Chemical Bonds and Energy

I. Handout: Chemical Bonds Notes

II. Types of Bonds

i. Strong Bonds - Covalent and Ionic

I. Electronegativity and Bonding

i. Strong bonds form between atoms when they share 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.)

Table of Pauling Electronegativity

H 2.1																	He 0
Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne 0
Na 0.93	Mg 1.31											Al 1.61	Si 1.9	P 2.19	S 2.58	Cl 3.16	Ar 0
K 0.82	Ca 1.0	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.9	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.0
Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.2	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.6
Cs 0.79	Ba 0.89	La 1.1	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.2	Pt 2.28	Au 2.54	Hg 2.0	Tl 2.04	Pb 2.33	Bi 2.02	Po 2.0	At 2.2	Rn 0
Fr 0.7	Ra 0.89	Ac 1.1															

vi. When two atoms are bonded together there are three basic ways to pair them up:

a. **Two atoms with the same electronegativity**, either both high or both low.

1. This will cause the electrons that are shared in the bond to be evenly shared between the atoms.
2. When atoms share electrons evenly between each other the bond formed is called **a non-polar covalent bond**.

b. **One atom with a somewhat higher electronegativity than the other.**

1. 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.
2. When atoms share electrons unevenly but not very unevenly the bond formed is called **a polar covalent bond**.

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

1. 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.
2. 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

a. If the difference in the electronegativity between the two bonded atoms is less than 0.4 then the bond formed is considered to be non-polar covalent.

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

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

ii. Polar covalent bonds

a. If the difference in the electronegativity between the two bonded atoms is between 0.4 and 2.1, then the bond formed is considered to be polar covalent.

b. 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.

c. 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 0.4
Polar Covalent	between 0.4 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

a. Handout: Electronegativty Tables

b. 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.

III. 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

IV. Molecules tend to fall into 4 broad categories:

i. Small molecules

a. These molecules consist of a small number of atoms strongly bonded together.

b. Most room temperature liquids, and gasses consist of small molecules.

c. Some examples include: water, ammonia, butane, gasoline, air (nitrogen and oxygen)

ii. Large molecules

a. These molecules consists of a large number of atoms strongly bonded together.

b. Many biologically important substances consist of large molecules.

c. Some examples include: vitamins, hormones, various cellular signaling molecules

iii. Polymers

a. These molecules consist of repeating small molecules bonded together to form larger molecules.

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

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

iv. Network molecules

a. 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.

b. 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.

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

1. Handout: eChem guide sheet

c. 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 attracts 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).

d. Homework: Types of Substances Sheet

ii. Weak Bonds (van der Waals attractions)

I. Weak intermolecular forces (Van der Waals forces)

i. Dipole-dipole attraction

a. 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.

b. 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

c. 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.**

2. 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 sofwtare. (Windows only.)

3. 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.

d. **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

a. 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.

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

c. 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.

d. **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.**

a. Demo: Viscosity

b. See the molecules in the viscosity demo.

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

1. Melting point
2. Boiling point
3. Evaporation rate

d. **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. Chemical Formulas and Equations

i. Elements

a. The names of the elements are given on the periodic table.

b. Formulas are written differently depending on the element.

1. The formula for most elements is just its symbol. For example, Na for sodium or Xe for xenon.
2. Some elements naturally come in diatomic molecules. When expressing this element in its pure form we would write a formula indicating this state. There are seven diatomic elements: H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2. You should memorize these.

ii. Compounds

a. Ionic Compounds

I. Ionic compounds are formed between oppositely charged ions usually consisting of a metal and one or more non-metals.

II. An ion can be a single charged atom or a small group of atoms (molecule) with a charge.

III.Binary Ionic Compounds (compounds composed of two single atom ions)

i.Naming

a.We can form ionic compounds from choosing a metal and a nonmetal. This is best taught by example

b.Sodium and chlorine form Sodium Chloride.

c.Magnesium and oxygen form Magnesium Oxide.

d.Calcium and sulfur form Calcium Sulfide.

e.Binary ionic compounds are named by removing the end of the name from the nonmetal and adding -ide.

ii.Formula writing

a.To write the correct formula you must know the charges present on each ion. To determine this you would look on the periodic table or your common ion sheet.

b.The positive and negative charges must exactly balance each other in order to have the correct ratio of ions to form a neutral compound.

c.Sodium can form a +1 charged ion and is written: Na^{+1}

d.Sulfur can form a -2 charged ion and is written: S^{-2}

e.The formula for Sodium Sulfide is Na_2S

f.Some other common ions that you should memorize: K^{+1}, Ag^{+1},Mg^{+2}, Zn^{+2}, Al^{+3}, Ca^{+2}, O^{-2}, Cl^{-1}

g.Click here to see how binary ionic compounds dissolve compared to molecular compounds.

h.Try some examples below:

Calcium Fluoride =	Click here for answer.	Potassium Chloride =	Click here for answer.
Lithium Oxide =	Click here for answer.	Aluminum Sulfide =	Click here for answer.

IV. Polyatomic Ionic Compounds

i. Sometimes a group of atoms can have a charge. This is called a poly atomic ion.

ii. Some common poly atomic ions which you should memorize are: nitrate NO_3^{-1}, sulfate SO_4^{-2}, carbonate CO_3^{-2}, bicarbonate (or hydrogen carbonate) HCO_3^{-1}, and hydroxide OH^{-1}

iii. Notice that the names of these ions end in -ate.

iv. When you see a name ending in -ate it probably implies that it is a polyatomic ionic compound.

v. The groups of atoms can be thought of as a single entity with a charge, just like a single atom can have a charge. For example, Sodium Nitrate needs one +1 sodium ion to neutralize one -1 nitrate ion, so the formula is $NaNO_3$.

vi. If you need more than one polyatomic ion then you put parenthesis around it in the formula. For example, Calcium Nitrate needs one +2 calcium ion to neutralize two -1 nitrate ions, so the formula is $Ca(NO_3)_2$.

vii. Click here to see how polyatomic ionic compounds dissolve.

viii. Try some examples below:

Sodium Sulfate =	Click here for answer.	Zinc Phosphate =	Click here for answer.
Barium Hydroxide =	Click here for answer.	Ammonium Sulfate =	Click here for answer.

V. Ions with multiple charges

i. Some atoms can commonly form 2 or 3 different charges. These atoms are typically transition elements.

ii. Copper, for example, usually forms +1 or +2 charged ions.

iii. This can cause problems if a compound is named Copper Oxide. This could have the formula CuO or Cu_2O depending on the charge of the copper atom.

iv. To clear up this ambiguity we can name the ions by specifically adding on a number to their name. Cu^{+1} is Copper(I) and Cu^{+2} is Copper(II). So the names of the copper compounds listed above are Copper(II)Oxide for CuO and Copper(I)Oxide for Cu_2O.

v. Try some examples below:

Iron(II)Oxide = Click here for answer.	Click here for answer. $=CuSO_4$
Iron(III)Oxide = Click here for answer.	Click here for answer. $= Cr(NO_3)_3$

b. Handout: Solubility Rules and Common Ions

c. Homework: Binary Ionic Naming Sheet

d. Homework: Binary Ionic With Roman Numerals.

e. Homework: Polyatomic Ion Sheet.

f. Molecular Compounds

I. Molecular naming falls into two groups - organic and inorganic. We will talk about inorganic for now. Molecular compounds consist of non-metal atoms.

II. Prefixes which indicate the number of atoms of each element are used in the naming of inorganic molecular compounds. You should memorize the following:

mono- = 1	di- = 2	tri- = 3	tetra- = 4	penta- = 5
hexa- = 6	hepta- = 7	octa- = 8	nona- = 9	deca- = 10

III. When given a formula the prefixes above are applied to the words that would be used to name the compound as if it were ionic. For example, P_2O_3 would be named Phosphorous Oxide if it were ionic, but it consists of two nonmetals, so it would be named Diphosphorous Trioxide.

IV. Whenever there is only one atom of the first element in a formula we drop the term Mono-. For example CO is Carbon Monoxide, not Monocarbon Monoxide.

V. To write formulas you just interpret the prefixes on the names and write the appropriate symbolic representation. For example, Sulfur Dioxide is SO_2.

VI. Try some of the following examples:

Carbon Tetrachloride = Click here for answer.	Click here for answer. $=N_2O$
Trinitrogen Pentoxide = Click here for answer.	Click here for answer. $= CO_2$

VII. There are special cases where we use common names for molecular compounds. The only two that I want you to memorize are: Water = H_2O and Ammonia = NH_3 (not to be confused with the amonium ion = NH_4^{+1}

g. Ban dihydrogen monoxide! DMHO Fact Sheet- Join the movement by clicking here.

h. Acids

i. **Homework:** Naming Various Chemicals Sheet

j. Get some extra practice on naming substances at the ChemTeam website.

iii. General naming and formula writing strategy

I. General naming strategy

i. Determine if the formula depicts an ionic compound (metal and nonmetal), a molecular compound (two nonmetals), or an acid (begins with hydrogen).

ii. If ionic, determine the names of the ions and write the name putting the metal first.

iii. If molecular, determine the names of the nonmetals and add the appropriate prefixes before writing the name.

iv. If an acid, it is either in the form hydro-______-ic acid or it is one of the ones you memorized.

II. General formula writing strategy

i. Determine if the name depicts an ionic compound (metal and nonmetal), a molecular compound (two nonmetals), or an acid (has the word acid in its name).

ii. If ionic, determine the charges on the ions and write a formula that will yield a neutral compound.

iii. If molecular, write a formula using the prefixes in the name to determine the subscripts in the formula.

iv. If an acid, then it is either one of the ones you memorized or it's H__. (hydrogen followed by some single element)

iv. Writing Chemical Equations

I. A chemical equation is a symbolic representation of what happens during a chemical reaction.

II. To describe the reaction you did between baking soda and hydrochloric acid in words you would write: Sodium Bicarbonate reacts with Hydrochloric Acid to produce Carbon Dioxide, Water, and Sodium Chloride.

III. In symbolic form we would write: $NaHCO_3 + HCl \rightarrow CO_2 + H_2O + NaCl$

IV. Everything to the left of the arrow is called the reactants, everything to the right, the products.

V. One can even add the state of each substance in the reaction.

(s) = solid (l) = liquid

(g) = gas (aq) = aqueous (dissolved in water)

Using the above, the reaction becomes: $NaHCO_{3(s)} + HCl_{(aq)} \text{-->} CO_{2(g)} + H_2O_{(l)} + NaCl_{(aq)}$

VI. Reactions with only ionic compounds as reactants.

i. Typically ionic compounds won't react with each other unless they are dissolved in water. Therefore, most of our reactions with ionic compounds will be in the aqueous phase.

ii. A reaction only occurs if one of the products formed would be insoluble in water. When an insoluble compound is formed from a reaction between two aqueous solutions, we call this compound a precipitate. See the Precipitation Rules Sheet to learn if an insoluble compound would form.

iii. When combining two aqueous ionic compounds you basically have four different ions floating around in solution. The positive and negative ions from each compound have the opportunity to come in contact and react. If the new compound formed is insoluble then a precipitate forms.

iv. We can write the reaction between Sodium Chloride and Lead(II)Nitrate in several ways.

a. In words it would be:

Sodium Chloride + Lead(II) Nitrate --> Sodium Nitrate + Lead(II) Chloride

b. In formulas it would be

$NaCl_{(aq)} + Pb(NO_3)_{2(aq)} \text{-->} NaNO_{3(aq)} + PbCl_{2(s)}$ (see an illustration of this below)

c. Notice that the $NaNO_3$ is still dissolved. Basically, the sodium and nitrate ions did not really do anything. They were floating around dissolved in solution before and after the reaction.

VII. Reactions can occur with all different kinds of substances. The one described above is typical of how ionic substances react with each other. However, elements and compounds (ionic, molecular, and acid) also react together, although in more or less predictable ways. For example:

i. $CuSO_{4(aq)} + Zn_{(s)} \text{-->} ZnSO_{4(aq)} + Cu_{(s)}$

ii. $Ca_{(s)} + Cl_{2(g)} \text{-->} CaCl_{2(s)}$

iii. $H_2SO_{4(aq)} + Mg_{(s)} \text{-->} MgSO_{4(aq)} + H_{2(g)}$

iv. $NaHCO_{3(s)} + HCl_{(aq)} \text{-->} NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

a. Demo: Precipitation of NaCl and $Pb(NO_3)_2$

b. Lab: Predicting Precipitates

c. **Homework:** Seven Solution Practice

v. Balancing Equations

I. Remember Lavoisier's Law of Conservation of Mass? So far the chemical equations we have written have not taken this law into consideration.

II. Let's recall the reaction between Sodium Chloride and Lead(II) Nitrate:

$$NaCl_{(aq)} + Pb(NO_3)_{2(aq)} \text{ --> } NaNO_{3(aq)} + PbCl_{2(s)}$$

Where did the second chloride ion come from, and where did the other nitrate ion go?

III. Every atom that appears on the left side of the arrow must also appear on the right side. It might be tempting to fix this problem by rewriting $PbCl_2$ as PbCl, but that would be the incorrect formula for Lead(II) Chloride. Pb^{+2} must pair up with two Cl^{-1}s.

IV. We resolve this by placing coefficients in front of the formulas indicating that you can have different ratios of the substances reacting to form different ratios of products. To fix the above reaction we would rewrite it as:

$$2\ NaCl_{(aq)} + Pb(NO_3)_{2(aq)} \text{ -->} 2\ NaNO_{3(aq)} + PbCl_{2(s)}$$

V. The 2 in front of NaCl give you 2 Nas and 2 Cls. The 2 in front of the $NaNO_3$ gives you 2 Na, 2 N, and 6 O. If you count up all the atoms on the left and right of the arrows you will have the same number of each element. The equation is now balanced.

VI. An unbalanced equation is like having a recipe with no quantities for each ingredient.

a. **Homework:** Balance Practice

b. If you want even more practice on balance equations click here to see the ChemTeam's set of problems.

c. Lab: Seven Solution Lab

vi. Types of Reactions

I. Reactions can placed in broad categories.

i. Synthesis

a. This occurs when the number of products is fewer than the number of reactants.

b. In symbolic form: A + B --> AB

c. An example of this is the formation of water: $2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(l)}$

ii. Decomposition

a. This occurs when the number of products is more than the number of reactants.

b. In symbolic form: AB --> A + B

c. Concrete example: $2\ H_2O_{(l)} \rightarrow 2\ H_{2(g)} + O_{2(g)}$

iii. Single Replacement (or Displacement)

a. This occurs when one element replaces another element in a compound.

b. In symbolic form: A + BC --> AC + B or A + BC --> BA + C

c. Concrete example: $Zn_{(s)} + CuSO_{4(aq)} \rightarrow ZnSO_{4(aq)} + Cu_{(s)}$

iv. Double Replacement (or Displacement)

a. This occurs when two sets of elements switch places in a reaction.

b. In symbolic form: AB + CD --> AD + CB

c. Concrete example: $2\ NaCl_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow 2\ NaNO_{3(aq)} + PbCl_{2(s)}$

d. When the reaction is between an acid and a base (any compound that forms hydroxide ions), water is formed as one of the products. This is called Neutralization. For example:

$H_2SO_{4(aq)} + 2\ NaOH_{(aq)} \rightarrow 2\ H_2O_{(l)} + Na_2SO_{4(aq)}$

v. Combustion

a. Typically combustion occurs when a hydrocarbon reacts with oxygen to produce carbon dioxide and water. Hydrocarbons are a class of compounds that primarily consist of hydrogen and carbon.

b. In symbolic form: $C_xH_y + O_2 \rightarrow CO_2 + H_2O$

c. Concrete example: $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$

II.Some common chemical reactions that you should be familiar with:

a.acid + base ----> water + ionic compound

b.metal + oxygen ---> ionic compound

c.metal + acid ---> hydrogen gas + ionic compound

d.ionic compound1 + ionic compound2 ---> ionic compound3 + ionic compound4

e.acid + carbonate ---> carbon dioxide + water + ionic compound

f.metal1 + ionic compound1 ---> metal2 + ionic compound2

g.hydrocarbon + oxygen ---> carbon dioxide + water

a.**Homework:** Types of Reactions

b.Lab: Common Chemical Reactions

c.Lab: Copper Conversion Lab

IV.Chemical Potential Energy

i.Demo: Various Types of Potential Energy

ii.Types of Potential Energy

I.Potential energy is a type of energy that is "hidden" in some way. It is a type of energy that can be converted to other forms and often is related to some attractive or pushing forces.

i.Elastic Potential Energy

a.Anything that can act like a spring or a rubber band can have elastic potential energy.

b.Let's take the rubber band for example. To stretch the rubber band you have to use energy. That energy has now been turned into elastic potential energy. To get that energy back, just let go of the rubber band and its potential energy is converted primarily into kinetic energy.

c.Springs work the same way, but you can either stretch or compress them. Wind-up watches store potential energy in an internal spring when you wind them and slowly use this energy to power the watch.

ii.Gravitational Potential Energy

a.There is a constant attractive force between the Earth and everything surrounding it, due to gravity.

b.To lift something off the ground it takes energy, so just by lifting an object, that object now has higher gravitational potential energy.

c.Gravitational potential energy is typically converted into kinetic energy (an object falling) before it is converted into any other type of energy.

d. Hydroelectric power is generated this way. As the water falls, it turns a turbine, which pushes electrons around, creating an electric current.

iii. Chemical Potential Energy

a. A chemical bond can be thought of as a spring, or an attractive force between atoms that is much stronger than gravity.
b. Because of this, chemical bonds can have potential energy.
c. Anytime two atoms form a strong covalent or ionic bond or two molecules form a weak van der Waals bond, energy is released, usually in the form of heat and light.
d. Film: Releasing energy in ionic bond formation.
e. Film: Releasing energy in molecular bond formation.
f. The amount of energy in a bond is somewhat counterintuitive - the stronger or more stable the bond, the less potential energy there is between the bonded atoms.

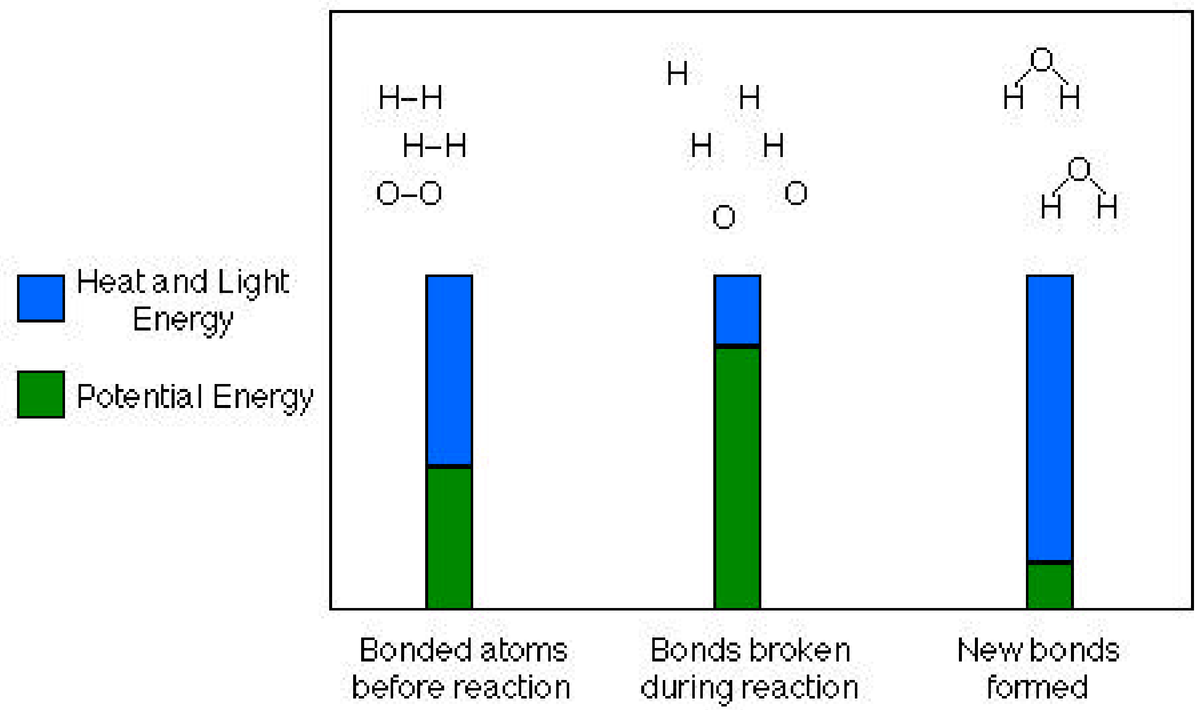
Strong bonds have low potential energy and weak bonds have high potential energy.

g. Lot's of heat and/or light energy is released when very strong bonds form, because much of the potential energy is converted to heat and/or light energy. The reverse is true for breaking chemical bonds. It takes more energy to break a strong bond than a weak bond. The breaking of a bond requires the absorption of heat and/or light energy.
h. During a reaction some chemical bonds are broken and new ones are formed. The trick for how a chemical reaction produces energy is to have the new bonds be stronger or more stable than the old bonds. If this is true, some of the potential energy that was present when atoms were more weakly bonded together must be released as heat and/or light to form the new more stable bonds. A chemical reaction happens in several steps:
    a. First, energy of some form, usually heat or light, is absorbed by two bonded atoms. This causes them to separate, breaking their chemical bond and increasing their chemical potential energy. (Some of the heat or light energy was converted to chemical potential energy.)

b. Then, two different atoms collide with each other. As these two atoms get closer together, they begin to form a chemical bond, decreasing the chemical potential energy and releasing heat and/or light energy.

i. In the example above, the newly bonded atoms often have lower potential energy than the two previously bonded atoms. This means that, after the reaction is done, more of the energy in the system exists as heat or light than before the reaction. See the diagram below of Hydrogen (H-H) and Oxygen (O-O) reacting to form Water (H-O-H):

In chemical notation the reaction is written: $2\ H_2 + O_2$ -------------------------------> $2\ H_2O$



Notice that the proportion of heat and light energy after the new bonds form is greater than it was before. In other words, heat and light energy are released as the reaction is completed. On average the bonds in the Water (H-O-H) molecules are stronger than those of the Hydrogen (H-H) and Oxygen (O-O) molecules. Because the (H-O) bonds in water are stronger, they have less potential energy, therefore some of that chemical potential energy must be converted to heat and light.

Also, note that the reaction could be run in reverse, but heat and light energy must be absorbed and converted to chemical potential energy in order to break up water to form Hydrogen and Oxygen.

Your body stores its energy in the ATP molecule. Explore the process of storing and retrieving energy by going to this web page. (Netscape Required.)

a. **Homework:** Draw potential energy curves for the following situations - an exothermic reaction that is likely to be completely spontaneous, a reaction which would require a constant input of energy to complete, an exothermic reaction with a high activation energy, an endothermic reaction with a low activation energy, a reaction in which the reactants and products contain approximately but not exactly the same chemical potential energy.

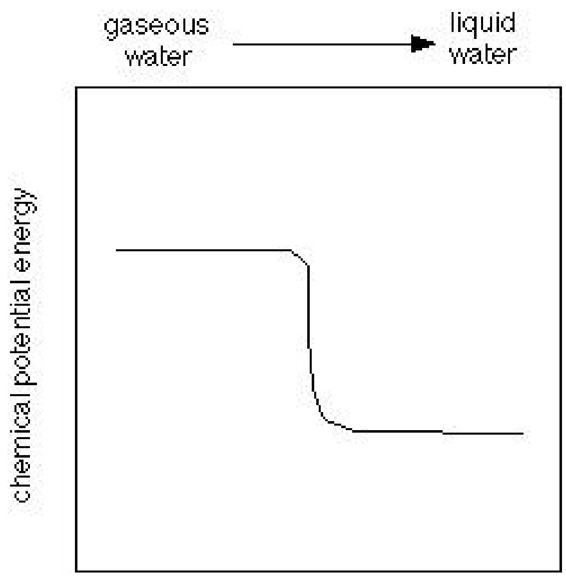
iii. Phase Change and Chemical Potential Energy

I. There is chemical potential energy in the weak van der Waals bonds as well as in the strong ionic and covalent bonds.

II. Because the van der Waals bond/attraction is so much weaker there is much less chemical potential possible here, but it is a major factor in the energy of phase changes.

III. As you can see below, the chemical potential energy of water is less than that of water vapor (or steam) which helps to explain why steam burns can be so severe. Not only is the steam hot, but it tends to condense into a liquid when it touches something. When forming a liquid the chemical potential energy that is released when the van der Waals bonds form can add substantial heat energy to the process of heat transfer from the water molecules to whatever they are condensing on.

IV. See a diagram of the energy below:



V. A process that releases chemical potential energy as heat is called exothermic.

VI. A process that absorbs heat to reduce chemical potential energy (break bonds) is called endothermic.

VII. Going from a gas to a liquid is exothermic because chemical potential energy is released as heat in the process.

VIII. Chemical reactions in which new molecules or ionic substances are formed are also exothermic or endothermic.

a. Lab: Sodium Thiosulfate Lab

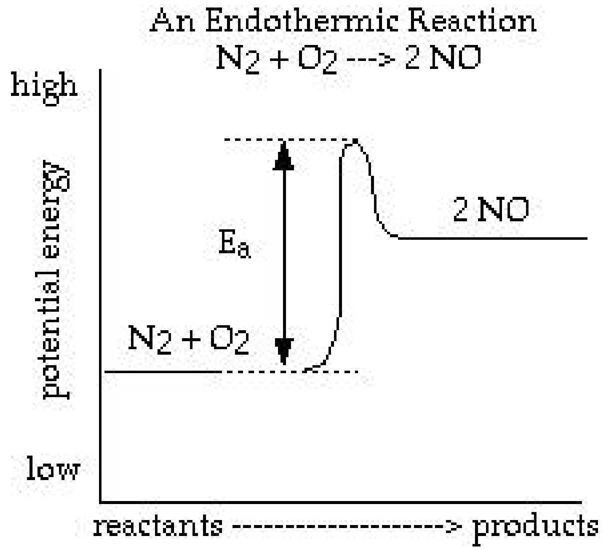
b. Lab: Ice to Steam Lab

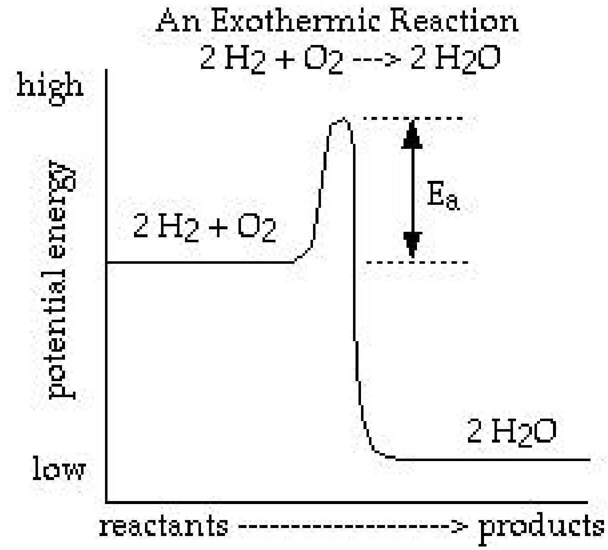
c. Homework: Lab Questions

iv. Activation Energy

I. As mentioned before, to get a reaction to happen bonds usually have to be broken first. This takes some energy, usually in the form of heat (fast moving molecules or atoms colliding) or light.

II. The energy needed to break the initial bonds of the reactants is called the Activation Energy and is abbreviated Ea. Below is a picture of two different reactions and how the chemical potential energy of the substances changes during the reaction.





III. Notice that the endothermic reaction needs a continual input of energy to continue reacting. However, the reaction between H_2 and O_2 gives off so much energy that it can supply the left over reactants with the activation energy needed to form water and complete the reaction. Once this reaction is started with a spark or a flame it continues until there are no more reactants.

IV. Exothermic reactions with a very low activation energy will occur spontaneously, and one's that tend to give off a lot of energy tend to be very reactive or even explosive.

v. Catalysts

I. A catalyst is a substance which speeds up the rate of reaction without being used up in the reaction. You will have the same amount of catalyst at the beginning and end of a reaction, but the reaction will occur much more quickly.

II. Catalyst examples:

i. Most of the enzymes in your body are catalysts. Many of the chemical reactions that are necessary for life to occur run too slowly without being catalyzed by an enzyme. Without catalysts we could not exist.

ii. The catalytic converter that is part of all modern car exhaust systems. This turn many of the pollutants (primarily hydrocarbon (C-H) fragments and carbon monoxide (CO) in the exhaust into carbon dioxide (CO_2) and water (H_2O).

iii. When you generated oxygen from hydrogen peroxide earlier this year you used a Manganese metal catalyst.

$$2\ H_2O_2 \text{ ---> } 2\ H_2O + O_2$$

Notice that the MnO2 is not written as part of this reaction. That is because it is not consumed during the reaction. It can be reused over and over again.

iv. Chlorine atoms from CFCs (chloroflourocarbons) catalyze the breakdown of ozone into oxygen: $O + O_3$ ---> $2\ O_2$

III. See some catalyzed reactions:

i. Film: Forensic Catalysis

ii. Film: Formic Acid Decomposition

IV. A catalyst works by lowering the activation energy necessary to complete a reaction

V. Because it takes less energy to form products, the reaction occurs more quickly.

VI. Film: Catalysts and Activation Energy

VII. Film: Hydrogen Catalysis

V. Handout: Chemical Bonding Review