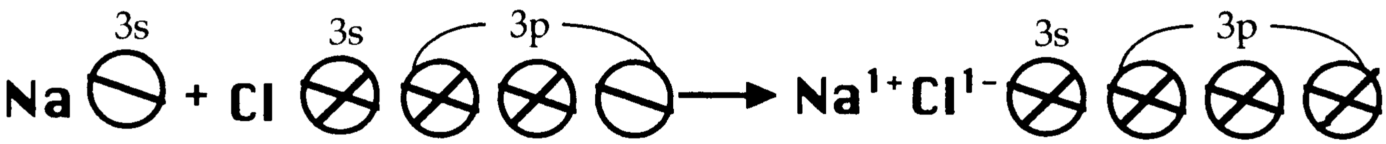
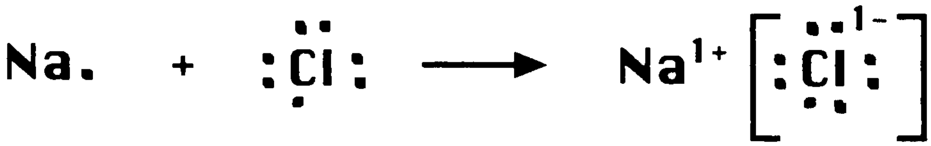
Electron Arrangement

$Na\ 1s^2\ 2s^2\ 2p^6\ 3s^1 + Cl\ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5 \longrightarrow Na^{1+}\ 1s^2\ 2s^2\ 2p^6 \quad Cl^{1-}\ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$





and Chemical Periodicity

Unit 7

Unit 7: Electron Structure & the Periodic Table

I. The Periodic Table......a Search for Similarities

By the mid-1800s, more than 60 elements had been identified and described by their physical and chemical properties. It was becoming increasingly more difficult to keep track of their individual characteristics and behaviors, and the search began for any similarities or patterns that may exist. Early attempts at classification tried to link the atomic masses of the elements with related properties, but were not very successful in predicting properties of other elements. In 1860, a scientific meeting was held to standardize the methods used to calculate atomic and molecular masses throughout the world. Once all masses were calculated by comparing to a common standard, it was possible to accurately induce any relationships between the elements' masses and properties. As an understanding of atomic structure developed some fifty years later, chemists were also able to use these theories to explain the similarities and differences in chemical and physical behavior.

A. In 1863, John Newlands arranged the known elements in order of their increasing atomic masses. He noticed that there appeared to be a repetition of properties every eighth element(Note that the Group VIIIA(8A) inert gases were not yet discovered at this time) and referred to his arrangement as the Law of Octaves. This pattern worked for the lighter elements up to calcium, but then began to fall apart.

Example:	Li	Be	B	C	N	O	F
	Na	Mg	Al	Si	P	S	Cl
	K	Ca	etc.				

B. Dmitri Mendeleev revised this idea in 1869, and suggested that the elements should be placed in periods (horizontal rows) of various lengths in order to match the pattern of repeating similar physical and chemical properties. He left several blank spaces on his table, because there were no known elements whose properties fit those predicted to occur. Mendeleev used his arrangement to predict the characteristics of the missing elements, such as Ga, Ge and Sc, with amazing accuracy. The table showed that the properties of the elements vary in an orderly fashion as a periodic function of their atomic masses, which is called the <u>periodic law</u>. There were a few minor discrepancies with this table when the elements were arranged strictly according to increasing atomic masses(Check the positions of Te compared to I on the table). Mendeleev knew that similar properties were the most important factor in placing the elements on the table, and he thought that the atomic masses of Te and I were incorrectly determined(However, new measurements simply confirmed them as being correct).

Problem:
Name at least two more pairs of elements that are out of order, if the table were arranged strictly according to increasing atomic masses?

C. Henry Moseley resolved the discrepancies in Mendeleev's table caused by the atomic mass values when he calculated the atomic number(nuclear charge) for each element in 1913. He showed that the nucleus has a whole-number positive charge, based on the number of protons present. When the elements are listed by increasing atomic number, all of the properties match those predicted by the element's position on the table. The modern periodic table is arranged by increasing atomic number.

D. Horizontal rows of elements on the table are called **periods** or **series**. The properties of elements in each period change from being metallic to nonmetallic to an inert gas, as you go from left to right across the table. The sequence of change remains the same for each of the 7 different periods.

E. Vertical columns of elements are called **groups** or **families**. All the elements in a particular group have similar physical and chemical properties. The representative elements that illustrate the entire range of chemical properties for all elements are designated in the U.S. system of labeling as Groups 1A through 8A. The subgroups are designated as B groups. (Note: For European countries, the A and B group designations are opposite of the U.S. system. A compromise has been suggested using numbers from 1 to 18 to label each of the groups)

II. Similarities in Atomic Structure

Similarities and patterns of change in the properties of the elements in a particular period or group can be related to the similarities and patterns of change in the atomic structure for the elements. Chemical reactions involve the exchange or sharing of the outermost electrons in atoms, therefore an understanding of electron arrangements is needed to be able to accurately explain and predict properties of the elements. In today's quantum mechanical model of the atom, the electron is considered to have a wave-like motion, moving towards and away from the nucleus with a specific energy, frequency and wavelength. Erwin Schrodinger developed a mathematical equation that related the amplitude of the electron-wave to any point in space surrounding the nucleus. Solving this equation generates a 3-dimensional graph of the region in space around the nucleus where the probability of finding the electron is high. This region is called an (space) **orbital**. There are four quantum numbers that provide the "energy address" of each electron in an atom, much like your home address distinguishes your house from all others. No two electrons can have the exact same set of four quantum numbers(this is called the Pauli Exclusion Principle). Electrons try to obtain their lowest possible energy levels, called the **ground state**.

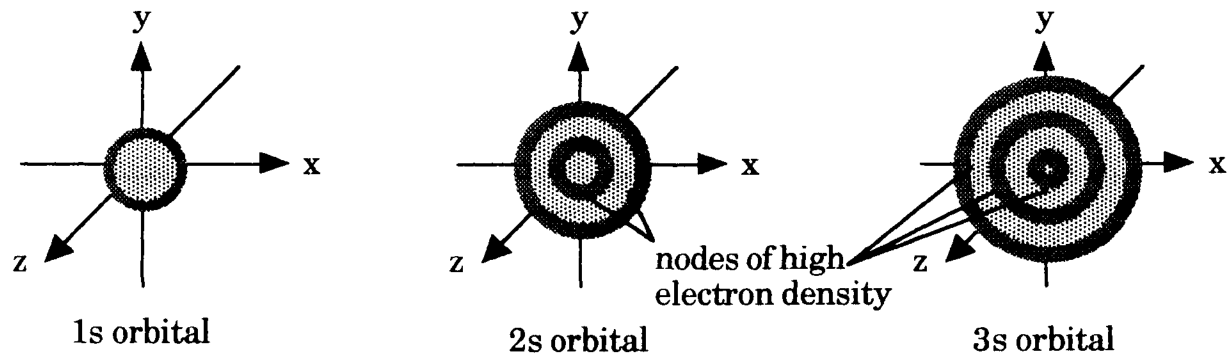
Quantum Numbers....the Electron's Home Address

A. Principal quantum number (n) - represents the main energy level(or shell) that the electron is occupying, ranging from levels 1, 2, 3,....to n. The electrons closest to the nucleus are in the lowest energy level, #1, and those farthest from the nucleus have the highest energy level. *The maximum number of electrons that can occupy a particular level is $2n^2$.* Therefore, level 1 can hold at most two electrons(2×1^2); level 2 an contain a maximum of 8 electrons(2×2^2); level 3 can hold 18 electrons (2×3^2); level 4 has a limit = 32 electrons, etc.

- All elements in the same horizontal row(period) on the periodic table have the same number of occupied main energy levels.

Examples: The elements in the second period from Li to Ne(atomic numbers 3 to 10) all have electrons in the first and second main energy levels.
The elements in the sixth period from Cs to Rn(atomic numbers 55 to 86) all have electrons in the first six main energy levels.

B. Orbital quantum number (l) - represents the shape of the space orbital that the electron occupies, which is an energy sublevel of the main energy levels. The shapes of these orbitals vary from spherical and dumbbell-shaped to more exotic four-leaf clover and octahedral structures. Sublevels are named **s, p, d and f** and can contain a maximum of 2, 6, 10 and 14 electrons, respectively. The number of different sublevels in a main energy level is equal to n. Therefore, level 1 has only the s sublevel; level 2 has the s and p sublevels; level 3 has s, p and d sublevels; level 4 has s, p, d and f sublevels, etc. The only difference between the same type of sublevel in different main energy levels is the size of the orbitals. Examine the comparison of the sizes of different s sublevels shown below:



Any orbital can hold a maximum of 2 electrons. Since the p sublevel has three separate orbitals, it will hold 6 electrons(the p sublevel limit). There are 5 orbitals in any d sublevel and 7 orbitals in any f sublevel.

- Elements in Groups 1A and 2A on the periodic table put their last electrons into an s sublevel. Those in Groups 3A to 8A put their last electrons into a p sublevel. The Group B elements(also known as the transition metals) begin to fill the d sublevels with their last electrons. The rare earth elements(also known as the lanthanide and actinide series) fill an f sublevel with their last electrons.

  Examples: All of the elements in Group 1A would put their last electron into an s sublevel. Hydrogen would have 1 electron in the s sublevel of the *first* main energy level. This is written as $1s^1$, which is called the **electron configuration** for a hydrogen atom. Lithium would first fill the 1s sublevel with two of its electron, then put its last electron into the s sublevel of the *second* main energy level. This is written as $1s^22s^1$. Sodium would fill up to the $3s^1$ sublevel, potassium would fill up to the $4s^1$ sublevel, etc.

  For the elements boron to neon, the first 4 electrons would fill up the 1s and 2s sublevels(written as $1s^22s^2$). The last electrons would enter the 2p sublevel and would have the configuration of $2p^1$, $2p^2$, $2p^3$, $2p^4$, $2p^5$, and $2p^6$, respectively.

- Elements with similar electron configurations are listed in the same groups(vertical columns).

Problem:

Fill in the supplied blank chart for each element with the electron configuration of the last sublevel that gets electrons. For example, write the configuration $3s^1$ in the blank for a sodium atom.

transition metals

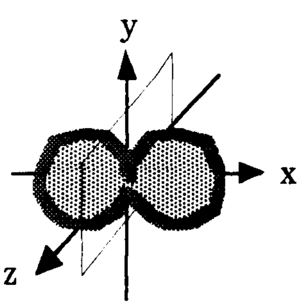
1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
1																	2
3	4											5	6	7	8	9	10
11	12											13	14	15	16	17	18
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
55	56		72	73	74	75	76	77	78	79	80	81	82	83	84	85	86
87	88		104	105	106	107											

rare earth elements

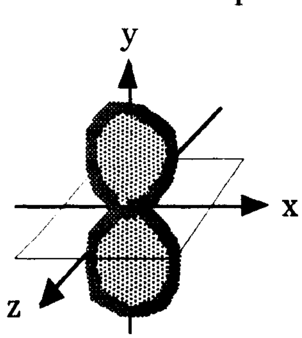
lanthanide series	57	58	59	60	61	62	63	64	65	66	67	68	69	70	71
actinide series	89	90	91	92	93	94	95	96	97	98	99	100	101	102	103

C. <u>Magnetic quantum number</u> (m) - describes the relative orientation of a particular orbital along the x, y and z axes of a 3-dimensional coordinate system. This number is used to distinguish different orbitals that are part of the same energy sublevel. Since an s sublevel consists of only one spherical-shaped orbital, the electron pair will be evenly distributed around the nucleus. The positions of the orbitals in the p, d and f sublevels need 3, 5 and 7 different magnetic quantum numbers, respectively, to specify which particular orbital the electron is occupying.

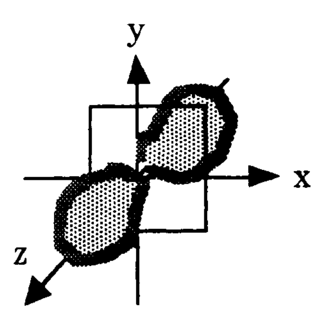
The p sublevel consists of 3 separate dumbbell-shaped orbitals at 90° angles to each other.



2p_x orbital

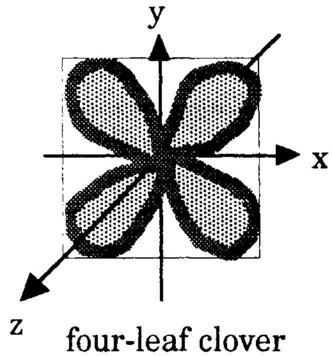


2p_y orbital

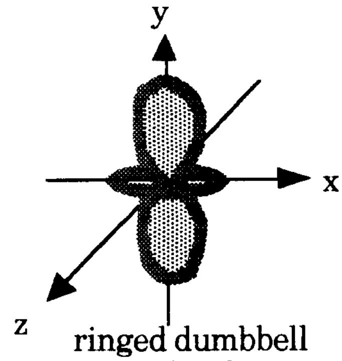


2p_z orbital

Four of the five d orbitals have the same four-leaf clover shape but different orientations in space. One o the d orbitals has a dumbbel shaped orbital with an extra ring of electron density around the middle(diagram a the far right).

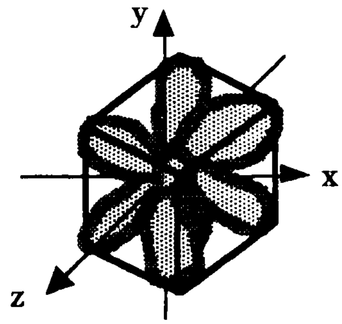


four-leaf clover
d orbital

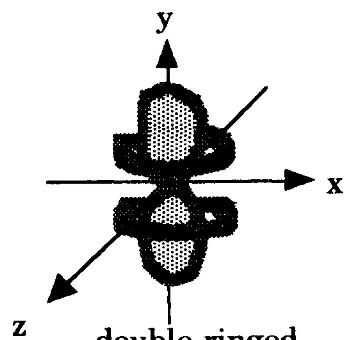


ringed dumbbell
d orbital

Four of the seven f orbitals have the same octahedral shape but different orientations in space. Three of the f orbitals have the same dumbbell-shape with two extra rings of electron density around the middle but different orientations in space. (diagram at the far right).



octahedral-shaped
f orbital



double-ringed,
dumbbell-shaped
f orbital

D. <u>Spin quantum number</u> (s) - electrons in the same orbital will spin in opposite directions. This is because a spinning electron creates a magnetic field around the atom with a north and south pole. (Remember that elementary school science project when you wrapped a wire around a nail and connected it to a battery to produce an electromagnetic? This is the same thing.) The second electron spins in the opposite direction to create a magnetic field having an opposite orientation of its north and south poles, which helps to hold the pair together. Since the electrons in the pair are both negatively charged, they have a natural tendency to repel each other.

Collectively, the four quantum numbers provide enough information to determine the position of every electron in an atom or ion. The numerical values assigned to the quantum numbers for each electron are used in solving Schrodinger's equation to generate the space orbital with a high probability of containing the electron. Since this equation is every graduate student's nightmare, chemists have developed other simplified methods based on the quantum number values to represent the energy levels of the various electrons. Each method emphasizes a particular aspect of the electron structure, such as the number of electrons in the highest main energy level.

III. Methods of Representing Electron Arrangements

The key points of these alternate methods are described on the next two pages. Note which specific aspect(s) of the electron's structure is provided by each method.

Electron Structure & the Periodic Table

The electron configuration for an atom shows the distribution of electrons in the various energy sublevels surrounding the nucleus. In the ground state, the electrons fill the sublevels according to the **Diagonal Rule** illustrated to the right.

$1s^2$
$2s^2\ 2p^6$
$3s^2\ 3p^6\ 3d^{10}$
$4s^2\ 4p^6\ 4d^{10}\ 4f^{14}$
$5s^2\ 5p^6\ 5d^{10}\ 5f^{14}$
$6s^2\ 6p^6\ 6d^{10}$
$7s^2\ 7p^6$

Group IA

H $1s^1$

Li $1s^2\ 2s^1$

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

K $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

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

Cs $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^1$

Fr $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^1$

The orbital notation and electron dot diagrams for all Group IA elements contain a single electron in the s orbital of successively higher main energy levels.

Example:

3s

Na.

Group VIIA

The orbital notation and electron dot diagrams for all Group VIIA elements contain 7 electrons in the s and p orbitals of successively higher main energy levels.

F $1s^2\ 2s^2\ 2p^5$

Cl $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Br $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$

I $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^5$

At $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^5$

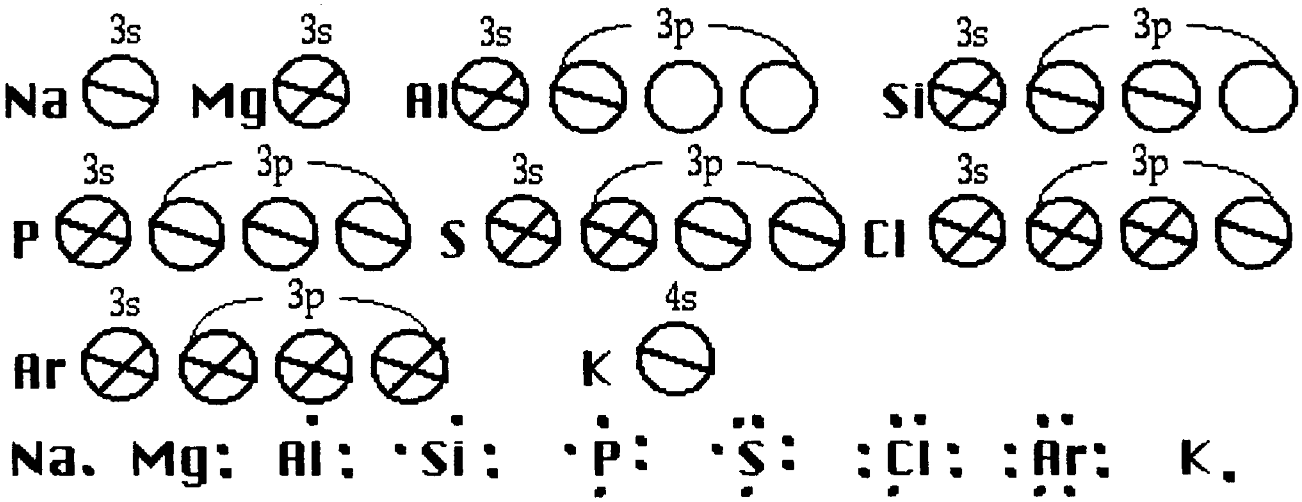
4s 4p

:Br:

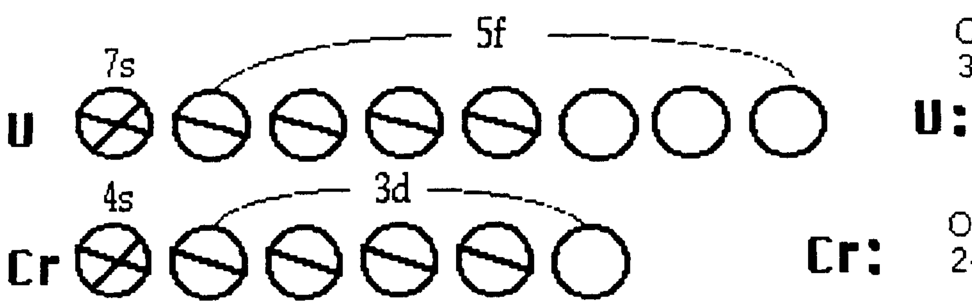
The number of **valence electrons** (those in the highest main energy level) in the atoms of a particular element are equal to the A Group number to which it belongs. Group IA through Group VIIIA elements contain 1 to 8 valence electrons, respectively.

Orbital Notation & Electron Dot Diagrams

Successive elements from Group IA to Group VIIIA across a period add an extra electron to the same main energy level.



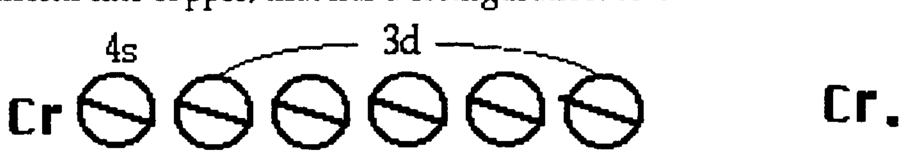
Note how the electrons entering the p sublevel enter different orbitals and have the same direction of spin. Pairing up does not occur until there is 1 electron per orbital first. This is called **Hund's Rule**. The octet of electrons in argon is very stable, because all of the electrons are paired up and their magnetic fields are balanced. With the next element, potassium, the pattern starts over again with the 4s orbital. Transition metals in the Group B block of the periodic table usually have a filled s sublevel as their highest energy level, and successive elements add electrons to a d sublevel of a lower energy level. The rare earth elements at the bottom of the chart have the same valence electron structure, but fill f sublevels with their last electrons. During chemical reactions, the s sublevel electrons are used first, but d and f sublevel electrons are also available for bonding. These d and f sublevel electrons are included in orbital filling diagrams only when the sublevel is partially filled, but they are never included in the electron dot diagrams. For example, the diagrams for uranium and chromium are shown below:



Oxidation states:
3+, 4+, 5+, 6+

Oxidation states:
2+, 3+, 6+

Although the predicted electron structure for chromium has a pair of electrons in the 4s orbital and 4 single electrons in the 3d sublevel, the actual structure is shown below. It has been determined that a half-filled 3d sublevel has lower energy and is more stable than the $4s^2$ arrangement. A similar exception to the pattern occurs with elements like copper, that has a configuration of $4s^1$ $3d^{10}$ rather than $4s^2$ $3d^9$.



Problems:
Write the electron configuration, orbital notation and electron dot diagrams for the following atoms or ions:

1. arsenic atom

2. iron atom

3. calcium ion

4. plutonium atom

5. iodideion

6. silver atom

7. barium atom

IV. Explaining Properties of the Elements Based on Their Atomic Structure

You should have noticed the patterns in electron structure that occurred in the last section. These periodic variations can be used to explain and predict the similarities and differences in the properties of the elements in a period or group. There are 3 factors that influence elemental properties:

1. Number of main energy levels(shells) - As the number of main energy levels increases going down a group, the size of the atoms increases. This causes the valence electrons to be farther from the nucleus, which means they are not held as tightly in the larger atoms of a particular group The lower level electrons create a "shielding effect" that block the pulling power of the positive nucleus on the electrons in the outermost shell.

2. Number and distribution of electrons in the outermost energy level(valence shell) - The Group VIIIA elements have a balanced distribution of 8 valence electrons(He has only 2e), which makes them chemically very stable. Elements in other groups will chemically react to gain, lose or share electrons to try to also obtain eight valence electrons(called the **Octet Rule**). Atoms or ions that have completely filled energy sublevels, as well as those with half-filled p, d or f sublevels, have also been found to have a more stable electron arrangement.

3. Net effective nuclear force(kernel charge) - This is a relative measure of the attraction between the protons in the nucleus and the valence electrons. The strength of this force is determined by subtracting the number of lower energy level electrons from the number of protons in the nucleus. As the kernel charge increases from +1 to +8 when going left to right across a particular period, the protons are able to pull the valence electrons in tighter. Note that most transition elements have a +2 kernel charge. Even though successive transition metals in a period(such as La to Hg) add an extra proton in the nucleus, they also add an extra electron to an inner shell which increases the shielding effect.

Problems:
Use your periodic table to determine the number of shells, valence electrons and kernel charges for the following atoms or ions:

1. lead atom

2. francium atom

3. krypton atom

4. rubidium ion

5. bromide ion

V. Trends in Periodic Properties of the Elements

Due to the changes in the number of shells, valence electrons and kernel charges described in the previous sections, the following trends in the periodic properties of the elements can be observed:

A. Metals vs. Nonmetals

- In general, elements with 3 or less valence electrons are classified as metals, which try to lose valence electrons to become stable. *Francium* is the most reactive metal because it loses its single valence electron the easiest due to its large size(7 shells) and small kernel charge(+1).
- Elements with 5 or more valence electrons are classified as nonmetals which prefer to gain electrons to acquire an octet, but will also share electrons to become more stable. *Fluorine* is the most reactive nonmetal because it gains a single valence electron the easiest due to its small size(2 shells) and large kernel charge(+7).
- The elements in Groups 3A, 4A and 5A vary from being nonmetallic at the top of the group to metallic at the bottom of the group as the atomic mass increases. Some of these elements are classified as **metalloids**, which have some properties characteristic of both a metal and a nonmetal. Metals lose electrons during chemical reactions forming positively charged ions.
- Typically a metal is a shiny solid that is a good conductor of heat and electricity, and can be easily bent and shaped without breaking. Nonmetals have the opposite properties, existing as gases or easily vaporized solids and liquids. They oxidize(burn) when heated in the air and absorb electricity, rather than conduct it. Nonmetallic solids are brittle, and will shatter when bent or hammered on. Nonmetals try to gain electrons in reactions to form negative ions, but they can share electrons to obtain an octet.

B. Atomic and Ionic Radii

- The electron orbitals have no sharply defined boundaries to provide an exact limit to the atom's size. However, there are ways to estimate the relative radius of an atom, such as using a technique called x-ray diffraction.
- The size of atoms generally increases going down a group due to the increased

number of shells and the greater shielding effect caused by additional lower energy level electrons.

- Going left to right across a period of elements, the size of the atoms generally decreases. This is due to an increased kernel charge from Group IA to Group VIIIA which has a greater pull on the same number of electron shells.
- When metals lose electrons to form positive ions, the ions always get smaller in size. When nonmetals gain electrons to form negative ions, the ions always get larger in size.

Example 1: Sodium atoms have 11 protons and an electron configuration of $1s^22s^22p^63s^1$. When they lose 1 electron to become isoelectronic with neon, there are still 11 protons, but the configuration becomes $1s^22s^22p^6$. An entire shell is lost and the kernel charge of the ion jumps to + 9, shrinking the atom to about half of its original size.

Example 2: Oxygen atoms have 8 protons and an electron configuration of $1s^22s^22p^4$. When they gain 2 electrons to become isoelectronic with neon, there are still 8 protons, but the configuration becomes $1s^22s^22p^6$. The kernel charge and number of shells remains the same, but the two additional electrons in the valence shell are harder to hold together which causes the ion to expand in size.

Problems:
State which particle is larger and explain the reasons for your choice.

1. Ar or Xe	2. Li or C	3. P or P^{3-}
4. Ne or Mg^{2+}	5. N or As	6. Cl or O
7. Ca or V	8. I^{1-} or Te^{2-}	9. Ti or Ti^{4+}

ANS: Xe, Li, P^{3-}, Ne, As, Cl, Ca, Te^{2-}, Ti

C. Ionization Energy, Electron Affinity and Electronegativity

The strength of the attractive forces holding the valence electrons in an atom can be measured from several different perspectives. These include how difficult is it to remove an electron from an atom, the change in energy that occurs when an atom gains an electron or the relative tendency to take an electron away from another atom. As a general rule, metals have less attraction for electrons than nonmetals.

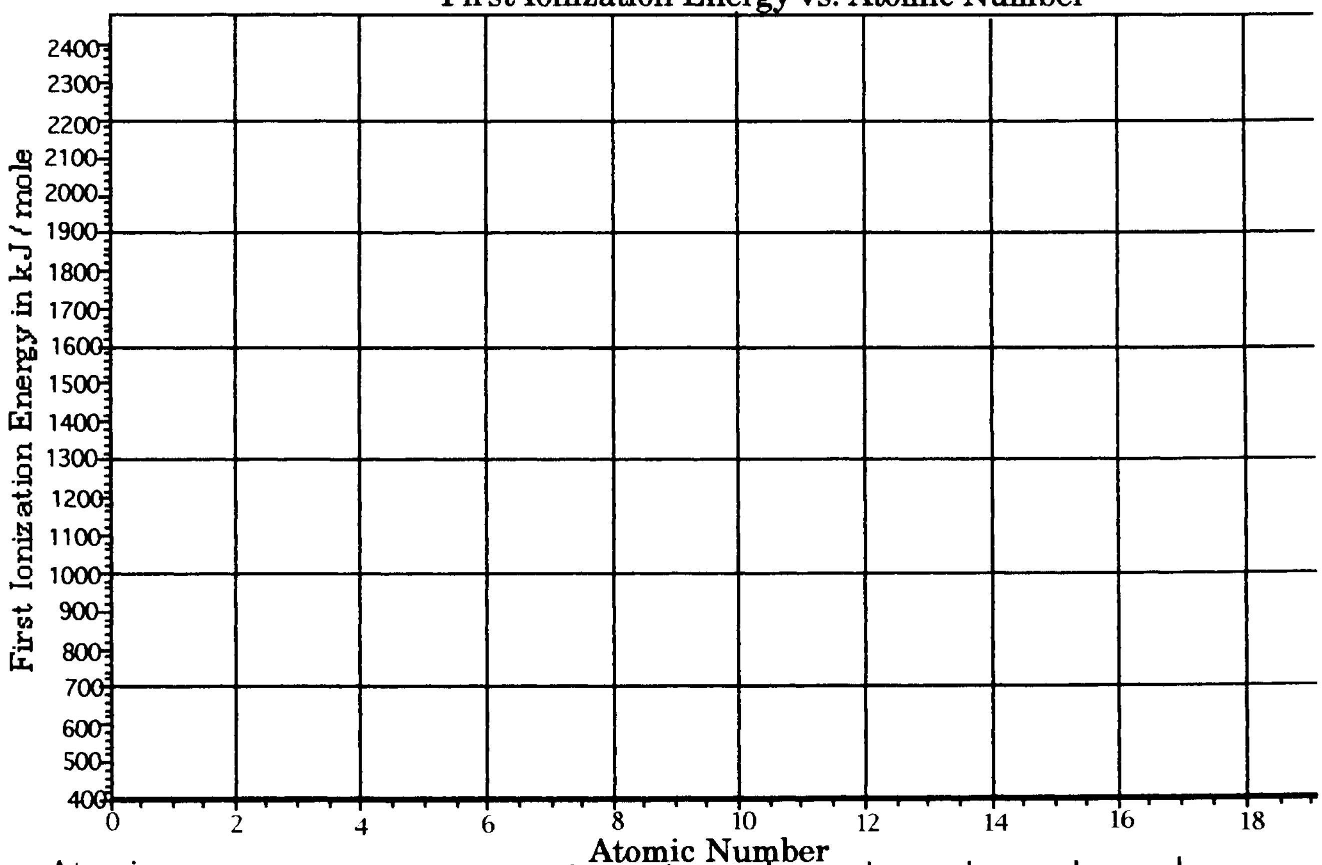
1. First **ionization energy**(IE) is the amount of energy needed to remove the most loosely held electron from a gaseous atom to produce a positive ion. An equation to represent this change is: atom(g) + ionization energy -----> $ion^{1+}(g)$ + e^-
The higher the ionization energy for an element, the greater the attraction between the nucleus and the valence electrons.

- Going down a group, the ionization energy decreases as the valence shell gets farther away from the pull of the nucleus.
- Going across a period from left to right, the ionization energy has a general

tendency to increase with increasing kernel charges. Elements in Groups 2A, 2B, 5A and 8A have unusually high ionization energies because of their filled and half-filled energy sublevels, which makes the periodic trend more varied.

Problem 1:
Plot the data on the graph below for the first ionization energy values vs. atomic numbers. Label each point with the symbol of the element and the electron configuration for <u>only the last</u> sublevel that receives electrons.

First Ionization Energy vs. Atomic Number



Atomic Number	1	2	3	4	5	6	7	8	9	10	11
first IE (kJ / mole)	1312	2371	520	900	800	1086	1402	1314	1681	2080	496

Atomic #	12	13	14	15	16	17	18	19	20
first IE	738	577	786	1012	1000	1255	1520	419	590

Problem 2: Trends in Ionization Energy
Explain the trends that you observe in terms of electron configuration.

- It is possible to remove more than one electron from an atom, but it requires increasingly more and more energy to remove each successive electron. The most dramatic rise in ionization energy occurs once the positive ion has obtained the electron structure of a noble gas. This particularly stable noble gas structure is difficult to "break".

2. **Electron Affinity**(EA) represents the energy change that occurs in a gaseous atom when it gains an extra electron. This can be an exothermic or endothermic reaction, depending upon whether the additional electron makes a negative ion that is chemically more stable or less stable than the atom.

   - The reactive nonmetals will become the most stable negative ions, so they will have the highest electron affinity values. The exothermic reaction to add an electron to a fluorine atom is written as: $F(g) + 1e^- \text{-----}> F^{1-}(g) + 328kJ$ (the electron affinity value) Since energy is being lost by the atom, the electron affinity is expressed as a negative value(EA = - 328 kJ / mole.)
   - The noble gases(and Group 2A) that have filled energy sublevels will become <u>less stable</u> if an extra electron is added. The atom must be forced to gain an electron by adding energy along with it, which makes the change an endothermic reaction. The reaction for a neon atom gaining an electron is written as: $Ne(g) + 1e^- + 29kJ \text{-----}> Ne^{1-}(g)$
     The electron affinity is expressed as a positive value, EA = +29 kJ / mole. Note that Be has an EA value that is less than Ne (EA for Be = 19 kJ / mole), possibly because the Be atom has a smaller kernel charge than a Ne atom to hold onto the extra electron.
   - Although metals become *most* stable by losing electrons, they can gain extra electrons in an electrical field to become slightly more stable. The Group 1A elements, for example, obtain an octet in their valence shell and become most stable by losing one electron. However, in an electron rich environment, they can become a little more stable by gaining an electron to fill their s sublevel in the valence shell.
   - Trends in electron affinity are similar to those for ionization energy. In general, EA increases as you move from left to right across a period and decreases as you move down a group. The elements in Groups 2A, 2B, 5A and 8A have unusually low electron affinity values, because they already have stable filled or half-filled sublevels which are disrupted by an additional electron.

Problem:
Predict which of the following elements would have the largest and smallest electron affinity values and explain the reasons for your choices: Li, Mg, S, Cl, As

ANS: largest = Cl, smallest = Mg

3. **Electronegativity** is the relative tendency of an atom to attract electrons to itself when it is chemically combined with another element. Scientists have developed several equations for calculating electronegativities based on the various factors of atomic structure or energy, and the values range from 0.7 for Fr to 4.0 for F. The noble gases do not have any calculated electronegativities because they do not form any compounds in nature.

- Metals have low electronegativities due to their small kernel charges and large atomic radii. Nonmetals have high electronegativities due to their large kernel charge, small size and increased chemical stability gained with the extra electron.
- General trends are for electronegativity values to increase going across a period

and decrease going down a group.

- In compounds, the element with the highest electronegativity will be assigned a negative oxidation number, the other element will have a positive number.
- It has been experimentally determined that if the difference in electronegativities between the two atoms is greater than 1.67, then a complete transfer of electrons occurs to form ions. A difference of less than 1.67 usually results in the atoms sharing valence electrons to form a molecule. (the differences are based on individual bonds between two atoms)

Problems:
Which pairs of elements will react to form ions and which will form molecules. Explain your choices.

1. Na and Cl 2. Ca and O 3. N and I 4. He and S 5. H and C

ANS: I, I, M, neither, M

D. Predicting Oxidation Numbers

When elements react to form compounds, the oxidation numbers reflect the number of electrons each atom tries to lose or gain to become more stable.

- Group 1A, 2A and 3A will lose 1, 2 and 3 electrons, respectively, and are assigned oxidation numbers of 1+, 2+ and 3+.
- Transition metals usually have a valence shell with 2 electrons in an s sublevel. They will lose these electrons first, giving them a typical oxidation number of 2+. In a stronger oxidizing environment, some lower energy level electrons in the d sublevel can be lost one at a time. Manganese, with a configuration of $[Ar]4s^2 3d^5$ will have oxidation numbers ranging from 2+ to 7+. *Notice* that Mn is classified as a Group 7B element, because it has a maximum of 7 electrons that can be lost in a reaction. The other B groups follow a similar pattern. In contrast to Mn, iron with a configuration of $[Ar]4s^2 3d^6$ will have oxidation numbers of 2+ or 3+ only. To make the oxidation number of iron higher than 3+ would require removing electrons from a particularly stable half-filled d sublevel, which normally will not occur.
- The heavier metals in Groups 3A to 5A follow a similar change in oxidation numbers. For example, tin has a configuration of $[Kr]5s^2 4d^{10} 5p^2$ and will lose valence electrons from the 5p sublevel to have an oxidation number of 2+ or also remove the additional 2 electrons from the 5s sublevel for a 4+ state. Electrons will not be removed from the stable, filled 4d sublevel.
- Nonmetals in Groups 5A, 6A and 7A will try to gain 3, 2 and 1 electrons, respectively, and are assigned oxidation numbers of 3-, 2- and 1-. They may also share some or all of their valence electrons with another nonmetal in an effort to make both elements more stable. The nonmetal with the lower electronegativity is assigned a positive oxidation number with a value dependent upon the number of electrons being used to share with the other element. For example, chlorine has a configuration of $[Ne]3s^2 3p^5$. In a reaction with sodium, the chlorine will gain 1 electron and have a 1- oxidation state. When combined with oxygen that has a higher electronegativity, the chlorine usually shares 1, 3, 5 or all 7 of its

valence electrons, depending on how many oxygen atoms it combines with. Chlorine's oxidation numbers would then be 1+, 3+, 5+ or 7+.

- Although we might expect the same behavior from the rare earth elements with electrons in the f sublevel, the lanthanide series has a 3+ as the most stable oxidation state, while the actinide series is usually 3+ or 4+.

Problems:
Looking only at the position of the following elements on the periodic table, write what you predict the oxidation states could be for each:

1. W 2. Al 3. Xe 4. Ba 5. At 6. H 7. S 8. Zn

VI. A Thumbnail Sketch of the Representative Elements

By learning the characteristic properties of a few representative elements, and knowing how changes in atomic structure effect the properties, you can predict the chemical and physical behavior of all 109 known elements.

A. **Hydrogen** - Although hydrogen is positioned in Group 1A because it has 1 valence electron like the alkali metals, it has its own unique properties and should be considered as a family by itself. Hydrogen is the most abundant element in the universe, and it is believed that all of the other elements were created by stellar nuclear fusion of hydrogen into helium and other heavier elements. Hydrogen is an explosive gas that reacts with oxygen in the air to form water. The hydrogen shares its electron with the oxygen. In acids, hydrogen loses its electron to produce an H^{1+} ion, which is simply a proton. When combined with reactive metals(Group 1A or 2A), the hydrogen will gain an electron to form a hydride ion, H^{1-}. Life would not be the same without hydrogen, because it is able to form bridges between atoms and holds the strands of DNA together.

B. **Alkali and Alkaline earth metals** - These are elements in Groups 1A and 2A, respectively, and are not found uncombined in nature because they are very reactive. Freshly cut metal is shiny, but will become dull quickly as it combines with oxygen in the air. Both groups will react with water by a single replacement reaction to form a basic solution of metal hydroxide and hydrogen gas. The reaction with the heavier Group 1A metals occurs so quickly that the heat of reaction causes the liberated hydrogen gas to explode. The alkaline metals are denser, harder, less reactive, and have higher melting and boiling points than the corresponding alkali metals in the same period.

C. **Groups 3A to 5A** - All three of these groups have nonmetals at the top of the group and metals at the bottom, so properties vary accordingly. Usually the focus is placed on an element or two in each group that has a significant importance or application. In Group 3A, aluminum is studied as the third most abundant element in the Earth's crust as a major component of many rocks and minerals. It is used to make lightweight alloys(mixture of metals) because it is corrosion-resistant. Carbon is the element of life because of its ability to share its 4 valence electrons in many different ways. All living organisms are built by chains of carbon atoms linked together to make proteins, cells structures, etc. Carbon chemistry is the basis of an entire branch of study called organic chemistry. Nitrogen and phosphorus are also essential elements for living organisms. Adenosine triphosphate(ATP) is involved in the utilization of energy by living systems. DNA and RNA contain both nitrogen and phosphorus as key parts of their structure.

D. **Group 6A and the Halogens** - These are the most reactive nonmetals, with fluorine being most active of all. They will cause other elements to be oxidized by losing their electrons to them. Oxygen is the most abundant element in the Earth's crust since it reacts with every element except He, Ne and Ar. In the atmosphere, it also exists in a triatomic form, O_3, which is called ozone. Acid rain is produced when water in the atmosphere combines with oxides of sulfur, carbon and nitrogen. The name "halogen" given to the Group 7A elements means "salt forming". All of these elements react vigorously with metals to form ionic compounds(salts). Many of these elements(HINClBrOF) exist as diatomic molecules because they are so reactive that they will share electrons with themselves in an attempt to become more stable.

E. **Group 8A, the Noble Gases** - These gases are chemically inert and will not form compounds in nature. Helium was first detected in the sun in 1868 when a new line spectrum was observed. In 1962, Neil Bartlett synthesized the first noble gas compounds, $XePtF_6$, but this was under high pressure and temperature. Noble gases are used to create a nonreactive atmosphere for welding and to fill light bulbs to prevent the filaments from burning up.

F. **Transition metals** - Typically, these elements have several oxidation states because they can use both their s and d sublevel electrons to form chemical bonds. The very small, highly charged positive ions that form attract negatively charged ions, including those in large organic molecules. They define the shape of such molecules as hemoglobin(with Fe). Many of the compounds with transition elements are colored, so they are used as pigments in paint and dyes. Also, these metals are very dense and hard, therefore they are frequently used as structural building materials.

G. **Rare earth(also known as inner transition) metals** - The lanthanide series contains elements that are not common, everyday names, but they are part of our daily lives. For the most part, these metals are used to make alloys and ceramic materials that have better conductivity and magnetic properties. They are used for improving the brightness of your color TV, making your stereo sound better with smaller, more powerful magnets in the speakers, and increasing the information on the Internet through fiber optic cables. All of the elements in the actinide series beyond uranium have been synthetically made by nuclear reactions. All of these elements are radioactive, because they have too many protons and neutrons. Uranium(and plutonium) are the typical fuels used to produce energy by nuclear fission. Americium can be found in most homes in America in your smoke detectors.

Problem:
Look at the labels of some of the consumer products you have in your home(or conduct a literature search) to compile a list of how elements from each of these groups have an impact on your life.

Unit 7 Objectives

Having studied the unit notes and done the problems, you should be able to:

1. List early attempts at classification of the elements and the limitations of these attempts, including Newland's law of octaves and Mendeleev's periodic table.

2. Describe the organization of the modern periodic table, and what is the significance of the vertical columns and horizontal rows..

3. Define the 4 quantum numbers and how they relate to the electrons, orbitals, energy levels and sublevels.

4. Write the electron configuration for an atom or ion using the periodic table or the diagonal rule.

5. Be able to draw an orbital notation or dot diagram for any atom or ion.

6. Write and explain the electron configurations, orbital notations or dot diagrams of the elements based on their position on the periodic table.

7. Distinguish between groups, families, series and periods on the periodic table.

8. Describe the properties of metals, nonmetals and metalloids.

9. Use the periodic table to locate each of these groups of elements: alkali metals, alkaline earth metals, halogens, noble gases, transition elements, lanthanides, and actinides.

10. Determine whether an atom is unstable or stable using the octet rule, or filled versus half-filled sublevels, and recognize the stability of the noble gases.

11. Explain why there are exceptions to the predicted sequence for filling electron energy sublevels based on the diagonal rule.

12. Describe the 3 factors of atomic structure that influence the trends of properties for elements in a particular group or period.

13 Relate noble gas configuration and filled or half-filled sublevels to the stability of an atom or ion.

14. Predict and explain the oxidation number of an element given the element's location in the periodic table.

15. Compare the relative first ionization energies of two elements given their positions in the periodic table.

16. Explain the variations in ionization energies of elements as their atomic numbers increase in a period or family of the periodic table.

17. Explain the increase in subsequent ionization energies for an element, from first to second, to third, and so on.

18. Define electron affinity and predict trends in electron affinities for elements, based on their location in the periodic table.

19. Use electronegativity values to determine which element has a positive and negative oxidation state in a compound, and will ions or molecules be formed.

20. Be able to summarize the key characteristics of the representative elements.