Periodic Table and Bonding

I. Handout: Periodic Table and Bonding Notes

II. Periodic Properties and the Development of the Periodic Table

  I. The first periodic table was arranged by Dimitri Mendeleev in 1869.

    I. He was a professor of Chemistry. at the University of St. Petersburg in Russia and was confronted with the problem of how to teach about the various elements known at that time. He decided to organized the elements by arranging them into groups that reacted similarly.

    II. He also noticed that various properties would repeat "periodically" so he arranged a table of elements order of atomic mass such that properties would change regularly if you moved across a row while maintaining groups with similar chemical properties in a column.

    III. Go to: http://www.periodic.lanl.gov/mendeleev.htm to see a version of Mendeleev's first table.

  II. Groups with similar properties

    I. All the elements in a group (or column) are called families.

    II. Group 8: The Noble Gases, don't react with other elements.

    III. Group 1: The Alkali Earth Metals, all react with water in the following manner

$2\ Li + H_2O \longrightarrow H_2 + 2\ LiOH$

$2\ Na + H_2O \longrightarrow H_2 + 2\ NaOH$

...

$2\ Fr + H_2O \longrightarrow H_2 + 2\ FrOH$

    IV. These are just a few examples of how Mendeleev organized the columns or families.

  III. Periodic Properties

    I. As you move across a row various properties change regularly click on the images below to see a visualization of the various properties. All of these images are from www.webelements.com, one of the best periodic table sites on the web.

  IV. Early on the elements were divided into two broad categories -> metals and non-metals. This was done long before anyone knew any detail about the atoms or any of the periodic properties mentioned above.

V. As you can see what makes something a metal or a non-metal is based on other properties like ionization energy, atomic radius, and electronegativity

VI. When metal atoms are bonded together the electrons become delocalized, jumping from one atom to another. A common analogy is to say that the nuclei of atoms in a metal exist in a "sea of mobile electrons".

VII. This is due to the low ionization energy of these electrons, and is what gives metals the property of conductivity. A typical electric current can be described as electrons moving from one place to another. This can easily happen in metallic substances as depicted below.

VIII. Mendeleev organized his table based on properties of density, melting point, and oxide formula. At first people rejected his organization of the elements. However, he found that as he organized the table there seemed to be "holes". He predicted that new elements would be found and he predicted the properties of those new elements. When these new elements were found Mendeleev's periodic table was acclaimed as correct and became and indispensable tool for understanding Chemistry.

I. Lab: Organizing the "elements"/moon phases.

II. Film: Making Glass using the Periodic Table

III. Handout: Periodic Table

IV. Handout: Periodic Trends

V. **Homework:** Periodic Properties Questions

III. Valence Electrons

I. Electron Configurations

I. Although the original periodic table was arranged by properties of the elements, Mendeleev didn't realize that it was the underlying structure of the atoms that gave elements those properties. Today's table is based strictly on the underlying structure. It looks very similar to the early tables, but not exactly.

II. So, to better understand the periodic table the properties of substances we will need to explore the structure of the atoms.

I. An atom is made of a nucleus of protons and neutrons, and an outer region containing electrons in orbitals (s,p,d, or f type).

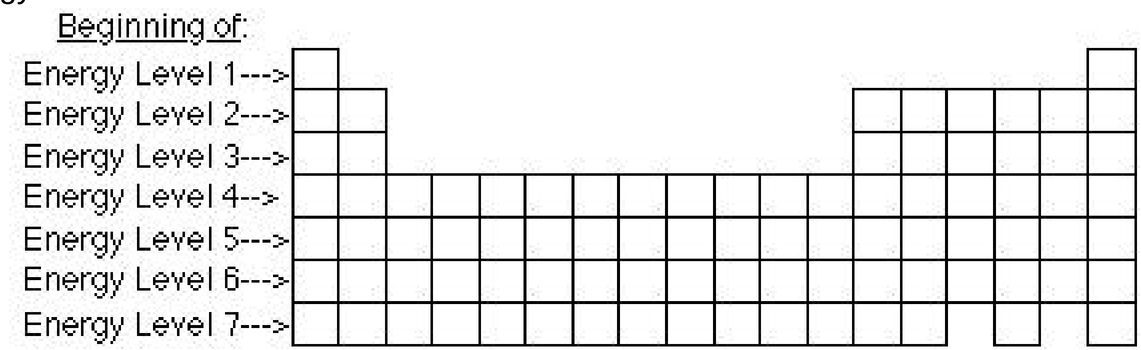
II. Each electron has a specific amount of energy associated with the orbital in which it is found.

III. Film: Orbitals

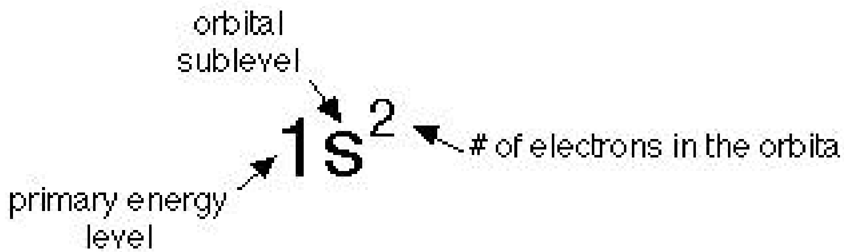
IV. The outer region of the nucleus is 10,000 times the size of the nucleus, so the nucleus is buried deep inside the atom.
V. Because the nucleus is tucked away beneath the electrons, it is the electrons that give an atom its properties.
VI. Specifically, the outermost or valence electrons will primarily determine how atoms interact with each other.

III. Energy Levels and Electron Filling Order

I. There are primary energy levels, and sublevels within each primary level.
II. Each row or period in the periodic table is considered to be the start of a primary energy level.



III. Each different type of orbital in a primary energy level is a sublevel.
IV. Each orbital can only hold two electrons and they must have opposite spin. This is called the Pauli exclusion principle.
V. Electrons will fill up lowest energy orbitals first.
VI. The lower energy sublevels of one primary energy level can overlap with the upper energy sublevels of another primary energy level. This can result in orbitals of a higher principle energy level filling before the orbitals in a lower principle energy level.
VII. To completely describe the electron configuration for an atom you need to specify how many electrons are in each orbital at each level. This is done with a specific kind of notation.



VIII. Electron Configuration Examples (click on links to see a graphical representation from www.webelements.com)

I. H = $1s^1$

II. He = $1s^2$

III. Li = $1s^2 2s^1$

IV. O = $1s^2 2s^2 2p^4$

V. Click here to see an applet which will display the electron configuration of any element.

VI. You can also use a shortcut in writing electron configurations by putting the previous closest nobel gas in brackets indicating that you start with the electron configuration for that element and add to it. For example,

Br = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$ = $[Ar] 4s^2 3d^{10} 4p^5$.

II. Intro to Lewis Dot Notation

I. Valence electrons are those electrons that are in the highest principle energy level. It is these electrons that primarily interact with other atoms.

I. Oxygen has 6 valence electrons: $1s^2 \mathbf{2s^2 2p^4}$

II. Sodium has 1 valence electron: $1s^2 2s^2 2p^6 \mathbf{3s^1}$

III. Bromine has 7 valence electrons: $1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{4s^2} 3d^{10} \mathbf{4p^5}$

IV. Xenon has 8 valence electrons: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 \mathbf{5s^2} 4d^{10} \mathbf{5p^6}$

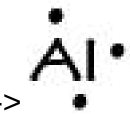
V. Pick two elements from any column and determine how many valence electrons those atoms have. You should notice a particular pattern.

II. Bonds form when valence electrons are shared or transferred from one atom to another, so our focus will mainly be on these highest energy electrons.

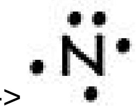
III.Lewis Dot diagrams are a graphical way of showing how many valence electrons an atom has. Below are some examples of electron configurations and the Lewis Dot diagram for each of those atoms.

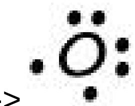
I.Sodium has 1 valence electron: $1s^22s^22p^6\mathbf{3s^1}$ with the Lewis Dot Diagram -> 

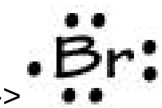
II.Beryllium has 2 valence electrons: $1s^2\mathbf{2s^2}$ --> 

III.Aluminum has 3 valence electrons: $1s^22s^22p^6\mathbf{3s^23p^1}$ --> 

IV.Germanium has 4 valence electrons: $1s^22s^22p^63s^23p^6\mathbf{4s^2}3d^{10}\mathbf{4p^2}$ --> 

V.Nitrogen has 5 valence electrons: $1s^22s^22p^6\mathbf{3s^23p^3}$ --> 

VI.Oxygen has 6 valence electrons: $1s^2\mathbf{2s^22p^2}$ --> 

VII.Bromine has 7 valence electrons: $1s^22s^22p^63s^23p^6\mathbf{4s^2}3d^{10}\mathbf{4p^5}$ --> 

VIII.Xenon has 8 valence electrons: $1s^22s^22p^63s^23p^64s^23d^{10}4p^6\mathbf{5s^2}4d^{10}\mathbf{5p^6}$ --> 

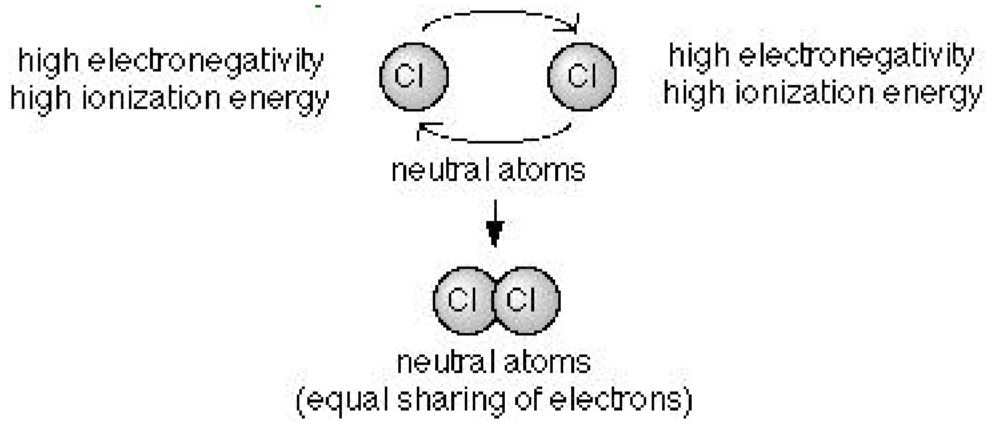
I.Homework: For each of these elements Li, C, Mg, Cl, Kr, and Ag write the following:

I.electron configuration - full version
II.electron configuration - shortcut version
III.Lewis dot diagram

- IV. Types of bonds
  - I. Strong Bonds
    - I. Ionic and Covalent Bonds Overview
      - I. Electronegativity and Bonding
        - I. Strong bonds form between atoms when they share or transfer electrons.
        - II. Depending on how even or uneven the sharing is between the atoms several deferent kinds of strong bonds can form.
        - III. The way to determine if the atoms will share their electrons evenly or unevenly is to examine the electronegativity of each atom.
        - IV. Electronegativity is how strongly an atom attracts electrons to itself when bonded with another atom.
        - V. The illustration below shows that atoms in the upper right corner of the periodic table tend to attract electrons very strongly when bonded, while the atoms in the lower left corner don't attract electrons to themselves very well. (Except under unusual conditions, the noble gasses don't usually form bonds, so electronegativity has no meaning for atoms which are not bonded to other atoms.)

          or if you prefer sphere sizes to represent the electronegativity:

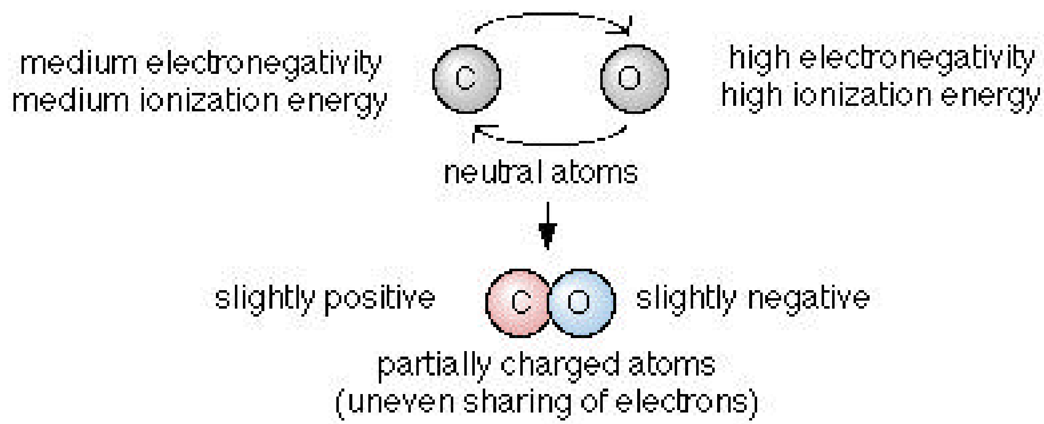
        - VI. When two atoms are bonded together there are three basic ways to pair them up:
          - I. Two atoms with the same electronegativity, either both high or both low.
            - I. This will cause the electrons that are shared in the bond to be evenly shared between the atoms.
            - II. When atoms share electrons evenly between each other the bond formed is called a non-polar covalent bond.



II.One atom with a somewhat higher electronegativity than the other.

I.This will cause the electrons to be shared unevenly, such that the shared electrons will spend more time on average closer to the atom that has the higher electronegativity.

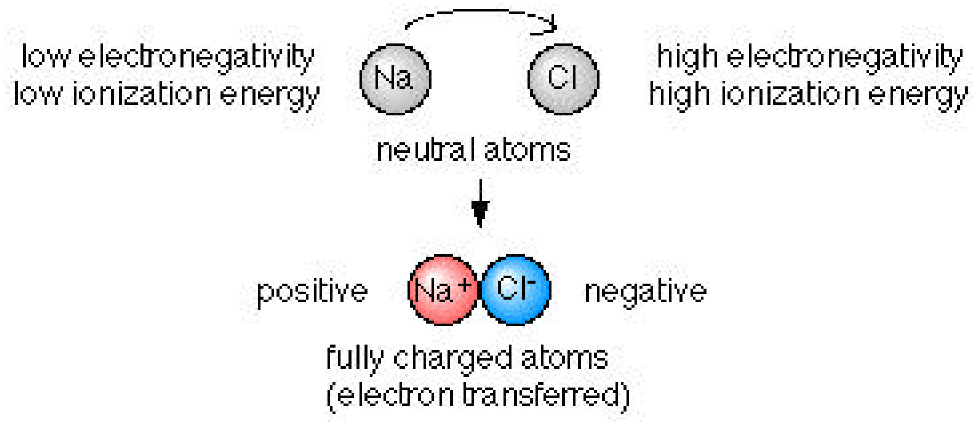
II.When atoms share electrons unevenly but not very unevenly the bond formed is called a polar covalent bond.



III. One atom's electronegativity is much higher than the other atom.

I. In extreme cases the electrons in the bond spend so much time closer to the atom with high electronegativity that the shared electrons are considered to be transferred to that atom. The "sharing" is so uneven that one atom basically "takes" one or more electrons from the other atom.

II. When the electrons being "shared" are so unevenly distributed between the atoms the bond that is formed is called an ionic bond.



II. Covalent Bonds

I. Non-polar covalent bonds

I. If the difference in the electronegativity between the two bonded atoms is less than 0.5 then the bond formed is considered to be non-polar covalent.

II. Each atom attracts the other atom's electrons about equally so that the electrons spend equal amounts of time near each atom.

III. Overall, both atoms will be neutral, having the same charge.

II. Polar covalent bonds

I. If the difference in the electronegativity between the two bonded atoms is between 0.5 and 2.1, then the bond formed is considered to be polar covalent.

II. One atom attracts the other atom's electrons better, so the electrons stay closer (on average) to that atom. This causes an imbalance of electric charge within the bond between the two atoms.

III. The atom that pulls the negative electrons better toward itself will be slightly negative and the other atom will be slightly positive.

III. Ionic Bonds

I. If the difference in the electronegativity between the two bonded atoms is greater than 2.1, then the bond is considered to be ionic.

II. Because one atom pulls the other atom's electrons so well toward itself, there is a great imbalance of electric charge. If for some reason the bond between the atoms is broken, the atom with the higher electronegativity will actually keep the electron for itself.

III. In this case the atoms with the higher electronegativity will be fully negative (due to the "gaining" of an electron) while the other atom is fully positive (due to its virtual loss of an electron).

IV. Summary of Electronegativity and Bond formation

I. Only the extreme cases are very clear. Very small differences in electronegativity result in non-polar covalent bonds, and very large differences in electronegativity result in ionic bonds. All other bonds are somewhere in-between.

Type of Bond	Difference in Electronegativity
Non-Polar Covalent	less than 0.5
Polar Covalent	between 0.5 and 2.1
Ionic	greater than 2.1

II. What kind of bond will form between the following atom pairs:

H and H

H and F

H and C

Li and F

C and O

I. Handout: Electronegativity Tables

II. Handout: Make your own bonds guide sheet

III. **Homework:** By using the periodic trends handout and the electronegativity tables determine what kind of bond will form between the following pairs of atoms: Na and Cl, C and O, Ca and F, N and N. Indicate if each pair will form an ionic bond or a covalent bond and describe the difference between these two bonds.

II. Ionic Substances

I. Ions form when the charge imbalance between bonded atoms is so large that one or more electrons are basically, transferred from one atom to another.

II. When this happens ions are formed (both positively charge and negatively charged ions).

III. If you put a bunch of positively charged and negatively charged ions in one place the opposite charges tend to attract strongly to each other forming clusters of ions containing equal amounts of positive charge and negative charge, resulting in a neutral substance.

IV. The cluster of ions formed can be of any size as long as there is an equal amount of positive and negative charge. For example, a tiny grain of table salt (NaCl), contains trillions, and trillions of sodium and chlorine ions.

V. We don't call these clusters of ions molecules. Instead they are referred to as crystals. (Any well organized group of ions or even molecules can be referred to as a crystal).

VI. Below are some examples of ionic substances:

Aluminum Oxide Al_2O_3	Calcium Fluoride CaF_2
Lead(II) Sulfide PbS	Sodium Chloride NaCl

VII. Notice above that the formula for an ionic substance specifies the RATIO of one element to another, not the number of atoms. This ratio depends on the charge of each of the ions.

VIII. The charge of an ion depends on how many electrons it loses or gains, and the number of electrons lost or gained depends on the number of valence electrons the atoms have.

IX. Through experimentation chemists discovered a pattern in which elements lose or gain electrons and how many they lose or gain.

I. Elements in the first column tend to lose one electron forming +1 charged ions.

II. Elements in the second column tend to lose two electrons forming +2 charged ions.

III. Elements in the halogen family (second to last column) tend to gain one electron forming -1 charged ions.

IV. Elements in the Noble Gas family (last column) don't form any charges because they tend not to lose or gain any electrons.

X. The fact that the Noble gasses don't lose or gain electrons indicates that they have the most stable electron configuration with 8 valence electrons.

XI. Atoms can lose or gain electrons to achieve these electron configurations.

I. Metals tend to loose electrons (forming positively charged ions) because they have relatively low ionization energy.

I. Notice the first column has all atoms with one valence electron.

Na = $1s^2 2s^2 2p^6 \mathbf{3s^1}$

K = $1s^2 2s^2 2p^6 3s^2 3p^6 \mathbf{4s^1}$

II. If these atoms loose their one loosely held valence electron, then they will have the same electron configurations as nearby Noble Gas elements and will now have 8 valence electrons.

Na+1 = $1s^2 \mathbf{2s^2 2p^6}$ = Ne

K+1 = $1s^2 2s^2 2p^6 \mathbf{3s^2 3p^6}$ = Ar

II. Non-metals tend to gain electrons (forming negatively charged ions) because they have relatively high ionization energy and high electronegativity.

I. Notice the third to last column has all atoms with 6 valence electrons.

O = $1s^2 \mathbf{2s^2 2p^4}$

S = $1s^2 2s^2 2p^6 \mathbf{3s^2 3p^4}$

II. If these atoms gain two more electrons, then they will have the same electron configurations as nearby Noble Gas elements and will now have 8 valence electrons.

O-2 = $1s^2 \mathbf{2s^2 2p^6}$ = Ne

S-2 = $1s^2 2s^2 2p^6 \mathbf{3s^2 3p^6}$ = Ar

III. So, as you can see, elements in the same column (or family) will form the same charge if they lose or gain electrons. By knowing any one element from a column you know the charge that will form on any of the others.

XII. By losing and gaining the appropriate amount of electrons ions are formed which then combine with each other in certain proportions to create a neutrally charged compound. For example:

- I. Sodium forms a +1 charged ion (Na+1) and Sulfur forms a -2 charged ion (S-2), so you would need two atoms of sodium to balance out each atom of sulfur. The formula for this substance would be written as Na2S.
- II. What would be the ratio of Calcium to Chlorine ions if they were to bond together?
  - I. Let's break this down into several steps:
    - I. How many valence electrons does each type of atom have when neutral?

      Ca

      Cl
    - II. Will they tend to lose or gain electrons?

      Ca

      Cl
    - III. What charge will each of the ions form?

      Ca

      Cl

IV.What will be the formula for the compound formed by calcium and chlorine ions:

I.Film: Bonding Basics

II.Homework:

Atoms of elements in the same column will form the same charge. Using examples (and showing the electron configurations of several elements) explain why elements in column 1 tend to form +1 charged ions while elements in column 16 tend to form -2 charged ions.

What would be the formula for the ionic compound formed between the following elements:

Magnesium and oxygen
Lithium and sulfur
Calcium and fluorine
Aluminum and Sulfur

III.Molecular Substances

I.When two or more atoms are bonded together with covalent bonds a molecule is formed.

II.Molecules are primarily formed from the nonmetal elements, because these elements have high ionization energy and high electronegativity, causing them to share electrons to forma covalent bond instead of giving them up to other atoms which would form an ionic bond.

III.When electrons are shared between specific atoms, covalent bonds link atoms together to form a molecule. See the molecule of Vitamin C below:

Vitamin C

IV.Grab the molecule above and drag it around.

V.Notice that some atoms above are bonded to one, two, three, or four other atoms and that some are double
bonded and some are single bonded.

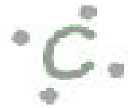
VI.To understand why we need to look at the valence electrons once again.

I.The red atoms above are oxygen atoms with 6 valence electrons.

O = $1s^2 2s^2 2p^4$ --> 

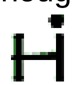
II. In order for the oxygen to have the stable 8 valence electron structure it must share two electrons with other atoms. In other words it must form two covalent bonds. Look at the red atoms above. You should notice that they always form two covalent bonds (either two single covalent bonds or one double covalent bonds).

III. The gray atoms above are carbon atoms with 4 valence electrons.

C = $1s^2 2s^2 2p^2$ --> 

IV. In order for carbon to have the stable 8 valence electron structure it must share four electrons with other atoms. Look at the gray atoms above to see if they always form four bonds.

V. The white atoms are hydrogen atoms with one valence electron. Hydrogen is closest to the Nobel Gas, Helium which only has 2 valence electrons. Therefore, hydrogen will be stable if it can share enough electrons to get two.

H = $1s^1$ --> 

VI. Notice that all the hydrogen atoms only form one bond above.

VII. Notice that the Lewis Dot diagrams for each of the atoms shows you how many bonds will form. Each unpaired electron will form a bond with another atom.

VIII. Using a piece of software called "eChem", you can experiment with how various nonmetal atoms will bond
covalently to form molecules. You can download "eChem" here:
http://www.investigationstation.org/sciencelaboratory/echem/download.html

IX. Molecules tend to fall into 4 broad categories:

- I. Small molecules
    - I. These molecules consist of a small number of atoms strongly bonded together.
    - II. Most room temperature liquids, and gasses consist of small molecules.
    - III. Some examples include: water, ammonia, butane, gasoline, air (nitrogen and oxygen)
- II. Large molecules
    - I. These molecules consists of a large number of atoms strongly bonded together.
    - II. Many biologically important substances consist of large molecules.
    - III. Some examples include: vitamins, hormones, various cellular signaling molecules
- III. Polymers
    - I. These molecules consist of repeating small molecules bonded together to form larger molecules.

II.Polymers are also large molecules, but they can be much larger than some of the large molecules listed above.

III.Some examples include: plastic, wood, DNA, proteins, enzymes (a type of protein with a special function).

IV.Network molecules

I.All previous examples involved molecules that were somewhat linear or sequential, with one atom bonded to the next and so on. Sometimes a network of bonds can form between many neighboring atoms.

II.Network molecules tend to have great relative strength because of the many covalent bonds connecting neighboring atoms. Some newly created molecules of this type are promising to revolutionize everything from drug delivery, to computer processing power.

III.Some examples include: diamond, buckyballs, and carbon nanotubes.

To see examples of the above molecule types, go to the Melange of Molecules web pages. Just close the window of molecules when you are done.

I.Film: Bonding and Electrons.

II.Handout: eChem guide sheet

III.Handout: Using the Chime Plug-in

IV.Homework: When forming covalent bonds, atoms of elements in the same column tend to form the same number of covalent bonds. Using examples (and showing the electron configurations of several elements) explain why elements in column 14 tend to form 4 covalent bonds and elements in column 15 tend to form 3 covalent bonds.

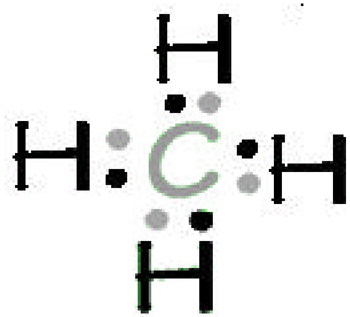
V.Lewis Dot Structures - Understanding Molecular Structure

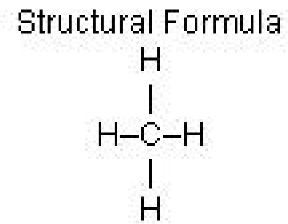
I.To help determine how atoms will covalently bond together into molecules we can use Lewis Dot Diagrams.

II.Lewis Dot Diagrams show only the valence electrons and utilize the fact that most molecular compounds have non-metal atoms sharing these electrons to have a stable 8 like the Noble Gases.

III.Below are some common Lewis Dot Symbols for some atoms:

IV.Each unpaired dot represents a valence electron that can be shared - a place where a bond can form. Methane with the formula CH_4 is diagrammed as:





I.Handout: Electron Dot Introduction

II.Homework: Electron Dot Practice Sheet 1 and Electron Dot Practice Sheet 2

VI.Shapes of Molecules

I.You may have noticed that sometimes there are multiple ways to construct a molecule from the atoms given in the formula. Take C_4H_{10} for example :

II.Each of the compounds displayed above is an isomer of butane. Isomers refer to different molecules with the same formula.

III.Click here to check out the Isomer Construction Set by Fred Senese at Frostburg State.

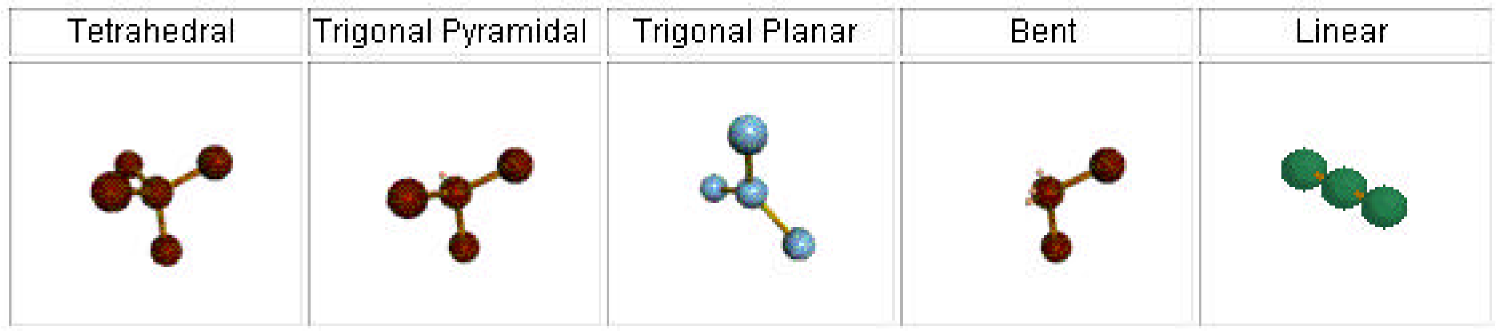
IV.Did you notice that the molecules shown above have a particular shape to them?

V.Molecules will form into shapes such that regions of high electron density (where electrons are being shared between atoms and where there are unshared pairs of valence electrons on the surface).

VI.Because all of these regions are negatively charged, they repel each other and try to move as far away as

possible from each other.

VII.Depending on how many atoms are bonded and if there are unshared pairs of electrons around, you will see the following common shapes:



IX.Check out the images of molecules below and see where you can find these shapes within them. The images are being displayed by the Chime plug-in which allows you to manipulate and take measurements on the displayed molecules. See the Using the Chime Plug-in sheet for instructions.

X.Click here to bring up a page that contains several molecules for you to explore using the Exploring Molecular Shapes guide.

XI.The shape of molecules is extremely important, especially for larger more complex molecules. Below are several examples showing how shape is the key factor in a molecules biological function.

I.Immune function: Our immune system has the capability to recognize foreign material that enters our body. It does this not by "thinking" about it. After the first time our immune system encounters a pathogen, it makes antibodies that are just the right shape to bond with the foreign antigen. Below is an example showing the antibody molecule in green and the antigen in red:

XII.The synthesis of ATP (the primary energy storing molecule in our body) happens in a series of steps using ATP synthase.

I.Film: See the entire ATP Synthase enzyme at work.

II.Film: Focus on the active region.

XIII.Click here to see how some molecules can change shape in order to better fit with another molecule.

I.Homework: Do #1 from the Exploring Molecular Shapes Guide.

II.Handout: Using the Chime Plug-in

II.Weak Bonds (van der Waals attractions)

I.Weak intermolecular forces (Van der Waals forces)

I.Dipole-dipole attraction

I.Some molecules form areas of positive and negative charge formed through an uneven sharing of electrons (polar covalent bonding). Water is formed with polar covalent bonds between hydrogen and oxygen. Below is water.

II. Because part of the molecule is partially positive (not as positive as an ion with a +1 charge) there are attractions between the negative portion of one molecule and the positive portion of another molecule. This attraction forms weak bonds between molecules

III. When hydrogen is one of the atoms within a molecule that is attracted to the dipole on another molecule, this somewhat stronger dipole-dipole attraction is called a hydrogen bond. **The hydrogen bond is the attraction between molecules, not the covalent bond which is formed between hydrogen and an atom from its own molecule.** Below are some examples of hydrogen bonding.

  I. The dotted line below show the attraction/hydrogen bond between two water molecules.

  (This video clip, used with permission, was developed at the NYU Scientific Visualization Center.)

  II. To see a 3D view of water and it's hydrogen bonds in motion go to: http://polymer.bu.edu/vmdl/Installers/water/install.htm and follow the instructions for installing the software. (Windows only.)

  III. Hydrogen bonding is also an important factor in helping to shape the structure of larger molecules. DNA is an excellent example. Click here to see how these bonds hold together our double helix.

IV. **A molecule can have more than one polar region, so the more polar regions a molecule has, the greater two molecules of this kind will attract to each other.**

II. London Dispersion forces

I. Even when atoms are sharing electrons equally, the electrons are not static objects. They are constantly in motion. Sometimes due to their random movement between the two atoms in covalent bond they just happen to be more on one side than another.

II. A fleeting instantaneous dipole (region of positive and negative charge) can be formed by the random distribution of electrons at any particular moment.

III. This instantaneous dipole can induce a dipole in another nearby non-polar molecule. They can then attract to each other in a similar way as the dipole-dipole attraction. However, the London dispersion force is much weaker than a dipole-dipole attraction.

IV. **The size of a molecule can affect the London dispersion force between two molecules. The more surface area there is on a molecule the greater chance there will be at least one instantaneous dipole at any particular moment. Therefore, the greater the surface area (generally this means the bigger the molecule) the stronger the attraction between two molecules of this type due to London dispersion forces.**

I. Demo: Viscosity

II. See the molecules in the viscosity demo.

III. Some properties that are affected by van der Walls forces

I. Melting point

II. Boiling point

III. Evaporation rate

IV. **Homework:** 1) Pick one of the properties above and explain how intermolecular forces play a role in creating these characteristics of materials. Choose a substance and speculate on which kind of Van der Waals attractions may be involved and describe how these weak bonds are formed. 2) Explain the differences between covalent bonds, ionic bonds, and van der Waals bonds.

III. Experiment with "feeling" the difference between the types of bonds.

V. Handout: Periodic Table and Bonding Review Sheet