Periodic Table and Bonding

	Definition
Alkali metal	An element in Group 1 of the periodic table. These elements are extremely reactive.
Alkaline earth metal	An element in Group 2 of the periodic table. These elements are very reactive.
Anion	A negatively charged ion.
Atomic radius	The size of an atom. Sometimes called “covalent atomic radius”.
Brittle	The ability to be crushed into pieces when hammered, a property of nonmetals.
Cation	A positively charged ion.
Diatomic molecule	A nonmetal atom that forms one or more nonpolar covalent bonds with another atom of the same element to form a molecule consisting of the two atoms when there is no other element to bond with. Elements that do this are Br, I, N, H, Cl, H, O and F.
Ductile	The ability to be stretched into a wire, a property of metals.
Dull	The lack of ability to reflect light efficiently, a property of nonmetals.
Group	Columns down the periodic table that denote elements with the same number of valence electrons and similar chemical properties.
Halogen	An element in Group 17 of the periodic table. These elements are extremely reactive.
Ionic bond	A bond formed when a metal atom loses its valence electrons to a nonmetal atom, forming positive and negatively charged ions that attract to each other.
Ionic Radius	The size of an ion compared to the original atom. Metal atoms lose electrons and form + charged ions that are smaller than the original atom, nonmetal atoms form – charged ions that are larger than the original atom.
Luster	The ability to reflect light, a property of metals.
Malleable	The ability to be hammered or rolled into thin sheets, a property of metals.
Metal	Elements that have low electronegativity and ionization energy and large radius that tend to lose electrons to form chemical bonds.
Metallic bond	A bond formed between metal atoms of the same element resulting from the atoms losing electrons to each other and sharing them loosely as a result.
Metalloid	Elements that exhibit properties of both metals and nonmetals.
Molecular orbital	A hybrid orbital made up of the shared unpaired valence electrons of two nonmetallic atoms. This orbital belongs to both of the bonded atoms rather than to any specific atom.
Monatomic molecule	An atom of noble gas, which is considered to be a molecule because there are no unpaired valence electrons.
Noble gas	An element in Group 18 of the periodic table. These elements are nonreactive.
Nonmetal	Elements that have high electronegativity and ionization energy and small radius that tend to gain or share electrons to form chemical bonds.
Nonpolar covalent bond	A bond formed between two nonmetal atoms when unpaired electrons of two atoms are shared equally, with an electronegativity difference of 0 to 0.4.
Nonreactive	Not capable of readily undergoing a chemical change.
Oxidation	The loss of valence electrons from an atom or ion, resulting in the increase in oxidation number of an element.
Period	Rows across the periodic table that denote elements with the same number of principal energy levels.
Polar Covalent bond	A bond formed between two nonmetal atoms when unpaired electrons of two atoms are shared unequally, with an electronegativity difference of 0.5 to 1.7.
Reactive	Capable of readily undergoing a chemical change.
Reduction	The gain of valence electrons from an atom or ion, resulting in the decrease in oxidation number of an element.
Semiconductor	An element that can act as either a conductor or insulator, depending on the situation. Used to manufacture microscopic on-off switches called transistors in computer chips.
Stock system	A method for naming ions of elements that can form more than one possible positive charge by using a Roman numeral after the ion name to denote the ion’s charge.
Transition metal	An element in Groups 3-12 of the periodic table. Many of these elements have colored ions.

The Elements - written by Tom Lehrer, 1959

Antimony (Sb)	Indium (In)	Hafnium (Hf)	Cadmium (Cd)
Arsenic (As)	Gallium (Ga)	Erbium (Er)	Calcium (Ca)
Aluminum (Al)	Iodine (I)	Phosphorous (P)	Chromium (Cr)
Selenium (Se)	Thorium (Th)	Francium (Fr)	Curium (Cm)
Hydrogen (H)	Thulium (Tm)	Fluorine (F)	Sulfur (S)
Oxygen (O)	Thallium (Tl)	Terbium (Tb)	Californium (Cf)
Nitrogen (N)	Yttrium (Y)	Manganese (Mn)	Fermium (Fm)
Rhenium (Re)	Ytterbium (Yb)	Mercury (Hg)	Berkelium (Bk)
Nickel (Ni)	Actinium (Ac)	Molybdenum (Mo)	Mendelevium (Md)
Neodymium (Nd)	Rubidium (Rb)	Magnesium (Mg)	Einsteinium (Es)
Neptunium (Np)	Boron (B)	Dysprosium (Dy)	Nobelium (No)
Germanium (Ge)	Gadolinium (Gd)	Scandium Sc)	Argon (Ar)
Iron (Fe)	Niobium (Nb)	Cerium (Ce)	Krypton (Kr)
Americium (Am)	Iridium (Ir)	Cesium (Cs)	Neon (Ne)
Ruthenium (Ru)	Strontium (Sr)	Lead (Pb)	Radon (Rn)
Uranium (U)	Silicon (Si)	Praeseodymium (Pr)	Xenon (Xe)
Europium (Eu)	Silver (Ag)	Platinum (Pt)	Zinc (Zn)
Zirconium (Zr)	Samarium (Sm)	Plutonium (Pu)	Rhodium (Rh)
Lutetium (Lu)	Bismuth (Bi)	Palladium (Pd)	Chlorine (Cl)
Vanadium (V)	Bromine (Br)	Promethium (Pm)	Carbon (C)
Lanthanum (La)	Lithium (Li)	Potassium (K)	Cobalt (Co)
Osmium (Os)	Beryllium (Be)	Polonium (Po)	Copper (Cu)
Astatine (At)	Barium (Ba)	Tantalum (Ta)	Tungsten (W)
Radium (Ra)	Holmium (Ho)	Technetium (Tc)	Tin (Sn)
Gold (Au)	Helium (He)	Titanium (Ti)	Sodium (Na)
Protactinium (Pa)		Tellurium (Te)	

* as of this date, add Lawrencium (Lr), Rutherfordium (Rf), Dubnium (Db), Seaborgium (Sg), Bohrium, (Bh), Hassium (Hs), Meitnerium (Mt), Darmstadtium (Ds), Copernicum (Cp), Roentgenium (Rg) and a couple of other as of yet unnamed elements.

1) The Periodic Table (HW: p. 15-21)

Essential Question: How did the structure of the atom influence the design of the Periodic Table?

Groups

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H																	He
2	Li	Be											B	C	N	O	F	Ne
3	Na	Mg											Al	Si	P	S	Cl	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	La-Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra	Ac-Lr	Rf	Ha	Sg	Bh	Hs	Mt	Ds	Cp	Rg						

A) Development of the Periodic Table

The Periodic Table was developed by Dmitri Mendeleev and Lothar Meyer in 1869.

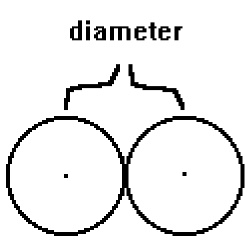
<u>The Modern Periodic Law</u>
The properties of elements are periodic functions or their atomic numbers.

HUH?

As atomic number increases, certain properties, like number of valence electrons, ionization energy and ion charge repeat periodically. Periodically? Yes, it means “at certain intervals”. This is why it’s called a PERIODIC table! Elements are ordered in rows called PERIODS and columns called GROUPS.

Direction	**Importance**	**Examples**
PERIODS (rows)	All elements in the same period have the same number of principal energy levels in their atomic structure	Na, Mg, Al, Si, P, S, Cl and Ar are all in Period 3. They all have three PEL’s in their atomic structure.
GROUPS (columns)	All elements in the same group have the same number of valence electrons, therefore lose or gain the same number of electrons, form similar chemical formulas and have similar chemical properties	Li, Na, K, Rb, Cs and Fr are all in Group 1. They all have one valence electron, they all lose one electron when forming +1 ions, and they are all extremely reactive. They have similar chemical properties and form the following formulas when bonding to oxygen: Li_2O, Na_2O, K_2O, Rb_2O, Cs_2O and Fr_2O.

B) Sizes Of Atoms
Atomic radius - 1/2 the measured distance between two nuclei of the element in the solid phase.



a) ***<u>Within a period of the table, the radius generally decreases as the atomic number increases.</u>*** This is due to an increase in nuclear charge as the atomic number increases, causing an increase in attraction between the nucleus and the valence electrons, pulling them closer together.
Example: In Period 3, the all of the valence electrons are in the same energy level (3). As the atomic number increases, the number of protons increases. Na has 11 protons, Mg has 12, Al 13, Si 14, and so on. As the number of positive protons increases, the attraction between the electrons and the nucleus increases, making each successive atom smaller than the previous one.

b) ***<u>Within a group of the table, the radius increases as the atomic number increases.</u>*** This is due to an additional PEL between the nucleus and the valence PEL, increasing the distance between the valence PEL and the nucleus. The more layers an onion has, the larger the onion is. Therefore, K is larger than Na because K has 4 energy levels, and Na only has 3.

C) Element Types

Elements on the Periodic Table are divided into three subgroups called metals, nonmetals and metalloids (semimetals). Here is the breakdown:

Groups

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H																	He
2	**Li**	**Be**											**B**	C	N	O	F	Ne
3	**Na**	**Mg**											**Al**	**Si**	P	S	Cl	Ar
4	**K**	**Ca**	**Sc**	**Ti**	**V**	**Cr**	**Mn**	**Fe**	**Co**	**Ni**	**Cu**	**Zn**	**Ga**	**Ge**	**As**	Se	Br	Kr
5	**Rb**	**Sr**	**Y**	**Zr**	**Nb**	**Mo**	**Tc**	**Ru**	**Rh**	**Pd**	**Ag**	**Cd**	**In**	**Sn**	**Sb**	**Te**	I	Xe
6	**Cs**	**Ba**	**La-Lu**	**Hf**	**Ta**	**W**	**Re**	**Os**	**Ir**	**Pt**	**Au**	**Hg**	**Tl**	**Pb**	**Bi**	**Po**	**At**	Rn
7	**Fr**	**Ra**	**Ac-Lr**	**Rf**	**Ha**	**Sg**	**Bh**	**Hs**	**Mt**	**Ds**	**Cp**	**Rg**						

Key:

Ba metal

B Metalloid

Cl Nonmetal

These elements have distinctive properties that give them identities.

D) Properties of Atoms on the Periodic Table

<u>Electronegativity: an atom's attraction for electrons in a chemical bond.</u> Elements with small atomic radius have a greater attraction for electrons, and therefore have a higher electronegativity. It is measured on a relative scale, with fluorine having the highest electronegativity (4.0). Electronegativity can be found on Reference Table S.

<u>First Ionization Energy: the energy required to remove the most loosely held valence electron from the atom to form a positive ion when the atom is in the gas phase.</u> This is directly proportional to the electronegativity, because the more tightly an atom is attracted to is electrons, the more energy it is going to require to remove that electron.

In general, metals have low electronegativities and ionization energies, and tend to lose their valence electrons to form + ions when bonding to nonmetal atoms. Nonmetal atoms have high electronegativities and ionization energies, and gain electrons from metal atoms to form - ions, or bond with other nonmetals to form covalent bonds.

<u>Metallic Character: the degree to which an element matches the characteristics of metals.</u> Metals lose electrons and form + ions, therefore elements that have low electronegativity and lose electrons easily have high metallic character.

<u>Nonmetallic Character: the degree to which an element matches the characteristics of nonmetals.</u> Nonmetals gain electrons and form - ions, therefore elements which have high electronegativity and gain electrons easily have high nonmetallic character.

E) Chemistry Of Metals, Nonmetals and Metalloids

Type	EN & IE	Radius	What they do with electrons	Ion Charge	Properties
Metals	Low	Large	Lose (ionic bonding)	Positive	• Excellent conductors of heat and electricity • Malleable (can be hammered or rolled into thin sheets) and Ductile (can be drawn into a wire) • Shiny (has luster) • Make up more than 2/3 of the elements • Metallic character increases as ionization energy decreases. Fr is the most metallic element on the Periodic Table.
Nonmetals	High	Small	Gain (ionic bond) Share (covalent bond)	Negative	• Poor conductors of heat and electricity • Brittle (shatters when struck) • Dull appearance, not shiny like metals
Metalloids	Med.	Med.	Usually share	Either	• Semiconductors (sometimes they conduct, sometimes they don't) • Used for making computer microchips • Luster (like metals) and Brittle (like nommetals)

F) Ions

For every electron an atom gains, it becomes more negatively charged. If an atom gains three electrons when forming a bond, the atom becomes a -3 ion. If an atom loses two electrons when forming a bond, it becomes a -2 ion.

<u>When an atom becomes an ion, it does so by gaining or losing in such a way that it ends up with 8 valence electrons</u> (called a ***<u>stable octet</u>***). The electrons are gained or lost from the valence s and p sublevels. If an element has 1-3 valence electrons, it will opt to lose them. Here is an example of an atom that does this:

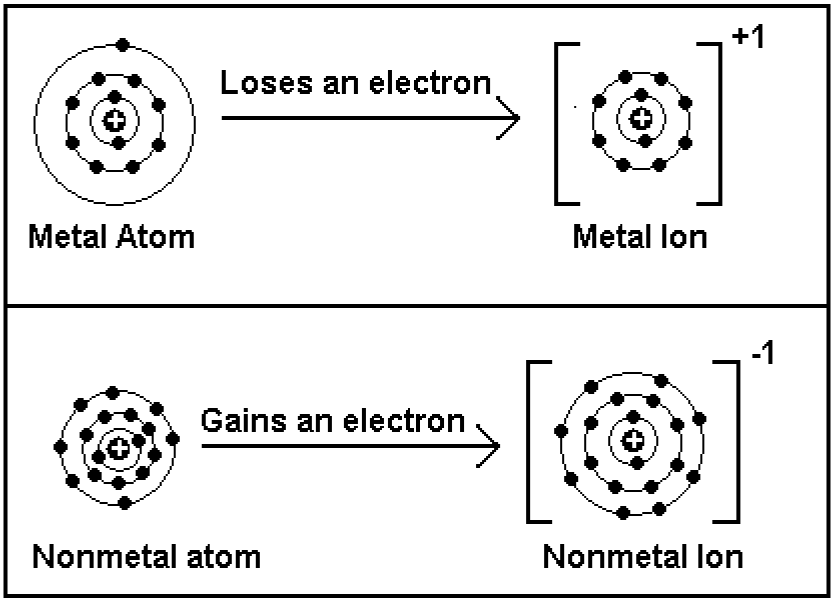
Sodium (Na): **2-8-1** Sodium has 1 valence electron. It can either gain 7 (to form the configuration 2-8-**8**) or lose 1 (to form the configuration 2-**8**). Either way, it ends up with 8 valence electrons. Nature tends to do things that take less energy, and the second option...that of losing the electrons...is what sodium does.

If the atom has from 5 to 7 valence electrons, it will opt to gain electrons, forming negative ions. For example: chlorine (2-8-**7**) has seven valence electrons, so it is no trick at all for it to just gain the one more that it needs to form a -1 ions and have its stable octet (2-8-**8**)

Group	Valence Electron Configuration	# Valence Electrons	How it forms an ion	Charge of ion	Ion Valence Electron Configuration
1	…-1	1	loses 1	+1	…-8
2	…-2	2	loses 2	+2	…-8
13	…-3	3	loses 3	+3	…-8
14	…-4	4	loses 4 (to a more electronegative atom) gains 4 (from a less electronegative atom)	+4 -4	…-8
15	…-5	5	gains 3	-3	…-8
16	…-6	6	gains 2	-2	…-8
17	…-7	7	gains 1	-1	…-8
18	…-8	8	doesn't need to	0	…-8

Note: Elements in the middle can lose electrons from both the valence energy level and the one below it when forming ions. This allows for a wide range of ion charges. Some of these elements can form more than one possible charge. For example, Cu can form charges of +1 or +2, and Fe can form charges of +2 or +3 depending on the circumstances of the reaction.

<u>When an atom gains electrons, its radius increases. When an atom loses electrons, its radius decreases.</u>



Na (2-8-1) loses its one valence electron. Its new electron configuration is 2-8. It now only has two PEL's instead of three. Na^{+1} has a smaller radius than Na^{0} had, because an energy level is lost.

Metal ions have a smaller radius than metal atoms.

Cl (2-8-7) gains one electron to make a stable octet. The one additional electron makes Cl^{-1} (2-8-8) bigger than Cl^{0} because the extra electron increases the repulsion between the electrons in the valence shell.

Nonmetal ions have a larger radius than nonmetal atoms.

G) Naming Elements

The names of the elements are provided for you in the back of your textbook. For the most part, the symbols reflect the names. The exceptions to this are the elements that have roots in antiquity, so they had Greek or Latin names. These are the elements that have been known since antiquity. Here is a listing of those elements:

Element	Symbol	Traditional Name	Element	Symbol	Traditional Name
Antimony	Sb	Stibnum (from Stibnite ore)	Potassium	K	Kalium
Iron	Fe	Ferrum	Copper	Cu	Cuprum ("from Cyprus")
Gold	Au	Aurum ("shining dawn")	Tungsten	W	Wolfram ("heavy stone")
Silver	Ag	Argentum	Tin	Sn	Stannum
Mercury	Hg	Hydrargyrum ("liquid silver")	Sodium	Na	Natrium
Lead	Pb	Plumbium			

H) Naming Ions

The names of ions are dependent on the charge of the ion in question.

<u>Positive ions retain the name of the element. If the atom is capable of forming more than one possible ion (like Fe can form charges of +2 or +3...these are indicated as Oxidation States in the upper right corner of each element box on the Periodic Table), then a Roman numeral is placed after the ion name, signifying the charge.</u>

Examples:

Elements with One Charge **Elements with more than One Possible Charge**

Ion	Name	Ion	Name	Ion	Name
Na^{+1}	sodium	Fe^{+2}	iron (II)	Pb^{+2}	lead (II)
K^{+1}	potassium	Fe^{+3}	iron (III)	Pb^{+4}	lead (IV)
Ca^{+2}	calcium	Cu^{+1}	copper (I)	Cr^{+2}	chromium (II)
Mg^{+2}	magnesium	Cu^{+2}	copper (II)	Cr^{+3}	chromium (III)
Ag^{+1}	silver	Au^{+1}	gold (I)	Sn^{+2}	tin (II)
Al^{+3}	aluminum	Au^{+3}	gold (III)	Sn^{+4}	tin (IV)

Using Roman numerals to identify the ion charge is called the **Stock system**.

Negative ions are named after the element, with the second syllable replaced with the suffix -ide. For example:

Ion	Element Name	Ion Name	Ion	Element Name	Ion Name
O^{-2}	oxygen	Oxide	S^{-2}	Sulfur	sulfide
N^{-3}	nitrogen	Nitride	P^{-3}	Phosphorous	phosphide
H^{-1}	hydrogen	Hydride	Cl^{-1}	Chlorine	chloride

I) Chemistry of The Groups of the Periodic Table

As we have already seen, the groups on the Periodic Table are broken into groups of similar chemical properties. The following table shows the breakdown, and the names of the various groups. A description of the properties of the elements will follow.

Groups

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H																	He
2	Li	Be											B	C	N	O	**F**	Ne
3	Na	Mg											Al	Si	P	S	**Cl**	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	**Br**	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	**I**	Xe
6	Cs	Ba	La-Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	**At**	Rn
7	Fr	Ra	Ac-Lr	Rf	Ha	Sg	Bh	Hs	Mt	Ds	Cp	Rg						

G.1 alkali metals **G.2** alkaline earth metals **G. 3-12** Transition elements **G. 17** halogens **G. 18** noble gases

Group	Name	# Val. e-	Ion Charge	Notes
1	Alkali Metals	1	+1	• Extremely reactive, only found in compounds • Can be extracted from compounds using electricity • Reacts violently with water to form H_2 (g) and a base • *Alkali* means base (as opposed to acid)
2	Alkaline Earth Metals	2	+2	• Very reactive, only found in compounds • Can be extracted from compounds using chemical reactions • Reacts quickly with water for form H_2 (g) and a base • *Alkaline* means base
3-12	Transition Metals	varies	+1 to +7	• Range of reactivities, some are quite reactive, others are nonreactive. • Some can be found in pure form in nature (Cu, Ag, Au) • Ions are colored, so compounds with transition elements in them are often colored. • Many form multiple charges, need the Stock system to name those that do
17	Halogens	7	-1	• Extremely reactive and corrosive, only found in compounds • Can be extracted from compounds using electricity • Reacts violently with metals for form halide compounds (like NaCl)
18	Noble Gases	8	0	• Completely nonreactive, never found in compounds • Xe and Kr can be forced to react with F_2 in the lab • They have a stable octet, so they don't need to bond

PHASES
All elements are solids at 25°C except the following:
Liquids - Hg, Br
Gases - N, O, F, Cl, H, He, Ne, Ar, Kr, Xe, Rn

Which of the following elements is a nonmetallic gas at 25°C?
a) Zn b) Br **c) Cl** d) Hg

Which of the following elements is a metallic liquid at 25°C?
a) Zn b) Br c) Cl **d) Hg**

Allotropes
These different forms arise because of the different conditions that elements may be found.
Examples:
Carbon: coal (amorphous mass), diamond (network bonded crystal), graphite (mineral where the carbon atoms are bonded in weakly connected sheet structures)
Oxygen: O_2 (diatomic oxygen), O_3 (ozone)
Phosphorous: red, white and yellow phosphorous

MOLECULE: a particle made of nonmetals atoms that are covalently bonded together. The number of atoms of each element is consistent from molecule to molecule (for example, a molecule of water ALWAYS contains two hydrogen atoms covalently bonded to one oxygen atom).

MONATOMIC MOLECULES: Since the noble gases do not react with other elements, the individual atoms are considered to be "molecules" of that gas. So, Kr is not just considered to be an atom, but a molecule as well.

He, Ne, Ar, Kr, Xe and Rn do not bond to other elements, so an atom of each of these can also be considered a molecule. A molecule made of one atom…MONatomic! Mono, meaning one!

Which of the following elements will form monatomic molecules?
a) Zn b) Na **c) Ne** d) N

DIATOMIC MOLECULES: Certain elements are found free and uncombined in nature, but are reactive. To become stable, they react with themselves, and are found as diatomic molecules, two atoms of the element bonded to each other. These elements are:

Br_2, I_2, N_2, Cl_2, H_2, O_2, F_2

Br I N Cl H O F

When these elements are free and uncombined with other elements, they are always found in their diatomic state. For example, H + H → H_2, H_2 is more stable than H. These can also be thought of as the "Twinkie Molecules", because they are always found in a two-pack.

Example: $Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$

The H_2 on the right side is diatomic hydrogen, when not bonded to another element, it forms diatomic molecules.

Which of the following elements will form diatomic molecules?
a) Zn b) Na c) Ne **d) N**

2) Bonding (HW: p. 22-24)

Essential Question: What makes atoms stick together to form compounds and molecules?

Chemical bond - results from the competition for valence electrons between two atoms. Chemical bonds are what hold atoms together to make compounds, and what are broken when compounds are decomposed back into the original elements.

IONIC BONDING

1) Occurs between a metal atom and a nonmetal atom

2) The nonmetal atom has a higher electronegativity than the metal atom, and therefore wins the competition for the valence electrons. The nonmetal atom gains electrons from the metal atom, which loses all of its valence electrons to the nonmetal.

3) The number of electrons lost or gained will be the number needed by each atom to form a **stable octet** (8 valence electrons)

4) The metal atom loses electrons (oxidation) and becomes a + charged cation.

5) The nonmetal atom gains electrons (reduction) and becomes a - charged anion.

6) The two oppositely charged ions attract each other. This forms an ionic bond. It is a surface attraction, and may be broken simply by melting or dissolving in water.

To determine for sure if a bond is ionic, look up the electronegativities of the two bonding elements on Reference Table S and take their difference. **If it is above 1.7 then the atom with the higher electronegativity has enough pull to remove the electrons from the atom with the lower electronegativity**.

How does an ionic bond form between Na and Cl?

Element Types? Na: METAL Cl: NONMETAL

EN of Na = 0.9 EN of Cl = 3.2 END = 2.3 (which is above 1.7, ionic)

1) The Na loses its valence electron (oxidation) and becomes Na^{+1} (a cation).
2) The Cl gains the electron that Na lost (reduction) and becomes Cl^{-1} (an anion)
3) The Na^{+1} and Cl^{-1}, having opposite charges, attract and form an ionic bond (NaCl).

This bond can be broken by melting the NaCl (heating it to its melting point gives the ions enough energy to separate from each other, like pulling on two magnets that are stuck together). Dissolving the NaCl in water will also break the bond, causing the NaCl to separate into Na^{+1} and Cl^{-1}, which cling to the water molecules that dissolved them.

PROPERTIES OF IONIC COMPOUNDS:

1) high melting and boiling points (NaCl melts at a temperature of 1074 K, compared to 273 K for water)

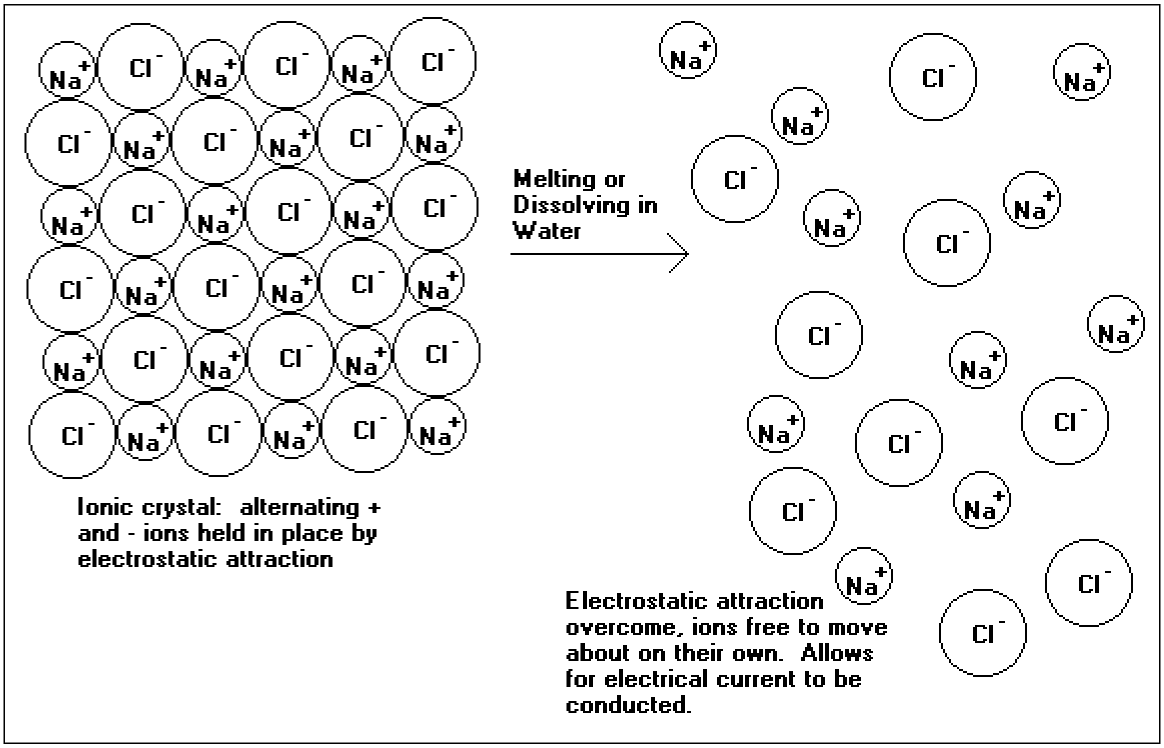
2) low vapor pressure (they don't tend to evaporate, where substances like water evaporate easily)

3) brittle (crushes easily into powder, unlike pure metals, which are malleable)

4) Ionic liquids and solutions conduct electricity because the charged particles are free to move around and carry their electrical charge from one place to another. Ionic solutions (like salt water) are called electrolytes, because they can conduct electricity through them. Electrolytes are used by the body to conduct current through nerves and muscles.

5) Ionic solids do not conduct electricity because the ions are held together in the crystal lattice and can't move around.

Ionic Crystal Structure, then adding heat (or dissolving in water) to break up the crystal into a liquid composed of free-moving ions.



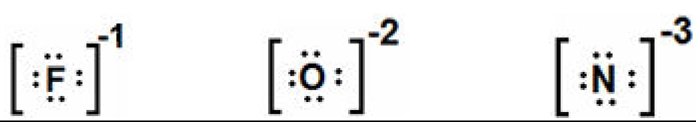
Notice in the SOLID IONIC CRYSTAL, the ions are locked in place, unable to move around. This is why ionic solids can't conduct electricity. When the ions are broken up by adding heat or dissolving in water, the ions are now free. If you place wires connected to a source of electricity (like a battery) into the water, the ions will carry the electrical charge from one wire to the other wire, allowing the electric current to keep flowing. If you keep mail carriers locked up inside of a room, they can't carry mail! If you set them free, they can carry mail to anywhere they need to.

<u>IONIC BONDING DOT DIAGRAMS</u>

When a metal loses its valence electrons, it now has 0 electrons in what was the valence shell. To write the dot diagram of the metal ion, write its symbol, put 0 dots around it, put it in brackets and write the charge outside the brackets, on the upper right side.

$[\text{Li}]^{+1}$ $[\text{Be}]^{+2}$ $[\text{B}]^{+3}$

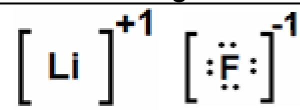
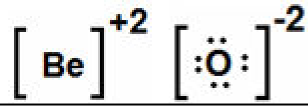
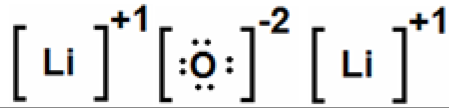
When a nonmetal gains valence electrons, it now has 8 electrons in the valence shell. To write the dot diagram of the nonmetal ion, write its symbol, put 8 dots around it, put it in brackets and write the charge outside the brackets on the upper right side.



To draw the dot diagram of the ionic compound, simply put the ion dot diagrams next to each other so that:

1) The ion charges cancel out (add up to zero)
2) The opposite charged ions are next to each other, and the like charged ions are as far away from each other as they can be.

Example Ionic Bonding Dot Diagrams:

Formula	Dot Diagram
LiF	
BeO	
Li_2O	
BeF_2	

COVALENT BONDING

1) Two NONMETAL atoms attempt to gain each other's valence electrons. They do not have enough difference in electronegativity to do so, therefore they share them.

2) The electrons shared are the unpaired valence electrons.

3) The bonded atoms actually become part of each other. This makes for a bond much stronger than an ionic bond. This bond can not be broken by dissolving in water or melting, so covalent compounds never conduct electricity, regardless of the phase (exception later this year).

To determine if a bond is covalent, check the electronegativity difference. If it is below 1.7, the atom with the higher electronegativity does not have enough pull to remove the other atom's electrons, so they share them. Each bonded atom has 8 valence electrons after bonding, except for hydrogen (H), which is only large enough to have 2 valence electrons after bonding (because of H's single energy level, which can only hold up to 2 electrons.

How many covalent bonds can a nonmetal atom form?

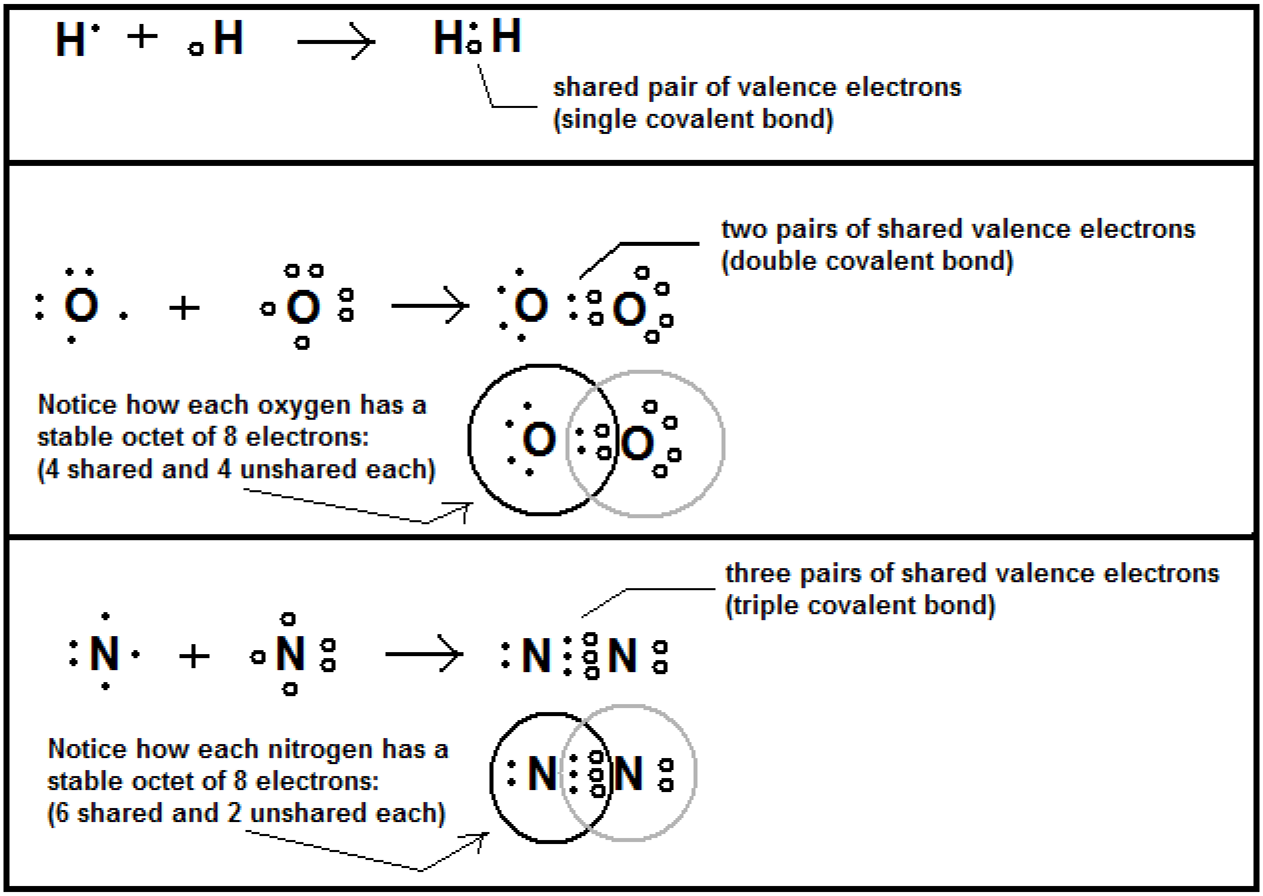
Nonmetal	Dot Diagram	# unpaired e-	# of covalent bonds	Nonmetal	Dot Diagram	# unpaired e-	# of covalent bonds
N	**N**	**3**	**3**	**S**	**S**	**2**	**2**
O	**O**	**2**	**2**	**Cl**	**Cl**	**1**	**1**
F	**F**	**1**	**1**	**P**	**P**	**3**	**3**
C	**C**	**4**	**4**	**Br**	**Br**	**1**	**1**
H	**H**	**1**	**1**	**I**	**I**	**1**	**1**

Making Molecules: Molecules are particles made from nonmetal atoms covalently bonding together. Each molecule of a substance has an identical molecular formula that tells you exactly how many atoms of each element are found in the molecule. H_2O (water) is a molecule made of two atoms of hydrogen bonded to one oxygen atom. CH_4 (methane natural gas) is a molecule made of one carbon atom with four hydrogen atoms bonded to it. NH_3 (ammonia) is made of one atom of nitrogen with three hydrogen atoms bonded to it.

THE TWO TYPES OF COVALENT BONDING

Nonpolar Covalent – formed between nonmetal atoms with an electronegativity difference of 0 to 0.4. The

electrons are being shared equally in the bond. Examples include the diatomic molecules (Br_2, I_2, N_2, Cl_2, H_2, O_2 and F_2) molecules, which form when nonmetal atoms that are unstable by themselves bond together to form more stable molecules, because each bonded atom now has a stable octet of 8 valence electrons.



Polar Covalent - formed between nonmetal atoms with an electronegativity difference of 0.5 to about 1.7. The electrons are being shared unequally in the bond. The electrons spend more time with the more electronegative atom, giving it a slight negative charge, and the less electronegative atom becomes slightly positive. The charged ends of the bonds form POLES (oppositely charged ends), which is why the bond is called "polar".

Writing "partial" all the time is a pain, so use the lower-case Greek letter "delta" instead: δ

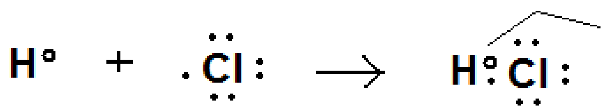
Partially positive = δ +　　　　Partially negative = δ-

Bonding Nonmetal Atoms	Electronegativity Of Each Atom (END)	Which Pole Is δ+ and which is δ-?
H and Cl	**H: 2.2 Cl: 3.2 (1.0)**	δ+ **H-Cl** δ-
H and O	**H: 2.2 O: 3.5 (1.3)**	δ+ **H-O** δ-
O and N	**O: 3.5 N: 3.0 (0.5)**	δ- **O-N** δ+
C and F	**C: 2.6 F: 4.0 (1.4)**	δ+ **C-F** δ-

How Atoms Covalently Bond To Form Molecules

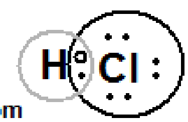
In each case below, notice how each bonding atom's unpaired valence electrons pair up, with the atoms becoming part of one another. Notice how this is different than how ionic bonds work, where electrons are transferred, and not shared. The valence electrons of one bonding atom are shown as dots, the valence electrons of the other bonding atom are shown as circles, so you can clearly see where the electrons in each bond are coming from.

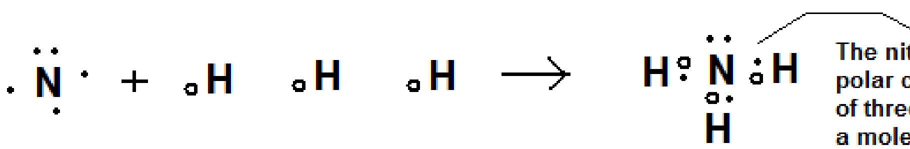
The particle formed by the bonding atoms is called a MOLECULE. Molecules can only be made of nonmetal atoms bonding to each other. Metals cannot be in molecules OK, that is not the whole truth (your blood's hemoglobin is a huge molecule with an iron atom at the center), but as far as this course is concerned, only nonmetal atoms will be in molecules.



The chlorine has formed a single polar covalent bond with a hydrogen atom. This forms a molecule of HYDROGEN CHLORIDE, HCl. When dissolved in water, it forms HYDROCHLORIC ACID, HCl(aq).

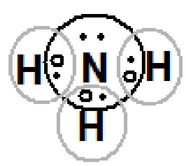
Notice how the chlorine has eight valence electrons, two that are shared with the hydrogen, and 3 unshared pairs of its own (shown here on the top, right side and bottom of the Cl atom). The hydrogen has 2, which is all it needs.

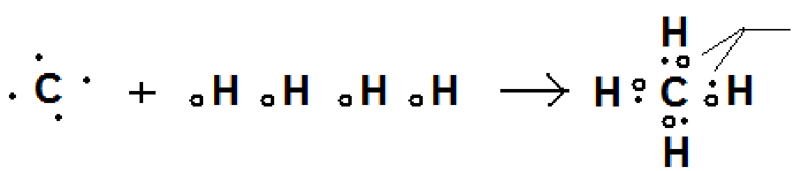




The nitrogen has formed three single polar covalent bonds, one with each of three hydrogen atoms. This forms a molecule of AMMONIA, NH_3.

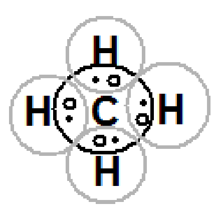
Notice how the nitrogen has eight valence electrons, two that are unshared (the ones on top) and three pairs of shared electrons. Each hydrogen now has two valence electrons (all from shared pairs), which is all the tiny structure of hydrogen can handle.





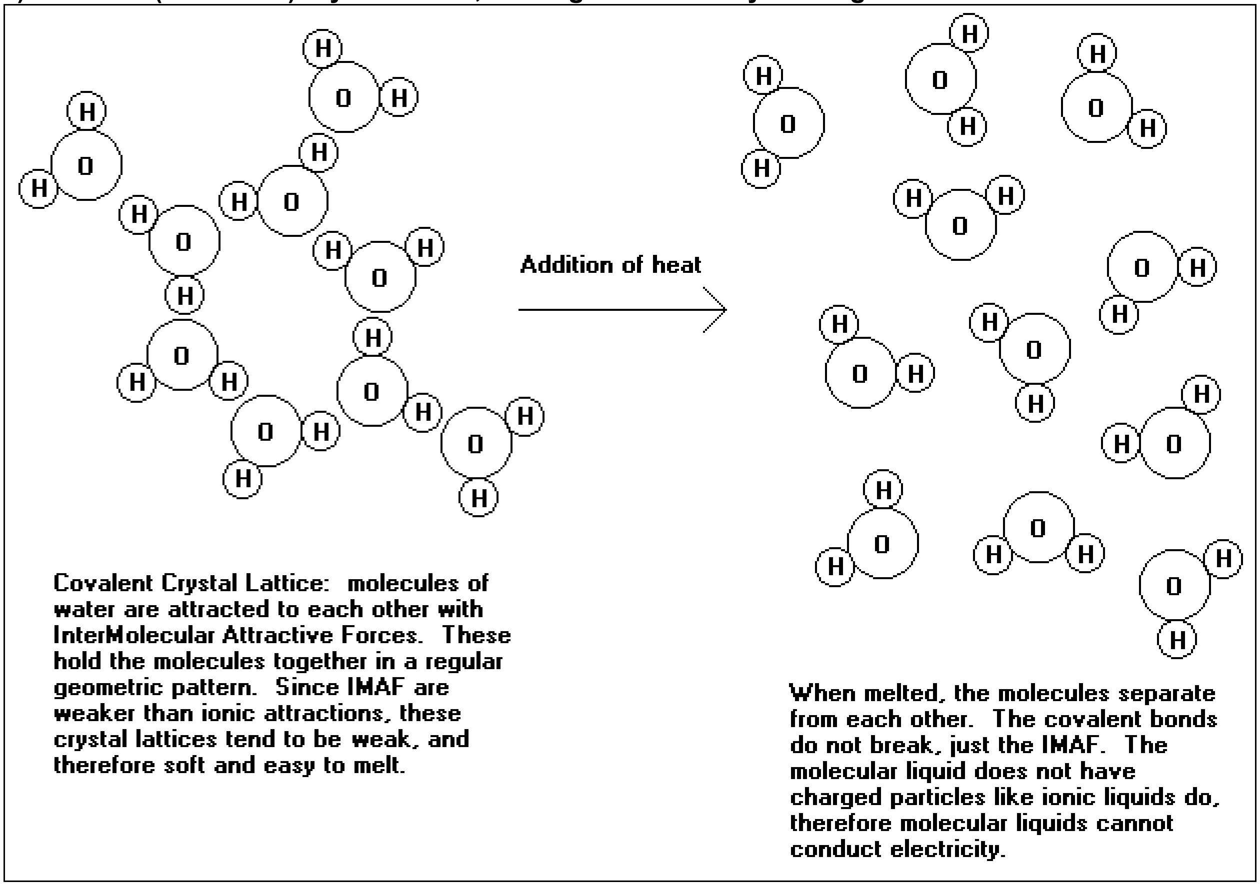
The carbon has formed four single nonpolar covalent bonds, one with each of four hydrogen atoms. This forms a molecule of METHANE, CH_4. Why nonpolar? C has an EN of 2.6 and H is 2.2, which is an END of 0.4. Between 0 and 0.4 is NONPOLAR.

Notice how the carbon has eight valence electrons, all from four shared pairs. Each hydrogen now has two valence electrons, which is all it needs.



When the atoms bond and share their unpaired valence electrons from their individual atomic orbitals, the newly shared pair of electrons forms a <u>molecular orbital</u>, which belongs to both bonded atoms. The bonded atoms, in essence, become part of each other. This makes the covalent bond much more difficult to break than ionic bonds. For example, dissolving an ionic compound in water or melting it is enough to break the ionic bond. Covalent bonds cannot, for the most part, be broken by dissolving or melting.

2) Covalent (molecular) crystal lattice, melting to form freely-moving molecules.



Notice in the SOLID MOLECULAR CRYSTAL, the molecules are locked in place, unable to move around. When the ions are broken up by adding heat or dissolving in water, the molecules are now free. If you place wires connected to a source of electricity (like a battery) into the water, there are no charged ions to carry the electrical charge from one wire to the other wire, so electric current cannot flow. Molecular substances are not made of charged particles, so, unlike ionic compounds, they cannot conduct electricity in any phase. The exception to this are ACIDS, which contain Hydrogen bonded to other nonmetals. The bond is so polar that water molecules can break up the acid molecule into ions, because acids have the highest ionic character of any molecules. Examples of acids include hydrochloric acid (HCl), sulfuric acid (H_2SO_4) and nitric acid (HNO_3). More about acids later in the course. <u>Except for acids, a covalent bond CAN NOT be broken by dissolving in water</u>.

SUMMING IT ALL UP

Bond Type	Bonded Elements	END	Substance is called	How the bond forms
Ionic	M – NM	1.7 +	Ionic	• Metal atom (low EN) loses e^- to nonmetal atom (high EN) • Metal is OXIDIZED, nonmetal is REDUCED • Metal atom forms + ion, nonmetal atom forms – ion • Oppositely charged ions attract • The attraction between the ions is the ionic bond
Covalent	NM – NM	0 – 1.7	Molecular	NONPOLAR COVALENT • Two nonmetal atoms with an END of 0 – 0.4 share their unpaired valence electrons EVENLY • No oppositely charged ends POLAR COVALENT • Two nonmetal atoms with an END of 0.5 or higher share their unpaired valence electrons UNEVENLY • The atom with the lower electronegativity develops a slightly positive charge (δ+) • The atom with the higher electronegativity develops a slightly negative charge (δ-)

More facts about bonding:

Ionic	• Forms ionic crystal lattices with very high melting and boiling points • Electricity is conducted by charged particles. • The ionic bond breaks when the compound is melted or dissolved in water • Breaking the bonds forms ions that are capable, like all liquid particles, of flow • Ionic liquids and solutions are good conductors of electricity • Ionic solids cannot conduct electricity, since the ions are locked in a crystal lattice and are not free to move around
Covalent	• Molecules are particles made up of covalently bonded nonmetal atoms • Since no ions are formed, molecular substances will never conduct electricity (exception to follow much later in the course) • Molecules can have partially charged ends, like a magnet, thanks to polar covalent bonds. Since these partially charged ends have much less charge than ions do, molecular substances have low melting and boiling points • The attraction of the δ^+ end of one molecule for the δ^- end of another is called an intermolecular attractive force • This attractive force is what causes water to have surface tension…so insects can walk across it and so it can form a meniscus when a cup of water is slightly overfilled • Atomic orbitals in each bonded atom combine in such a way that the shared electrons belong to both of the bonded atoms and form molecular orbitals.

Student Name:___________________________________ Grades: _____, _____
PT/30 Bond/20

1) The Periodic Table Homework

A) Development of the Periodic Table

1) Who developed the periodic table? In what year?

____2) In what order are the elements on the periodic table arranged today?
a) atomic number *b) atomic mass* *c) number of neutrons* *d) randomly*

____3) What is the significance of the periods?
a) Number of valence e- *b) Number of PEL's*
c) Number of electrons *d) Number of protons*

____4) What is the significance of the groups?
a) Number of valence e- *b) Number of PEL's*
c) Number of electrons *d) Number of protons*

5) Using the Periodic Table, determine the number of valence electrons in atoms of the following elements, and the Principal Energy Level in which they will be found:

Element	# Valence Electrons	PEL	Element	# Valence Electrons	PEL
Li			Na		
Mg			Ca		
Al			Ga		
Ge			Sn		
N			P		
Se			Te		
Cl			I		
Kr			Rn		

B) Sizes of Atoms

1) Define atomic radius.

____2) According to Reference Table S, as the elements Na to Cl are considered from left to right, what happens to the atomic radius of the atoms?
a) increases *b) decreases* *c) remains the same*

____3) According to Reference Table S, as the elements in Group 2 are considered from top to bottom, what happens to atomic radius?
a) increases *b) decreases* *c) remains the same*

____4) **Explain why the radius of Br is larger than the radius of F.**
a) More electrons *b) More PEL's* *c) More nuclear charge* *d) more neutrons*

____5) **Explain why the radius of O is smaller than the radius of B.**
a) More electrons *b) More PEL's* *c) More nuclear charge* *d) more neutrons*

6) Using Reference Table S to look up atomic radius, determine which of the following pairs of atoms of different elements is larger (remember to use the number of valence electrons and valence PEL as your guide).

Compare…	Which has the larger radius?	Compare…	Which has the larger radius?
N or O		Cl or I	
N or P		Mg or Na	
O or S		Mg or P	
P or S		Li or Be	
P or As		Ca or Ba	

7) Explain why the atomic radius increases as the atomic number decreases for the elements in Period 3.

8) Explain why the atomic radius increases as the atomic number increases for the elements in Group 2.

C) Element Types

1) Identify the following elements as being metals, nonmetals, metalloids or noble gases:

Element	M/NM/ML/NG	Element	M/NM/ML/NG	Element	M/NM/ML/NG	Element	M/NM/ML/NG
K		Ru		Sb		Si	
N		Br		Ne		Fr	

D) Properties Of Elements

1) Define electronegativity.

2) Define first ionization energy.

3) Why does an atom's first ionization energy increase as electronegativity increases?

____4) **What is the relationship between the electronegativity of an atom and its covalent atomic radius?**
a) Direct *b) Indirect* *c) No relationship*

____5) **Which element has the greatest metallic character?**
a) H *b) He* *c) Cs* *d) O* *e) Au* *f) Pb*

____6) **Which element has the greatest nonmetallic character?**
a) H *b) He* *c) Cs* *d) O* *e) Au* *f) Pb*

____7) **As the elements in period 5 are considered from left to right, what happens to the degree of metallic character?**
a) increases *b) decreases* *c) remains the same*

____8) **As the elements in group 14 are considered from bottom to top, what happens to the nonmetallic character?**
a) increases *b) decreases* *c) remains the same*

____9) **When Ca and F bond together, which pair of ion charges are correct?**
a) F^{+}, Ca^{-} *b) Ca^{+}, F^{-}*

10) What another name for a + ion? ________________________

11) What is another name for a - ion?________________________

E) Chemistry of Metals, Nonmetals and Metalloids

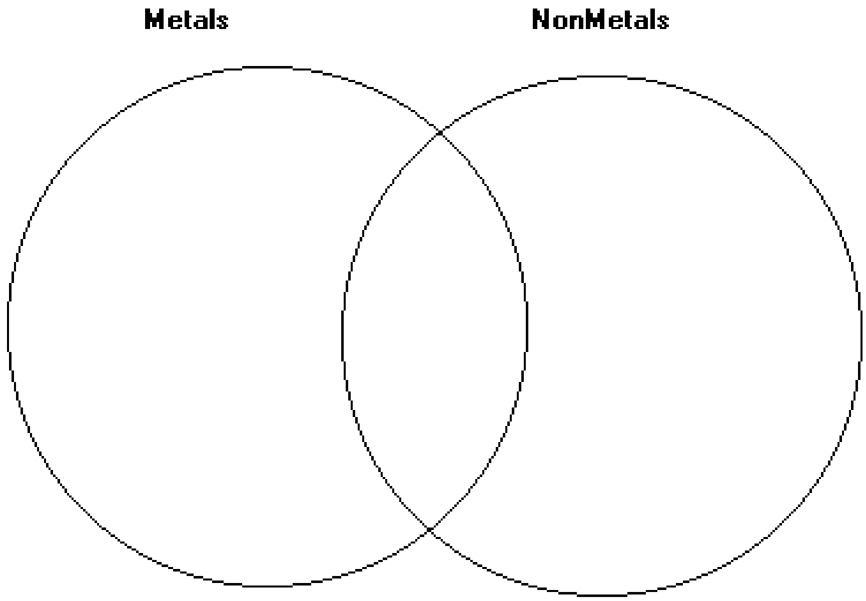
1) Are the following properties those of metals, nonmetals or metalloids? Check to appropriate column:

Property	Metal	Nonmetal	Metalloid
Conductor			
Luster			
Brittle			
Lose electrons			
Nonconductor			
Semiconductor			
Ductile			
Dullness			
Malleable			
Form + ions			
Gain electrons			
Undergo oxidation			
Ion name ends with –ide			
Form – ions			
Undergo reduction			

2) Complete the following chart. If an ion can have more than one IONIC charge, put them all.

Element	M/NM/ML Or NG	# Val E-	# PEL	Phase (S/L/G)	Malleable/ Brittle	Luster/ Dull	Conduct? (Yes, No, Sometimes)	Ion Charge(s)
Ca								
N								
Br								
Cu								
As								
P								
Xe								
Li								
Cl								

3) Complete the following Venn diagram to compare metals and nonmetals, including at least three similarities between the element types and six differences for each element type:



F) Ions

____1) How do metal atoms form ions?
a) gain electrons *b) lose electrons* *c) gain protons* *d) lose protons*

____2) How do nonmetal atoms form ions?
a) gain electrons *b) lose electrons* *c) gain protons* *d) lose protons*

____3) How many valence electrons do all ions have?
a) 1 *b) 2* *c) 6* *d) 8*

____4) When atoms form ions, how many valence electrons will they end up with?
a) 1 *b) 2* *c) 6* *d) 8*

____5) What happens to the radius of an atom when it becomes a + ion?
a) increases *b) decreases* *c) remains the same*

____6) What happens to the radius of an atom when it becomes a - ion?
a) increases *b) decreases* *c) remains the same*

G) Naming Elements

1) List the eleven elements whose symbols derive from the ancient names for these elements by name and symbol.

Name	Symbol	Name	Symbol

H) Naming Ions

1) Name the following ions. Check the Periodic Table to see if you need to use the Stock System for any of them.

Ion	**Name**	**Ion**	**Name**	**Ion**	**Name**	**Ion**	**Name**
Na^{+1}		Cr^{+6}		Co^{+2}		Sc^{+3}	
C^{-4}		Br^{-1}		F^{-1}		I^{-1}	

2) Using Reference Table S, determine which atom in each pair has the higher electronegativity and ionization energy

Compare…	**Higher EN**	**Higher IE**	**Compare…**	**Higher EN**	**Higher IE**
N or O			Cl or I		
N or P			Mg or Na		
O or S			Mg or P		
P or S			Li or Be		
P or As			Ca or Ba		

3) Determine the charge of ions of each of the following elements and indicate if the ionic radius is larger or smaller than the original atom.

Ion	**Charge**	**Ion larger or smaller than original atom?**	**Ion**	**Charge**	**Ion larger or smaller than original atom?**
Li			Mg		
Na			Ca		
K			Sc		
H			N		
P			O		
S			F		
Cl			Br		

4) Name the following ions:

Ion	**Name**	**Ion**	**Name**	**Ion**	**Name**
Na^{+1}		Au^{+3}			oxide
Br^{-1}		Mg^{+2}			lead (IV)
Zn^{+2}		Cu^{+1}			chromium (III)

I) Chemistry of the Groups of the Periodic Table

1) Identify the group that a particular element belongs to based on the properties:

Property	**Group Number**	**Group Name**
Highly corrosive nonmetallic gas, only found in compounds		
Metal that is only found in combination with another elements, has an ionic charge of +2		
Completely nonreactive nonmetallic gas, not isolated in any compound		
Extremely reactive metal, reacts violently with water		

2) How can you tell a transition metal compound just by looking at it?

3) List the elements that are gases at STP.

4) List the elements that are liquids at STP.

5) List the two allotropes of oxygen by name and formula.

6) Define *molecule*.

7) List the *monoatomic* molecules.

8) List the *diatomic* molecules.

____9) Which of the following elements exists as monoatomic molecules at STP?
a) Li b) Br c) Ne d) Au

____10) Which of the following elements exists as diatomic molecules at STP?
a) Li b) Br c) Ne d) Au

2) Bonding Homework

A) Ionic Bonding

_____1) Which of the following compounds is formed by ionic bonding?
a) C_2H_6 b) NO_2 c) Li_2O d) O_2

How were you able to tell this? Explain in terms of *electronegativity difference*: ________________________________

_____2) Which species will conduct electricity (is an electrolyte)?
a) NaCl (s) b) N_2 (s) c) LiF (aq) d) CaI_2 (s)

What gave it away?____

3) Fill in the missing information:

When K and Cl atoms bond together:

K __________ (Gains, loses) ___ (#) valence electron and becomes a _______ (charge) ion, called a(n)___________ (anion, cation) via ________________ (oxidation, reduction).

Cl__________ (Gains, loses) ___ (#) valence electron and becomes a _______ (charge) ion, called a(n)___________ (anion, cation) via ________________ (oxidation, reduction).

4) Draw the dot diagram of the *ionic* compound KCl.

5) Draw the dot diagram of the *ionic* compound CaF_2.

6) Qualitative analysis is used to determine what type of substance you have and what the basic composition is. It involves making use of the known properties of the different kinds of substances. You have learned about the properties of ionic compounds. Let's suppose you come across an unlabeled bottle of substance at work, or perhaps a small spill of this substance, and you need to identify what that substance is. You suspect it might be ionic (like salt), because it is brittle, not shiny like a metal, and forms little crystals. Sugar looks a lot like salt, but it is not ionic. Sugar is molecular, and when it dissolves in water, it does not conduct electricity. Sugar also has a much lower melting point than salt does…in fact, when sugar is heated in air, it starts to turn brown (caramelize). You could taste the substance to see, but there are plenty of other substances that form small white crystals…sodium cyanide, for instance, which would kill you within moments of tasting it.

a) Briefly describe an experiment that can perform to determine whether a substance is ionic or not:

b) What would the expected outcome of the experiment you just suggested be if the substance was ionic?

B) Covalent Bonding

_____1) Which of the molecules listed below has the most polar bond between the bonded atoms?
a) HF b) HCl c) HBr d) HI

How were you able to tell this? Explain in terms of *electronegativity difference*:________________________________

_____2) Which of the following compounds is formed by covalent bonding?
a) Na_2S b) $AlCl_3$ c) $C_6H_{12}O_6$ d) LiH

How were you able to tell this? Explain in terms of *types of elements*:________________________________

_____3) Which of the following molecules contains a nonpolar covalent bond?
a) H_2O b) HF c) F_2 d) NH_3

How were you able to tell this? Explain in terms of *electronegativity difference*:________________________________

_____4) Which of the following molecules contains a polar covalent bond?
a) H_2 b) PH_3 c) F_2 d) NH_3

How were you able to tell this? Explain in terms of *electronegativity difference*:________________________________

5) When an atom of H and an atom of F bond together:

The H will be partially _______ (charge), because it has _____________ (higher, lower) electronegativity than F.

The F will be partially _______ (charge), because it has _____________ (higher, lower) electronegativity than H.

6) Hydrogen and sulfur atoms combine to form a molecule of H_2S, called hydrogen sulfide.

a) Show, using the chart below, how hydrogen and sulfur combine to form a molecule of hydrogen sulfide:

Dot diagram of Hydrogen atom	**Dot diagram of Sulfur atom**	**Dot diagram of Hydrogen atom**	→	**Dot diagram of Hydrogen Sulfide**

b) How can you tell this molecule is formed by covalent bonding and not ionic bonding? Explain in terms of *electronegativity difference*.

c) Are the bonds between hydrogen and sulfur polar or nonpolar? Explain, in terms of *electronegativity difference*.

7) **Complete the following chart, drawing the dot diagram of each element in the molecule and then the dot diagram of the molecule. If the formula is H_2O, make sure you have two atoms of H and one of O in your dot diagram of the molecule.**

Formula	Dot diagram for:	Dot diagram for:	Dot Diagram of Molecule
F_2	F	F	
N_2	N	N	
HBr	H	Br	
H_2O	H	O	
NH_3	N	H	

8) Identify the following bonds as being polar covalent or nonpolar covalent. For the polar covalent bonds, label the δ^+ and δ^- ends.

Bond	END	Polar or Nonpolar?	If polar, label the δ^+ and δ^- ends
H – H			H – H
H – C			H – C
H – Cl			H – Cl

9) Regarding the nonmetallic element, *oxygen*:

a) Oxygen atoms have ____ unpaired valence electrons. This means that they can form ____ covalent bonds.
# #

b) Two O atoms have ____ unpaired electrons between them. When they form O_2, they share ____ electrons.
# #

c) Two O atoms form ____ bonds between them, also known as a __________________ bond.
 # type

d) Each covalent bond is a __________ of shared unpaired valence electrons that form ____________ orbitals.