of the p orbitals underwent energy differences (hybridization) to



accommodate the splitting up of the elemental 2s electrons. This is the hybridization that carbon undergoes when it has 4 single bonds around it.

Once we alter the orbitals, the shape (geometry) changes, as well. The shape of one  $sp^3$  orbital is teardrop. Carbon has four of them. When they are arranged around the carbon nucleus, there are  $109.5^\circ$  between each orbital. What this means is that the actual shape of an  $sp^3$  hybridized atom is that of a tetrahedron -- a three-legged milking stool with a flag-pole. Note, too, that another way in which to remember the geometry is to add up one s shell with three p shells to get four total shells.

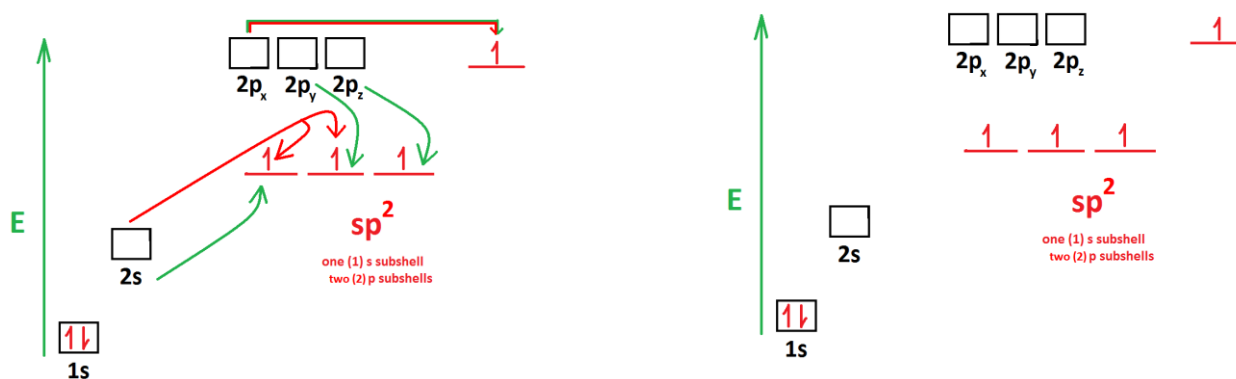


These four shells correspond to the placement for 4 individual electrons at the top, more or less, of each inverted teardrop-shaped orbital. BTW: I use the terms shells and orbitals interchangeably.



Not all carbon compounds have all single bonds about the carbon atoms. Some have double bonds. If a single bond consists of 2 shared electrons, it follows that a double bond consists of 4 shared electrons, i.e., 2 electrons per bond still. The simplest carbon compound that contains a double bond between carbon atoms is ethylene (common name), right. In order for the carbon to form these two bonds between each carbon atom, a different hybridization has to occur.

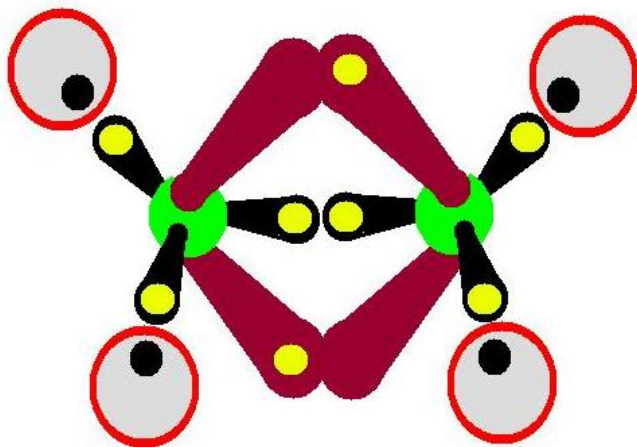
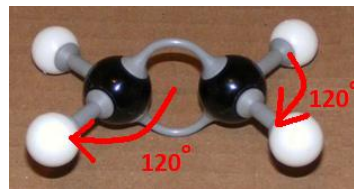
The difference between hybridization to make four single bonds around carbon and to make a double bond and two single bonds around carbon is the rearrangement of the electrons:



Note that in this case, that two of the 2p orbitals give up energy so that the 2s orbital may gain it and split up the two 2s electrons. During this process, one of the 2p electrons remains in a "standard" 2p sub-shell, below. This is significant and we'll touch on that in a moment.

When carbon has two single bonds and one double bond on it, it is in  $sp^2$  hybridization. The shape about the carbon atoms in this hybridization, given 3  $sp^2$  orbitals, is triangular, planar triangular or

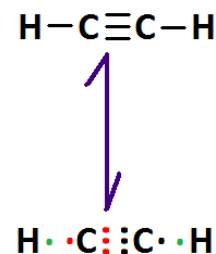
trigonal planar, right. Each hybrid orbital is separated by  $120^\circ$  bond angles. As with the  $sp^3$  hybrid, one may also derive the number of electrons from the hybridization, i.e., one s orbital and two p orbitals hybridize to separate 3 electrons. The fourth electron will be discussed, shortly.



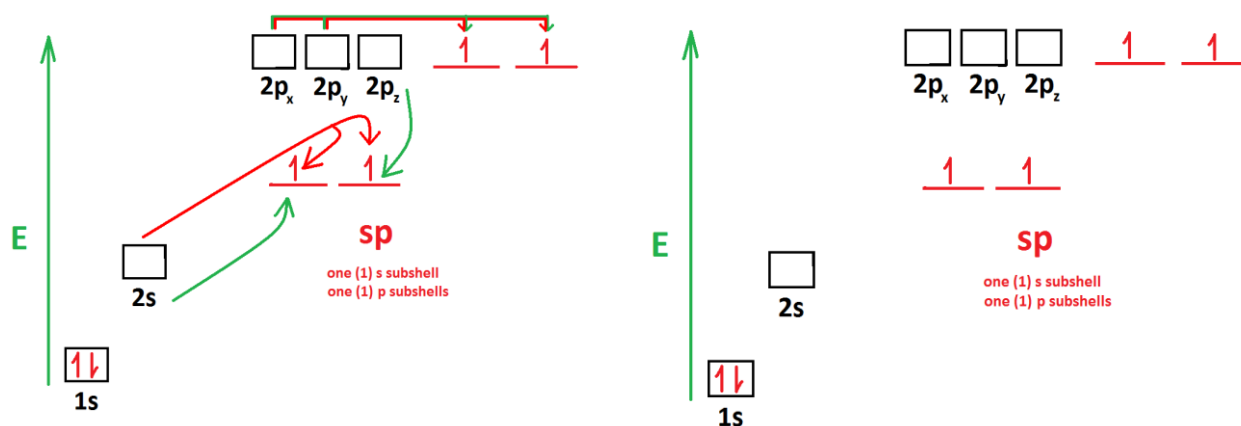
$sp^2$  hybridization accounts for only one of the two bonds in the double bonds. The second bond in the double bond comes from the lone 2p electron. Here's how it works: the p orbital's geometry is dumb-bell shaped. When it is overlapped with the  $sp^2$  hybrid orbitals, right, two kinds of bonds occur between the carbon atoms. The first is an end-to-end bond where the orbitals butt up against the other. This kind of bond is called a sigma ( $\sigma$ ) bond and comes from the  $sp^2$  hybrids.

The second bond to make the double bond comes from side-to-side, top-to-top and bottom-to-bottom overlap of the p orbital. This bond is called a pi ( $\pi$ ) bond. To reiterate, the first of the double bond is  $sp^2$ ; the second of the double bond is p. In order for the second bond to form completely, there must be overlap from top and bottom. This becomes very important in the lectures on organic chemistry that focus on aromatic compounds.

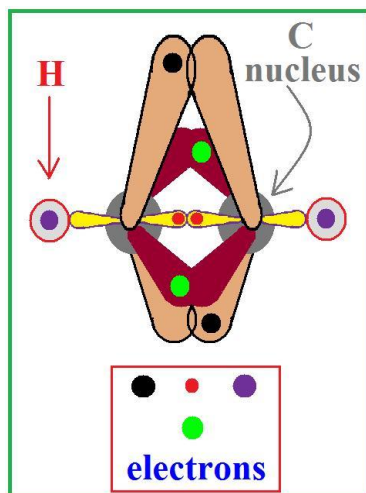
The third, and last, carbon-based bond we'll examine is the triple bond. The simplest carbon-based compound that has a triple bond (three bonds; total of 6 electrons shared) is acetylene (common name), Figure, right.



While there are still four bonds on each carbon atom, three of those bonds are between the two carbon atoms. Figure, above, illustrates the mechanics that must occur in order for this new hybridization to actuate.



Note that in this case, the 2s orbital absorbs the energy given off by one of the 2p orbitals to give the new hybrid, below. This new hybrid is called an **sp** hybrid; note that the two 2p electrons remain in "standard" 2p orbitals. When a carbon atom has one single bond and one triple bond about it, it is said to be in sp hybridization. What about the shape of the molecule? It is linear, i.e., the geometry about the two carbon atoms is arranged in a straight line with 180° bond angles between them. As the shortest distance between two points is a straight line, the sp orbital (one s shell and one p shell) may be easily remembered as  $1+1 = 2$ .

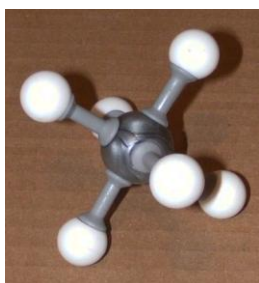


sp hybridization accounts for only one of the three bonds about the two carbon atoms in the triple bonds. The actual sp hybridized orbital provides electrons for the end-to-end sigma bonds; the remaining two 2p orbitals provide the last two sets of electrons to form pi bonds to "seal it up", Figure, left.

Carbon is not the only atom to undergo orbital hybridization in order to make bonds. In many cases the hybridization pattern may be concluded by examining the periodic table and counting electrons in outer shells, e.g., Be is in Group II and undergoes sp hybridization yielding linear (aka digonal) geometry, B is in Group III and undergoes  $sp^2$  hybridization yielding trigonal planar geometry.

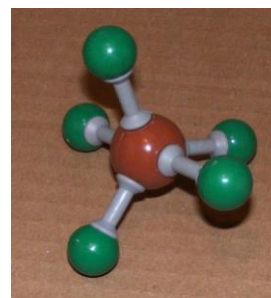


Less obvious are Pt, Pd and Ni which undergo  $dsp^2$  hybridization to give a square planar geometry (upper right) -- note that  $sp^3$  and  $dsp^2$  hybrids are distinctly different hybridizations.



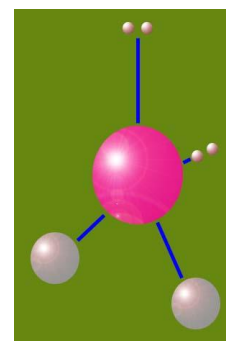
Likewise, it's not always possible to determine that P (in a +5 state) undergoes  $dsp^3$  hybridization to yield trigonal bipyramidal geometry, Figure, right. Do you see the 5 bonding sites?

Or that S (in a +6 state) undergoes  $d^2sp^3$  hybridization to give an octahedral geometry, Figure, left. Do you see all 6 binding sites?



A strange hybridization requires some additional discussion. This has to do with how oxygen hybridizes to form water. Oxygen is in group VI on the periodic table and its electronic configuration is  $1s^2 2s^2 2p^4$ , where there are 2 unpaired 2p electrons. In the older literature, oxygen was said to have undergone  $p^2$  hybridization. This gave a geometry around the oxygen in water that was angular.

The newer literature shows that oxygen undergoes a strange  $sp^3$  hybridization. This is a slightly different look at  $sp^3$  hybridization than what we have become

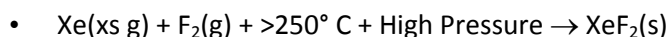


accustomed to. When we examined it earlier, we saw that there were 4 unpaired electrons at each  $sp^3$  hybrid orbital. In the case of oxygen in water, there are two  $sp^3$  hybrids with an electron each and two  $sp^3$  hybrids with one PAIR of electrons, each, giving the tetrahedral geometry, below right, previous page. The tetrahedral geometry better explains the surface tension of water and properties of ice (floats) than does the angular geometry. Be aware that both explanations are still "out there", though. More coming in the lecture on water.

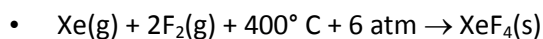
Returning to the statement that sent us off through Quantum Theory and Wave Mechanics, Xenon is water-soluble and has found some use in compounds as an anesthetic. Xe is capable of attaining one of three hybridizations:  $sp$ ,  $dsp^2$  or  $d^2sp^3$ .

The following reactions illustrate these hybridizations:

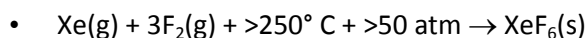
1.  $sp$  hybridization:



2.  $dsp^2$  hybridization:



3.  $d^2sp^3$  hybridization:



Xe and Kr have many compounds. These compounds are generally involved with fluorine, oxygen and nitrogen, which are the most electronegative elements on the periodic table. While He and Ne have no known compounds, recently, reports have been coming out about new compounds with Ar.

## Problem Set 2

1. Ba has how many protons? Electrons?
2. B has how many electrons in its valence shell?
3. Sr has how many electrons in its valence shell?
4. Write the electronic structure for Al.
5. Write the electronic structure for Be.
6. Write the electronic structure for S.
7. Write the electronic structure for Ca.
8. Write the electronic structure for the sodium ion.
9. Write the electronic structure for the chloride ion.

10. Write the electronic structure for the  $P^{3-}$  ion.
11. Write the electronic structure for the  $Be^{2+}$  ion.
12. Write the electronic structure for the fluoride ion.
13. Write the electronic structure for the  $H^+$  ion.
14. Given the following sequence of quantum numbers: 1, 0, 0,  $+\frac{1}{2}$ , "identify" the element (this is the only one you can do this with, BTW).
15. Given the following sequence of quantum numbers, 2, 1, -1,  $-\frac{1}{2}$ , describe the electron.
16. Given the following sequence of quantum numbers, 2, 0, +1,  $+\frac{1}{2}$ , describe the electron.
17. The "s" subshells may hold a maximum of \_\_\_\_\_ electrons.
18. How many "sub-subshells" are in the "p" subshells?
19. How many "sub-subshells" are in the "d" subshells?

#### Problem Set 4

1. Identify which of the following reactions are reduction or oxidation reactions:

- |                                      |                                    |                                       |
|--------------------------------------|------------------------------------|---------------------------------------|
| A) $Cu^{2+} + 1e^- \rightarrow Cu^+$ | B) $Fe \rightarrow Fe^{2+} + 2e^-$ | C) $Zn \rightarrow Zn^{2+} + 2e^-$    |
| D) $Mg^{2+} + 2e^- \rightarrow Mg$   | E) $Fe \rightarrow Fe^{3+} + 3e^-$ | F) $Hg_2^{2+} + 2e^- \rightarrow 2Hg$ |
| G) $Li \rightarrow Li^+ + 1e^-$      | H) $Be^{2+} + 2e^- \rightarrow Be$ | I) $F_2 + 2e^- \rightarrow 2F^-$      |
| J) $Al \rightarrow Al^{3+} + 3e^-$   |                                    |                                       |

2. Given the following reactions, determine which reactant is oxidized and which reactant is reduced:

- |  |  |
|--|--|
| A) $Mn + Zn^{2+} \rightarrow Mn^{2+} + Zn$ | B) $2Al^{3+} + 3Pb \rightarrow 2Al + 3Pb^{2+}$ |
| C) $Na + Cl \rightarrow Na^+ + Cl^-$       | D) $2Li + O \rightarrow 2Li^+ + O^{2-}$        |
| E) $2Na + S \rightarrow 2Na^+ + S^{2-}$    |  |

3. Based on your responses to #2, above, which reactant is the oxidizing agent and which reactant is the reducing agent?

4. Based upon your new understanding of "simple" redox reactions, balance the following reactions:

- |  |  |
|--|--|
| A) $Cu^{2+} + Zn \rightarrow Cu + Zn^{2+}$ | B) $Fe + Fe \rightarrow Fe^{2+} + Fe^{3+}$ |
| C) $Be^{2+} + Al \rightarrow Be + Al^{3+}$ | D) $Zn + Mg^{2+} \rightarrow Zn^{2+} + Mg$ |



### Problem Set 6

1. Write out the names of the following ions:

- A)  $\text{SO}_4^{2-}$       B)  $\text{PO}_4^{3-}$       C)  $\text{NO}_3^-$       D)  $\text{NO}_2^-$   
E)  $\text{S}^{2-}$       F)  $\text{Cl}^-$       G)  $\text{F}^-$       H)  $\text{I}^-$       I)  $\text{O}^{2-}$       J)  $\text{NH}_4^+$

2. Using Lewis structures, draw  $\text{CH}_4$ .

3. Using Lewis structures, draw  $\text{MgS}$ .

4. Using Lewis structures, draw  $\text{CCl}_4$ .

5. Using Lewis structures, draw  $\text{AlF}_3$ .

6. Using Lewis structures, draw  $\text{MgF}_2$ .

### Problem Set 11

1. Draw the following hybridizations (geometrically) and label them with their names:

- A)  $sp$       B)  $sp^2$       C)  $sp^3$       D)  $dsp^3$       E)  $d^2sp^3$

2. Predict the hybridization of the following elements upon reaction:

- A) B      B) Be      C) Ca  
D) C (all single bonds)  
E) C (1 double bond and 2 single bonds)      F) C (1 triple bond and 1 single bond)  
G) Xe (in  $\text{XeF}_6$ )      H) Al (in  $\text{AlF}_3$ )      I) N (in  $\text{NH}_3$ )

3. Write the electronic structures graphically for the following:

- A)  $\text{B}^0$       B)  $sp^2$  hybrid of B      C)  $\text{C}^0$       D)  $sp^2$  of C  
E)  $sp^3$  hybrid of C      F)  $sp$  hybrid of C      G)  $sp^2$  hybrid of N  
H)  $sp^3$  hybrid of N      I)  $\text{N}^0$       J)  $\text{Ne}^0$